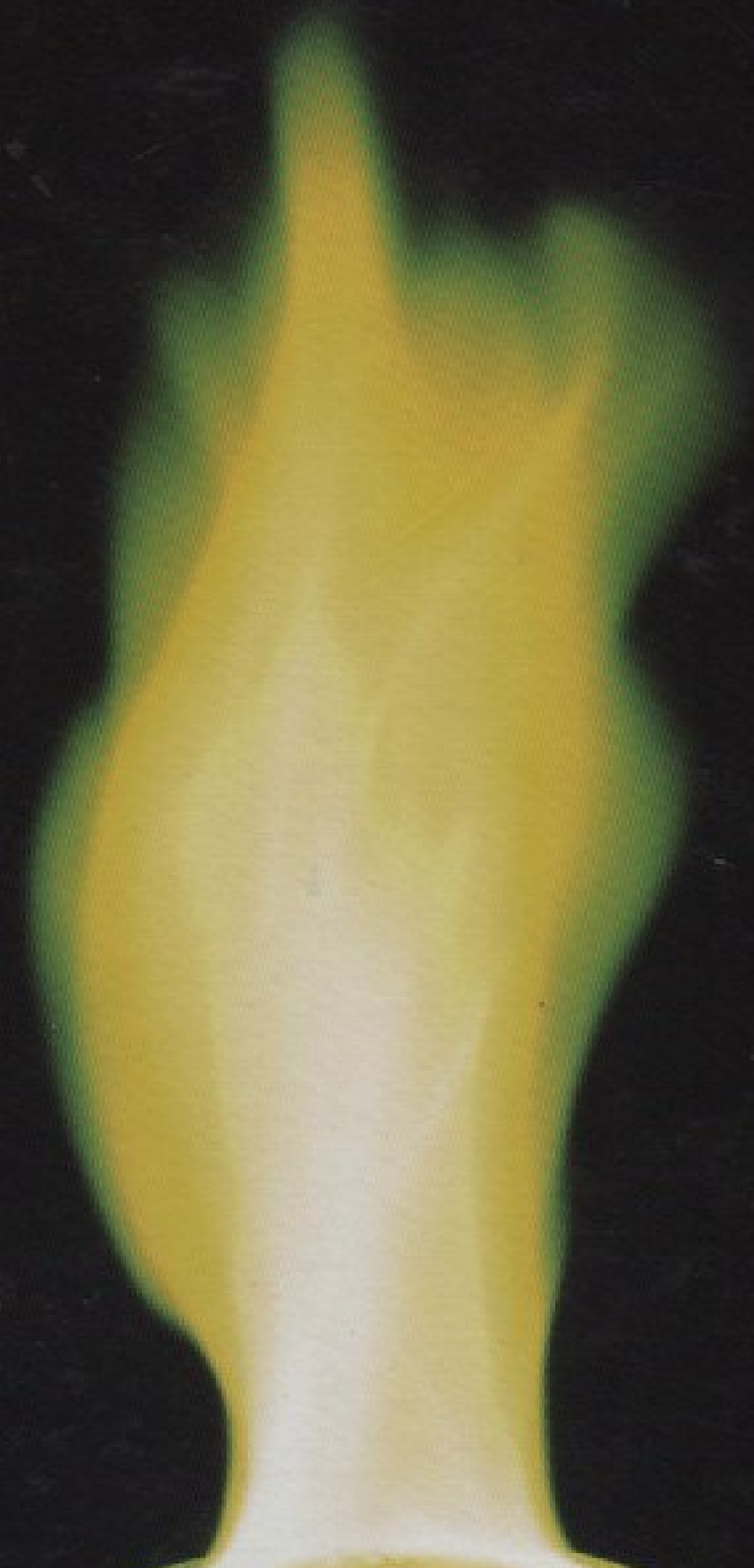
NELSON

UNIVERSITY PREPARATION

CHEMISTRY 11



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Printer
Transcontinental Printing, Ltd.

ISBN-13: 978-0-17-650432-8
ISBN-10: 0-17-650432-X

Printed and bound in Canada
2 3 14 13 12 11

For more information contact Nelson Education Ltd., 1120 Birchmount Road, Toronto, Ontario M1K 5G4. Or you can visit our Internet site at http://www.nelson.com.

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Contents

UNIT 1

Matter, Chemical Trends, and Chemical Bonding

OVERALL EXPECTATIONS

- analyze the properties of commonly used chemical substances and their effects on human health and the environment, and propose ways to lessen their impact
- investigate physical and chemical properties of elements and compounds, and use various methods to visually represent them
- demonstrate an understanding of periodic trends in the periodic table and how elements combine to form chemical bonds

BIG IDEAS

- Every element has predictable chemical and physical properties determined by its structure.
- The type of chemical bond in a compound determines the physical and chemical properties of that compound.
- It is important to use chemicals properly to minimize the risks to human health and the environment.

UNIT TASK **PREVIEW**

In the Unit Task you will research the properties of a conventional consumer product and the processes involved in its manufacture. You will then design a greener version of the product and pitch your product to potential investors.

The Unit Task is described in detail on page 134. As you work through the unit look for Unit Task Bookmarks to see how information in the section relates to the Unit Task.

EXPLORING GREEN CHEMISTRY

Look around you. Think about all of the different substances that you deal with on a daily basis. Consider your notebook or the materials that help to deliver power to your calculator. How was the cover of your notebook made? What will happen to it when you throw it away? What materials were used to make the batteries or the solar cells in your calculator? What will happen to these materials once they no longer work properly?

We are surrounded by a huge variety of materials with a wide range of chemical and physical properties. We are continually developing processes to create products with desired properties. Our ability to manufacture materials provides many conveniences, but also some drawbacks. We are using raw materials at an alarming rate. We are also discarding hazardous chemicals into our landfills or incinerating them and releasing chemicals into our atmosphere.

We are starting to understand more about the hazards of chemical processes. Industries in many countries are now paying attention to the fuels and raw materials they use and the by-products they release along with their intended products. Sometimes manufacturers simplify processes to fewer steps. This has many benefits, including reducing waste. Processes are burning less fuel, using fewer toxic reactants, and releasing fewer unwanted by-products. Industry is attempting to become "greener."

"Green chemistry" is a movement to make industries that involve chemicals more environmentally friendly and sustainable. Green chemistry asks the question: "Why generate pollution if there is a greener alternative?"

Developing a green alternative begins with considering the hazards of the required chemicals as well as their properties. Chemists then develop a manufacturing process so that every stage of product development is environmentally safe—from the raw materials to what happens to the product at the end of its useful life. In other words, the process is "benign by design."

A green chemistry solution may involve using safer chemicals. Liquid carbon dioxide, for example, is starting to replace toxic organic solvents used in dry cleaning. Greening a chemical process can also involve making a process more efficient. For example, the original makers of ibuprofen, an important pain reliever, found a way to make the drug in half the number of steps. The result is a process that generates less waste, uses less energy, and is more profitable!

As you progress through this unit, you will see how the principles of green chemistry can reduce the ecological footprint of some processes.

Questions

1. What can you do to "green" your day? How do these suggestions reduce your ecological footprint?
2. Brainstorm the criteria for a "green" product.
3. What manufacturing industries do you know about that affect the environment or human health in a negative way? What aspects of these industries could perhaps be improved? Brainstorm possible solutions to these problems.

CONCEPTS

- classification of matter
- Bohr–Rutherford model of the atom
- the periodic table
- ionic versus molecular compounds
- atoms versus molecules
- formation of ions

SKILLS

- draw Bohr–Rutherford diagrams for elements
- follow safety procedures in the laboratory
- write chemical formulas and name compounds

Concepts Review

1. Construct a concept map that shows how matter can be classified. Include the following terms: pure substance, mixture, element, compound, heterogeneous, homogeneous, and solution. K/U C
2. Classify each of the following substances as an element, a compound, or a mixture: K/U
   (a) pure water
   (b) pure silver
   (c) Caesar salad
   (d) salt water
   (e) butane
3. Answer the following questions about the periodic table: K/U
   (a) Distinguish between the terms "group" and "period."
   (b) How can we distinguish between a metal and a non-metal on the periodic table?
   (c) To which chemical family does each of the following elements belong?
   F Mg Xe Na
4. How many electrons are there in the outermost orbit of an element in
   (a) the alkali metal family?
   (b) the halogen family?
   (c) the noble gases?
   (d) Group 2? K/U
5. Distinguish between a cation and an anion. K/U
6. Distinguish between the terms "atomic number" and "mass number." K/U
7. List three physical properties and one chemical property for each of the following substances. K/U
   (a) liquid water, $H_2O(l)$
   (b) oxygen gas, $O_2(g)$
   (c) octane (the main component in gasoline), $C_8H_{18}(l)$
8. Draw Bohr–Rutherford diagrams for the following atoms: T/I C
   (a) Li-7
   (b) C-14
   (c) Cl-35
   (d) Ca-40
9. In your notebook, match each chemical name in the first column with the correct chemical formula from the second column. K/U C

(a) lead(II) nitrate	(i) $H_2O(l)$
(b) carbon tetrachloride	(ii) $HCl(aq)$
(c) sodium hydroxide	(iii) $NaOH(s)$
(d) hydrochloric acid	(iv) $Fe_3(PO_4)_2(s)$
(e) dihydrogen monoxide	(v) $Pb(NO_3)_2(s)$
(f) iron(II) phosphate	(vi) $CCl_4(l)$

10. (a) What does WHMIS stand for?
    (b) What does HHPS stand for?
    (c) Identify the hazard indicated by the label in **Figure 1**, and describe the precautions you would take in handling this product. K/U T/I

Figure 1 A can of degreasing solvent

Skills Review

11. In your notebook, copy and complete **Table 1**. You may need to refer to the periodic table on the inside back cover of this textbook. K/U T/I

12. (a) Use the data from Table 1 to draw a graph of atomic number versus mass number. Be sure to clearly label your axes and title your graph appropriately.
    (b) What type of relationship is shown in this graph? T/I C

13. Copy and complete **Table 2** in your notebook. K/U T/I

14. Describe all the safety precautions necessary for working with each of the following items in the laboratory: T/I C
    (a) dilute sulfuric acid
    (b) a Bunsen burner
    (c) a hot plate
    (d) a sample of hydrogen gas

15. Write out a clear procedure that would help a classmate learn how to light a Bunsen burner safely. K/U C

16. Sketch a floor plan for your laboratory/classroom. On your plan, indicate the locations of each of the following items: T/I C
    - first-aid kit
    - eyewash station
    - fire extinguisher
    - fire blanket and/or shower
    - waste container for broken glass
    - chemical safety goggles
    - lab aprons
    - fire exit
    - phone

17. Describe a standard test for each of the following gases: K/U T/I
    (a) oxygen
    (b) hydrogen
    (c) carbon dioxide

Table 1 Subatomic Particles for Selected Atoms or Ions

Name of atom or ion	Symbol	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
	C				6	
magnesium atom			25			
		16			16	
	S^{2-}		32			18
sodium ion					12	

Table 2 Descriptions of Selected Ionic and Molecular Compounds

Compound name	Chemical formula	State at room temperature	Soluble in water?	Physical appearance	Ionic or molecular?
sodium chloride					
	CO_2				
	H_2O		n/a		molecular
methane					
	$NaHCO_3$		yes	white powder	

CAREER PATHWAYS **PREVIEW**

Throughout this unit you will see Career Links in the margins. These links mention careers that are relevant to Matter, Chemical Trends, and Chemical Bonding. On the Chapter Summary page at the end of each chapter you will find a Career Pathways feature that shows you the educational requirements of the careers. There are also some career-related questions for you to research.

CHAPTER 1 Atomic Structure and the Periodic Table

KEY CONCEPTS

After completing this chapter you will be able to

- demonstrate an understanding of the need to use chemicals safely in everyday life
- explain the relationship between the atomic number and the mass number of an element
- differentiate between isotopes and radioisotopes of an element, and describe some of their uses
- explain the relationship between atomic mass and isotopic abundance
- state the periodic law, and explain how patterns in the electron arrangement and forces in atoms result in periodic trends

How Does Atomic Structure Affect an Element's Properties?

In order to understand the nature of matter, we try to imagine what the smallest pieces of matter look like. To what extent can we actually "see" individual atoms or molecules? Our sense of vision is quite limited when it comes to viewing the atomic world. Scientists use tools, such as the electron microscope or the scanning tunnelling microscope, to look at such small entities. At present, these devices give us a hint of what individual atoms might look like. However, this technology does not allow us to peek inside to see how these atoms are put together.

The image on the facing page was created by a scanning tunnelling electron microscope. It shows a tiny sample of gold on a graphite carbon surface. The gold atoms show up as yellow, red, and black. The carbon atoms appear green. The image is beautiful and intriguing but gives us a limited amount of information. If we are to connect the structure of an atom with its properties, we need to know more details about the atomic structure. We need to know how the individual subatomic particles—the particles that make up the atom—are arranged.

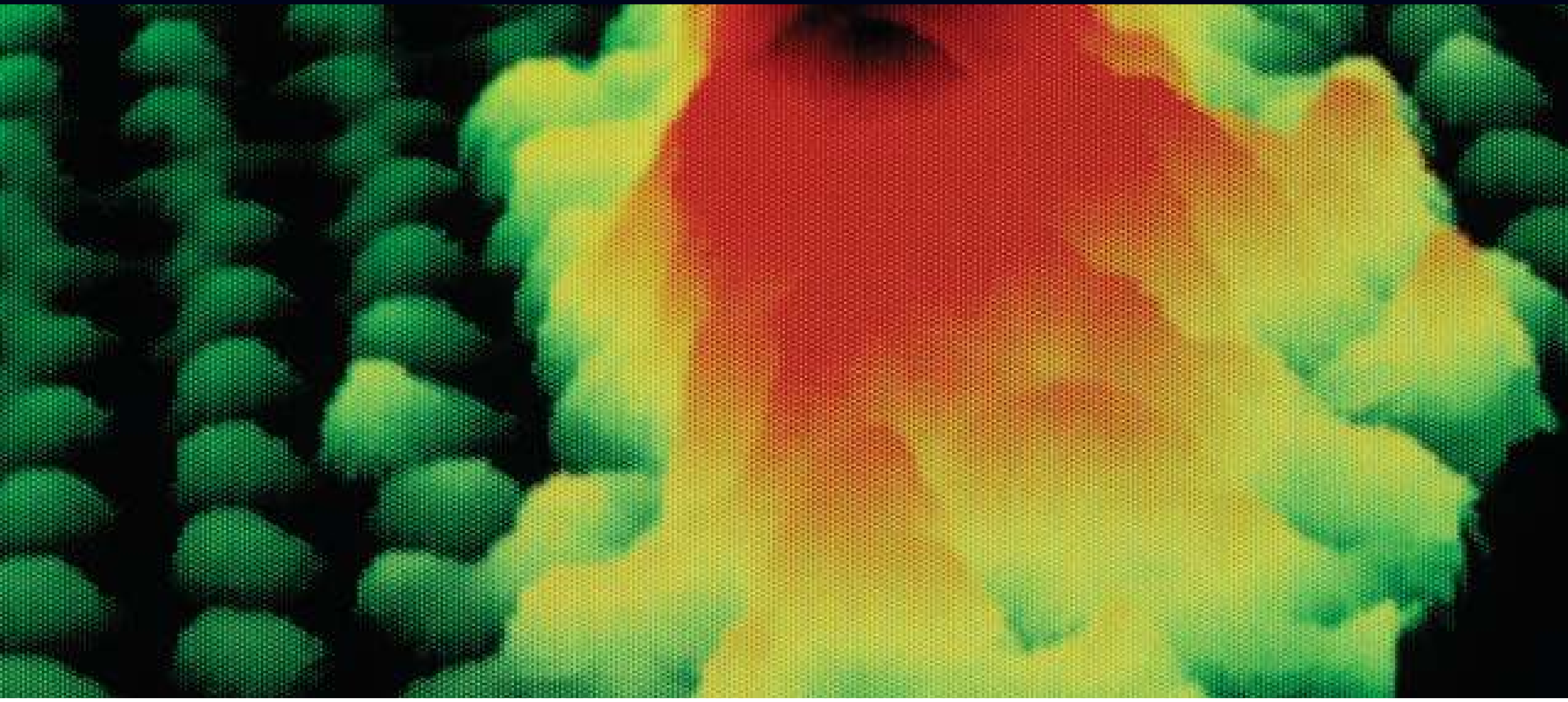
Consider a sample of charcoal made of pure carbon. Imagine that you divide your sample into smaller and smaller pieces in order to see what one individual atom of carbon looks like. From your previous studies in chemistry, you most likely have a picture in your mind of an atom of carbon. In science, a mental picture of something is called a model. Now consider a sample of pure magnesium. Again, imagine that you are able to see what an atom of magnesium looks like.

Now, consider some of the physical and chemical properties of carbon and magnesium. Why does carbon primarily exist in three different forms: charcoal, graphite, and diamond? Why can carbon form so many different types of molecular compounds? Why is magnesium a lustrous, silvery metal? What is the difference, at the atomic level, between a magnesium atom and a magnesium ion? Why and how do magnesium atoms become ions? Why does magnesium burn violently with oxygen to produce a brilliant white light? As we further develop our model of the atom, we will be able to answer some of these questions. With a good atomic model and some simple data, we can predict the properties of substances. Understanding their properties allows us to select appropriate substances for specific applications.

STARTING POINTS

Answer the following questions using your current knowledge. You will have a chance to revisit these questions later, applying concepts and skills from the chapter.

1. What does your model of a carbon atom look like?
2. How does your model of magnesium differ from your model of carbon?
3. How accurate is your model of the atom? Explain your answer.
4. How do models of the atom help us explain the physical and chemical properties of elements?
5. Why would it be useful to be able to "see" atoms and molecules?
6. Can we predict properties for elements yet to be discovered? Explain.



Mini Investigation

Mystery Tube of Science

Skills: Hypothesizing, Performing, Observing, Analyzing, Communicating

SKILLS HANDBOOK A2

This activity represents how scientific investigation leads to the development and evolution of theories. Your teacher will demonstrate the mystery tube (**Figure 1**). Watch carefully and try to develop a mental model—a theory—of the structure inside the mystery tube.

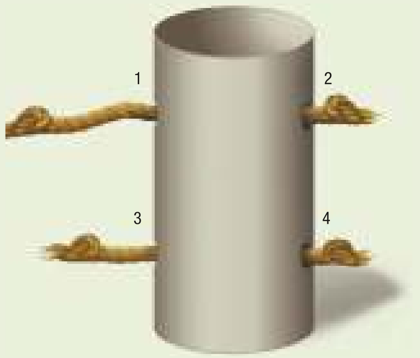


Figure 1 The mystery tube of science will make you think!

Equipment and Materials: mystery tube of science with ropes labelled 1 to 4; cardboard bathroom tissue tubes; string; scissors

1. Your teacher will pull one rope coming from the tube. Observe and record the result of this action.
2. Based on this observation, develop a mental model of the inside of the mystery tube.
3. Your teacher will now pull on a different rope. Observe and record your observations. If necessary, revise your model of the inside of the tube.
4. Your teacher will perform one more experiment. Revise your model as necessary.
5. Working in groups, discuss what you think is happening inside the tube. Build your own mystery tube with the materials supplied. Try to duplicate your teacher's tube.
6. Design a different mystery tube of your own. Challenge other groups to discover the tube's mysteries.

A. Summarize the properties of the mystery tube by listing each cause (in Steps 1, 3, and 4) and its effect. T/I

B. Do you think you would be able to find out more about the mystery tube if you could perform experiments on it directly, rather than by watching your teacher manipulate it? Explain your answer. T/I

C. How does your answer to B relate to the work of scientists investigating the structure of the atom? T/I

D. Research various magic tricks on the Internet that provide a similar illusion. Choose one, and design your own demonstration. T/I A

GO TO NELSON SCIENCE

1.1 The Nature of Chemistry

matter anything that has mass and takes up space

The study of chemistry explores matter and its interactions. **Matter** is defined as anything that has mass and occupies space. Matter might include the food you eat, the cosmetics and grooming products you use, the battery in your cellphone, or the gasoline in your car engine. We encounter thousands of different chemicals every day. The potential applications of chemistry are therefore enormous.

What Is Chemistry?

It might be an interesting challenge to turn this question around and ask, "What *isn't* chemistry?" Indeed, we are hard pressed to think of many things that do not relate to chemistry and its applications.

Chemistry is often called the central science. Chemistry helps to connect the physical sciences (physics) and the life sciences (biology). Concepts in chemistry are often based on an understanding of the laws of physics. For example, the concept of ionic bonding is based on the understanding that opposite electrical charges attract one another. These chemical concepts can, in turn, be extended to help explore concepts in biology and environmental science. For example, once we understand what ions are and how they behave, we can apply this knowledge to the human body. This helps us understand and predict how many body systems work.

The study of chemistry is subdivided into various branches. Inorganic chemistry, organic chemistry, nuclear chemistry, biochemistry, and physical chemistry are some of the different areas of chemistry.

Chemists ask questions about the nature of matter and how it interacts. They design and conduct experiments, collecting evidence in an attempt to answer these questions. They analyze and evaluate data. Scientists eventually develop explanations that help them to understand what is taking place. These explanations may involve changes on a very small scale. Often scientists use models to help them visualize something that cannot be seen. Chemists may then make conclusions and develop theories, which they go on to test further. Chemists need to alternate between action in the macroscopic world (the "big" world) and thinking about the microscopic world (the "small" world) (**Figure 1**).

LEARNING **TIP**

Conceptualizing in Chemistry
Understanding concepts in chemistry often requires trying to visualize things that we cannot see. There are many ways you can help yourself to be better at this. Try using hands-on manipulative models, working with computer simulations, or drawing images on paper.

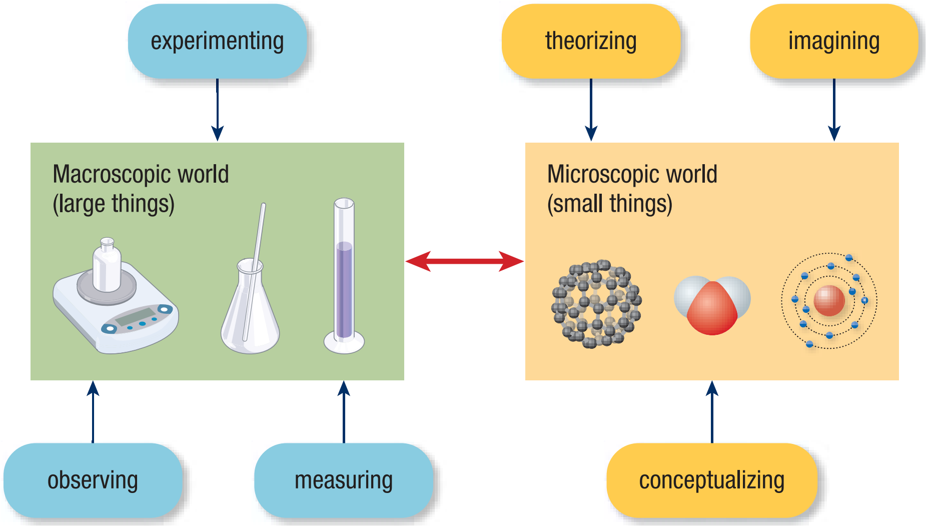


Figure 1 Chemistry involves moving back and forth between the macroscopic and the microscopic worlds.

The application of this knowledge leads to advances in technology. Science involves explaining observations about substances. Technology, however, often involves applying these explanations to produce something useful.

The relationship between science and technology is complex. Sometimes science leads to technological advances, but at other times technology leads to scientific discovery. As an example, scientists discovered the substance Teflon by accident. Since then, Teflon has been used in many applications, such as a coating on cookware. You will learn more about Teflon later in this unit.

Scientific discoveries and technological advances carry both risks and benefits. All scientists, engineers, and technicians must be aware of the risks and benefits related to their research. Long-term risks are often hard to foresee. For example, chlorofluorocarbons (CFCs) were once extensively manufactured and widely used. They were effective refrigerants, cleaning solvents, and propellants in aerosol cans. Consumers considered CFCs to be extremely beneficial substances. Then scientists discovered that CFCs deplete the stratospheric ozone layer. Safer alternative compounds have now replaced CFCs in most applications (**Figure 2**).

Figure 2 Technology has partly solved the CFC problem by finding much safer propellants for spray cans.

Chemistry at Work

Let's examine a specific example of chemistry at work. You are probably familiar with three forms of carbon: charcoal, graphite, and diamond. In the late twentieth century, scientists discovered traces of a new form of carbon, C_{60}. They found that pure C_{60} is an extremely stable substance that is able to withstand extremely high temperatures and that has low solubility in water. These experiments provided the chemists with **empirical knowledge**: knowledge that depends on observations. The chemists then pondered how they could explain these properties. This led them to propose models of the three-dimensional geometry of C_{60}.

empirical knowledge knowledge that comes from investigation and observation

The competition to isolate and identify C_{60} was intense as existing theories and new experiments worked hand in hand. Eventually, Harold Kroto, Robert Curl, and Richard Smalley were awarded the Nobel Prize in Chemistry in 1996 for their roles in its discovery. They proposed that its molecular shape resembles a soccer ball, with the carbon atoms forming the pentagons and hexagons on the surface of the sphere (**Figure 3**). Images of soccer balls and architectural domes helped chemists to imagine what these molecules might look like on a molecular scale. Knowledge based on ideas created to explain observations is known as **theoretical knowledge**.

theoretical knowledge knowledge that explains scientific observations

The discoverers named C_{60} "buckminsterfullerene," after architect Richard Buckminster Fuller, who was famous for his work on geodesic domes. This rather long name is often shortened to "buckyball." The family of carbon molecules with similar structures is known as "fullerenes." A fullerene is any type of sphere, ellipsoid, or tube made entirely of carbon.

WEB **LINK**

In 2010 a Canadian astronomer discovered buckyballs in space, using NASA's infrared telescope. To find out more about Jan Cami's discovery,

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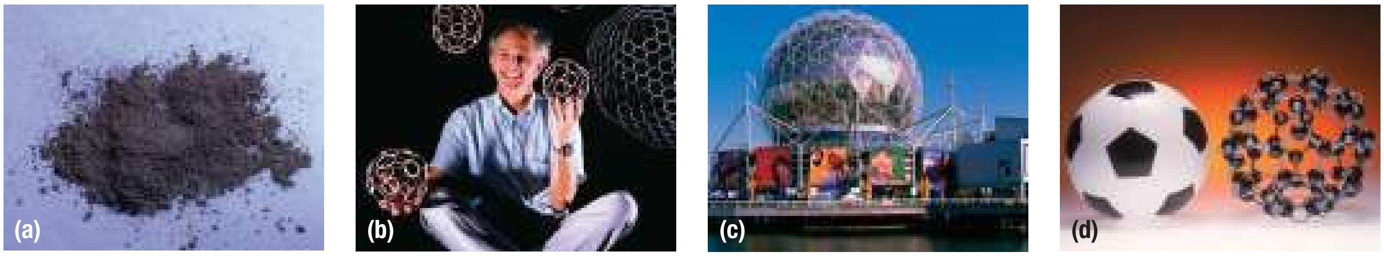

Figure 3 (a) Pure C_{60} is a greyish powder. (b) Professor Harold Kroto proposed these models of buckyballs. (c) Vancouver Science World boasts a geodesic dome designed by Buckminster Fuller. (d) Chemists now have experimental evidence for the structure of C_{60}.

The suggested structure of a buckyball is a theory. A **theory** is an explanation of evidence obtained by observation, experimentation, and reasoning. It represents our best understanding about why something happens. It allows us to predict future events and apply knowledge to new situations.

theory an explanation or model based on observation, experimentation, and reasoning

UNIT TASK **BOOKMARK**

Think of unlikely or serendipitous developments as you plan your green product for the Unit Task described on page 134.

As more experimental work is done and as technology improves, theories evolve or are discarded. Before the mid-nineteenth century, people thought that heat was a substance. This substance, called "caloric," appeared to flow from hot objects to cooler ones. Experiments eventually disproved this theory and it was discarded.

As theories evolve, the body of scientific knowledge grows. For example, researchers are looking for possible uses of the C_{60} molecule in materials science, electronics, and health sciences. Buckyballs may be effective medicine delivery "cages."

Of course, chemistry does not always happen in a series of logical steps. Creativity, luck, and serendipity (chance) are often important.

Everyday Chemistry

Whether you are aware of it or not, you practice chemistry in your everyday life. You might be colouring your hair, cooking, gardening, exercising, or creating art. You are continually experimenting to achieve various objectives. Are you looking for that perfect shade of purple for your hair? Perhaps you are trying to figure out the ideal mix of fertilizers to yield the best crop of tomatoes. Maybe you are experimenting to find the best foods to eat before a cross-country running meet. You observe and measure things as you proceed. You make connections between what you changed and what happened. Often, you theorize about why things happen the way they do. You are already a budding chemist!

IUPAC and the Scientific Community

WEB **LINK**

To find out more about the importance of the International Union of Pure and Applied Chemistry to the chemistry community,

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Chemistry is a human endeavour performed by scientists all around the world. These scientists are driven by human emotions and actions: ambition, curiosity, wonder, competition, hope, and perseverance. There is the need for cooperation, trust, and governance in the scientific community. The International Union of Pure and Applied Chemistry (IUPAC) was established in 1919 in order to help regulate standards and procedures in chemistry. Specifically, IUPAC's mission involves promoting the international aspects of chemistry and applying chemistry to the service of humanity. In so doing, IUPAC promotes the values, standards, and ethics of science.

1.1 Summary

- Chemistry is the study of matter and its interactions. It is also a process for acquiring knowledge. It is a human endeavour.
- Chemistry involves both experimenting in the macroscopic world and theorizing about the microscopic world.
- The interplay between empirical (experimentally determined) knowledge and theoretical knowledge is critical in the study of chemistry.
- The International Union of Pure and Applied Chemistry (IUPAC) sets rules and standards for chemists worldwide.

1.1 Questions

1. What isn't chemistry? List as many things as you can think of that do not involve matter and its interactions. A
2. (a) Distinguish between science and technology.
   (b) Give a specific example of a scientific achievement and the technology that developed as a result of it. K/U A
3. Sometimes it is hard to foresee the long-term effects of a chemical on the environment. K/U
   (a) Give an example of a chemical that has a negative environmental effect, and briefly describe that effect.
   (b) What is being done to reduce the negative effect?
4. Explain the difference between empirical and theoretical knowledge, with examples. K/U
5. There is a lot of interest in hydrogen-powered automobiles these days. Give an example of the science involved and an example of a technology involved. A
6. (a) What does the acronym IUPAC stand for?
   (b) Why was this organization set up?
   (c) Do you think such an organization is necessary? Defend your position. K/U A

Atomic Structure

One of the most intriguing theories in science is that of the atom. An **atom** is the smallest possible particle of an element.

atom the smallest particle of an element

The Development of Atomic Theory

The concept of the atom has been around for over 2000 years. Over that time our models of the atom have changed significantly as scientists used them to help explain how matter behaves.

Around 400 BCE, early Greek philosophers argued that matter could be subdivided into tiny indivisible particles called *atomos* (meaning "cannot be cut" or "indivisible"). Democritus went further: he stated that atoms have different sizes, are in constant motion, and are separated by empty spaces. This idea was rejected by Aristotle, the famous philosopher, who proposed that all matter is composed of four essential substances: earth, air, water, and fire. For almost 2000 years Aristotle's theory was widely accepted as being correct.

The Alchemists

The period of time from the first century CE to the seventeenth century is sometimes known as the period of alchemy. During this period, alchemists explored the nature of matter intensely. Alchemy was practised in many cultures around the world including Egypt, India, Persia (modern Iran), Europe, China, and Japan. Alchemists searched for such elusive prizes as the "elixir of life," with its hope of immortality, and the philosopher's stone, which was thought to transform common metals into gold. This period saw great technological advances. Their construction of lab glassware and equipment, development of alloys, handling procedures for dangerous chemicals, and various chemical processes are developments that are still in use today (**Figure 1**).

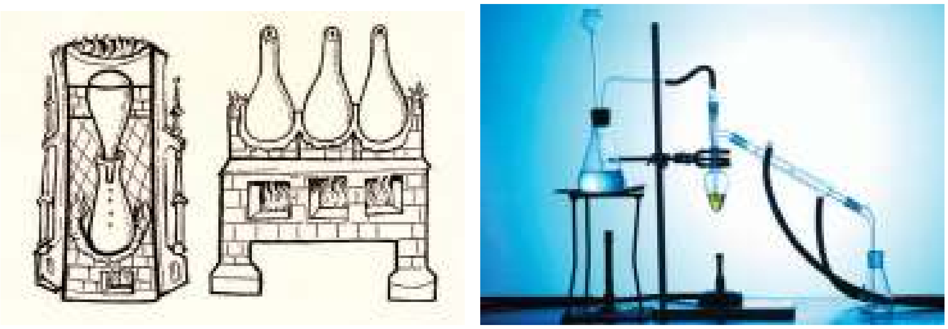

Figure 1 The alchemists developed lab procedures and equipment that were the precursors of modern apparatus.

Dalton's Atomic Theory

The scientific revolution in the seventeenth and eighteenth centuries increased the focus on scientific methods and the importance of evidence. Aristotle's ideas were eventually dismissed and scientists revisited Democritus' theory of the indivisible atom. In 1808, English scientist and schoolteacher John Dalton proposed a theory to explain the observations of matter (**Figure 2**). He suggested that atoms are solid spheres like billiard balls. More specifically, Dalton proposed that

- all matter is made up of tiny, indivisible particles called atoms
- all atoms of an element are identical
- atoms of different elements are different
- atoms are rearranged to form new substances in chemical reactions, but they are never created or destroyed

Dalton's theory was an important step forward. It was not the end of the story, however. Future observations would indicate that Dalton's theory could not explain some new evidence.

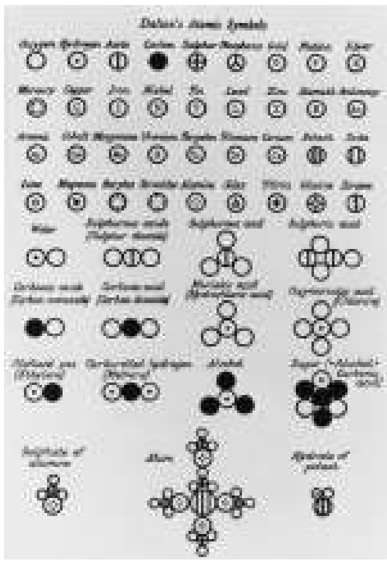

Figure 2 Dalton's list of the elements from 1808

Thomson's Discovery of the Electron

One of the problems was that Dalton's theory did not address how things acquire electrical charge. In 1897, J.J. Thomson used an apparatus called a cathode ray tube to discover a "new" type of particle: the electron (**Figure 3**). Through experiments, Thomson proposed that an **electron** is a negatively charged, extremely small part of an atom. A new model of the atom now included electrons spaced evenly in a positively charged sphere: the "plum pudding model." The fruit represents the electrons (**Figure 4**).

electron a negatively charged particle in an atom or ion

Figure 3 J.J.Thomson used a device known as a cathode ray tube to perform his experiments.

Figure 4 Thomson's plum pudding model of the atom could be called the blueberry muffin model.

Rutherford's Discovery of the Nucleus

In 1909 Ernest Rutherford performed a series of important experiments to test Thomson's model (**Figure 5**). Rutherford aimed tiny positive alpha particles at a thin sheet of gold foil. (An alpha particle, like a helium nucleus, has 2 protons and 2 neutrons.) He then measured how much the gold foil deflected the particles by surrounding the foil with a fluorescent screen. The screen glowed where the particles hit it. Rutherford made a prediction based on Thomson's model: if the electrons were equally distributed throughout the atom, the positively charged alpha particles should pass straight through the atoms of the foil. They would not be deflected (**Figure 6(a)**). After extensive trials, Rutherford observed that most of the alpha particles did indeed pass through the gold foil unaffected. However, a few particles were deflected at large angles (**Figure 6(b)**). Rutherford was amazed by his observations. He famously declared "it was almost as incredible as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you." From these observations, Rutherford reasoned that each atom contained a small, dense, positively charged central nucleus (**Figure 6(c)**).

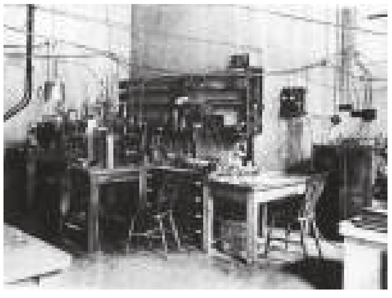

Figure 5 Ernest Rutherford, a brilliant physicist from New Zealand, spent nine years (1898–1907) working at McGill University in Montréal.

WEB LINK

To find out more about the work of Ernest Rutherford and his time at McGill University,

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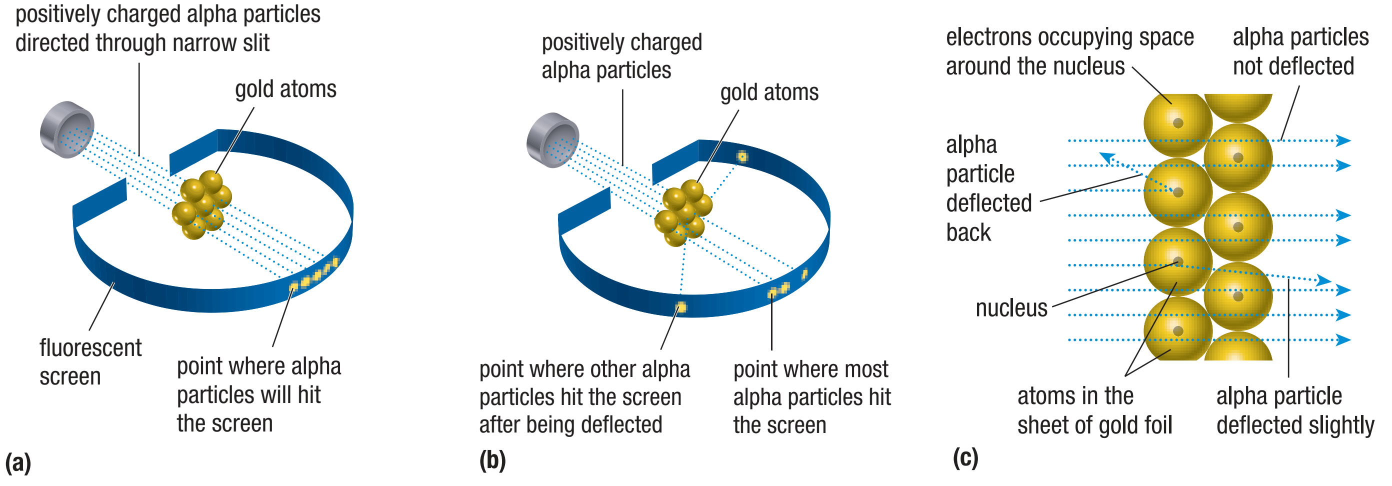

Figure 6 (a) Rutherford's prediction: if Thomson's model were correct, the alpha particles would pass through the gold atoms unaffected (b) Rutherford's observation: a small portion of the alpha particles were deflected by the gold atoms (c) Rutherford's explanation: each atom has a dense, positively charged nucleus that deflects alpha particles

In 1920, Rutherford proposed that the nucleus is made up of positively charged particles, each one called a **proton**. He also suggested that the nucleus is surrounded by mostly empty space occupied by electrons. This model of the atom is known as the nuclear model (**Figure 7**).

Just how far from the nucleus are the surrounding electrons? Imagine that you are standing at one end of a soccer field. If the nucleus of the atom is the size of a dime, the electrons are, on average, as far away as the other end of the field!

Rutherford also predicted the presence of a third subatomic particle with the same mass as the proton but with no charge.

proton a positively charged particle in the atom's nucleus

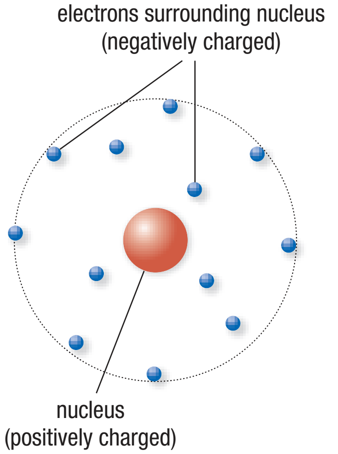

Figure 7 Rutherford's nuclear model of the atom (not to scale)

Chadwick's Discovery of Neutrons

Other scientists soon modified Rutherford's nuclear model. In 1932, James Chadwick's experiments confirmed that nuclei contain neutral particles as well as protons. These neutral particles became known as neutrons. A **neutron** is a particle in the nucleus of an atom or ion. A neutron has no charge: it is neither positive nor negative.

neutron a neutral particle in an atom's nucleus

Bohr's Proposal of Energy Levels

Further work was performed by Niels Bohr, a Danish scientist. He experimented with applying electricity and thermal energy to hydrogen gas. He observed that the hydrogen atoms emitted light when they were "excited" by the additional energy. Bohr directed the light through a prism with a screen behind it (**Figure 8**). He observed lines of only certain colours of light. He realized that hydrogen has a unique atomic spectrum: a pattern of coloured lines that is not produced by any other element. Astronomers have observed this same emission from stars and interstellar gas clouds.

WEB LINK

To view the atomic spectra for various elements,

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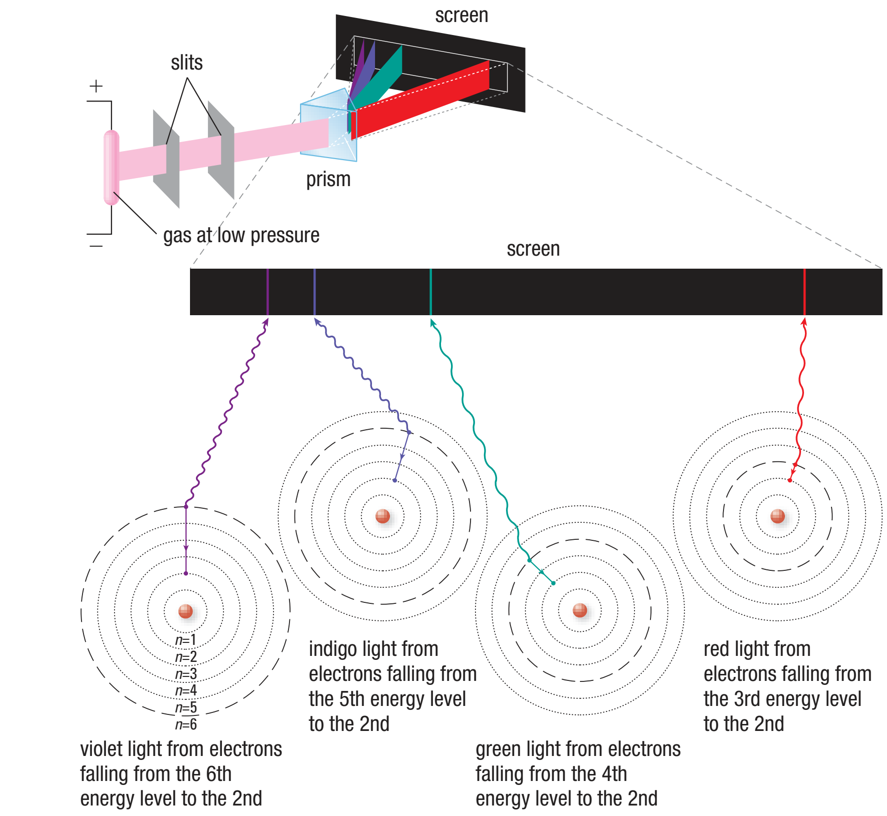

Figure 8 Bohr's work showed that electrons release bursts of light energy. The wavelength (colour) of the light depends on where the electrons are jumping from and to. Hydrogen atoms, shown here, can emit four colours of light. Each element emits a unique spectrum.

Bohr's observation led to a radical new proposal: that electrons orbit the nucleus of the atom in definite energy levels. Bohr had again revised the model of the atom. His revision is called the planetary model of the atom (**Figure 9**).

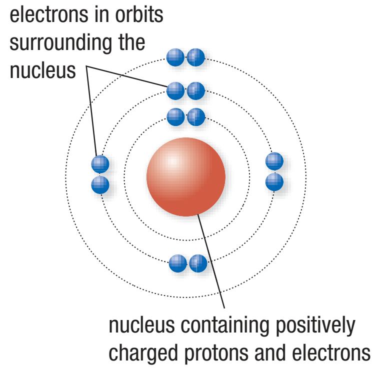

Figure 9 Bohr's planetary model of the atom

energy level a theoretical sphere around an atom where electrons exist; electron orbit

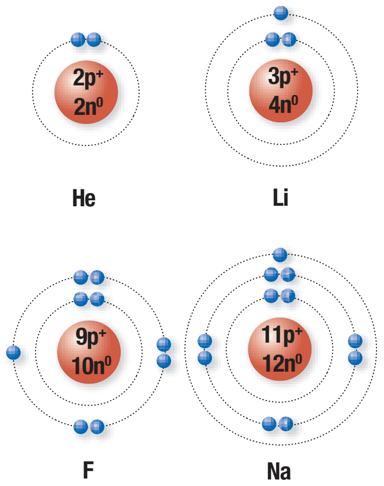

Figure 10 An atom of helium, He, has 2 valence electrons. Lithium, Li, and sodium, Na, both have 1 valence electron. Fluorine, F, has 7 valence electrons.

valence shell the outermost energy level or orbit of an atom or ion

valence electron an electron in the outermost energy level or orbit

In the planetary model, electrons only exist in certain allowed orbits. Each orbit has a specific quantity of energy associated with it. As a result, electron orbits are sometimes called **energy levels**. Electrons can jump from and to different energy levels within an atom. As they jump from higher energy levels back down to lower energy levels, they emit light energy. The colours of light emitted by the excited hydrogen atoms correspond to the changes of energy that the electrons experience as they move between energy levels (Figure 8). Since he observed only certain colours, Bohr suggested that an atom has only certain specific energy levels.

In Bohr's model of the atom, each energy level can hold a certain number of electrons. For the first 18 elements in the periodic table, the maximum number of electrons in the first, second, and third orbits is 2, 8, and 18 respectively.

Bohr–Rutherford Diagrams

We combine ideas from the work of Dalton, Thomson, Rutherford, Chadwick, and Bohr to draw a model of the atom called a Bohr–Rutherford diagram. We can use Bohr–Rutherford diagrams to show the number of each type of subatomic particle in a specific atom and to represent the arrangement of electrons around the nucleus. These diagrams are especially useful for the first 20 elements. In a Bohr–Rutherford diagram, concentric circles represent the different energy levels, the electrons are drawn in the appropriate locations, and the numbers of protons and neutrons are indicated in the nucleus.

Most chemical reactions involve only the electrons in the outermost energy level. The outermost energy level is also known as the **valence shell**. The electrons in the valence shell are **valence electrons**. **Figure 10** shows the Bohr–Rutherford diagrams for atoms of helium, He-4; lithium, Li-7; fluorine, F-19; and sodium, Na-23. (The number after each chemical symbol represents the total number of protons and neutrons in the nucleus.)

A Summary of Subatomic Particles

Table 1 summarizes the subatomic particles and their characteristics.

Table 1 Subatomic Particles

Subatomic particle	Symbol	Location in the atom	Charge	Approximate mass (kg)
electron	e^-	in energy levels outside the nucleus	–1	9.11×10^{-31}
proton	p^+	in the nucleus	+1	1.67×10^{-27}
neutron	n^0	in the nucleus	0	1.67×10^{-27}

The Value of Different Models

Looking back, we can see that the model of the atom has undergone many significant changes. The theory of the atom has evolved with continued experimentation and improvements in technology. The Bohr–Rutherford model shows electrons orbiting the nucleus like the planets orbiting the Sun. While the model is a simplified one, it is very useful. This model allows us to make connections between the atomic structures of different elements and their chemical and physical properties.

Many scientists have contributed to our understanding of the atom. In science there are often competing theories, controversies, and setbacks as scientists strive to explain their observations. Later scientists, including Albert Einstein and Erwin Schrödinger, worked to further extend our understanding of the atom. They conducted experiments and proposed explanations regarding the arrangement of electrons in different elements. You will explore more theories about electron arrangement as you move on to further chemistry courses.

WEB **LINK**

The Large Hadron Collider (LHC) is the world's largest particle accelerator. To learn about Canada's participation in LHC experiments to investigate the nature of matter and energy,

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Atomic Number and Mass Number

By the 1930s, scientists had found that the Bohr–Rutherford model of the atom was very useful. It helped to explain physical and chemical properties of elements (**Figure 11**). Scientists knew of the three subatomic particles and had a model of how these particles were arranged in the atom. They summarized this information in the periodic table. We can use the periodic table to find the number of protons and electrons in a neutral atom. As well, given the atomic mass of an element, we can determine the number of neutrons in an atom of that substance.

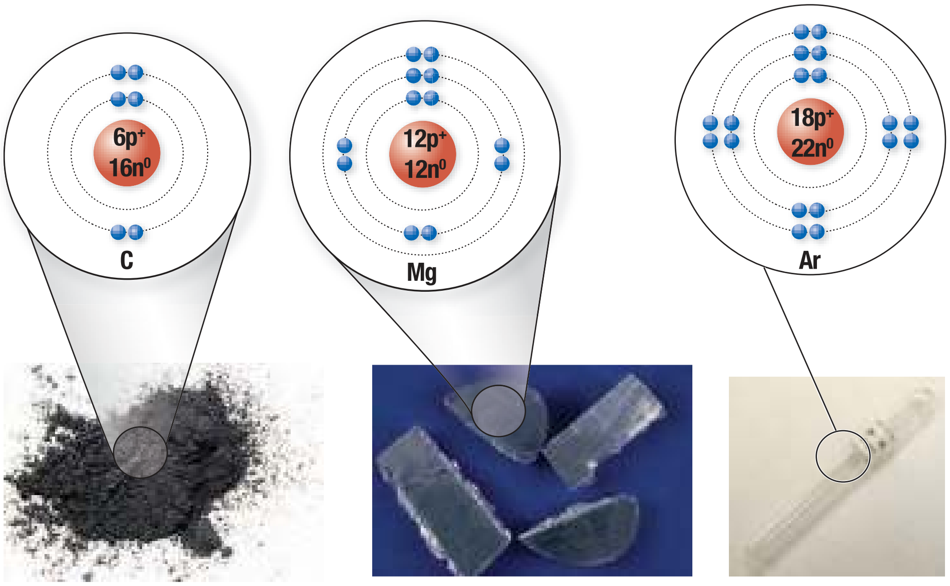

Figure 11 How are the atoms of these three substances different and how are they similar?

Atomic Number

The number of protons in one atom of a specific element is known as the **atomic number**. The symbol "*Z*" is often used to represent the atomic number. All atoms of the same element have the same atomic number. For example, a carbon atom always has 6 protons, a magnesium atom always has 12 protons, and an atom of argon always has 18 protons. The atomic number is usually written in the top left corner of an element's cell on the periodic table (**Figure 12**). The number of protons in a neutral atom is equal to the number of electrons. We can infer this because, in a neutral substance, the positive charges due to the protons must balance the negative charges from the electrons. For example, a carbon atom with 6 protons must have 6 electrons in order for it to be neutral.

atomic number (*Z*) the unique number of protons in one atom of an element

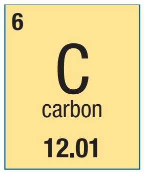

Figure 12 Each element appears in a different cell in the periodic table. The element's atomic number can be found in the top left-hand corner of its cell.

Mass Number

The **mass number** for an atom is the sum of all the particles in the nucleus. The number of protons plus the number of neutrons is equal to the mass number, *A*. We can therefore determine the number of neutrons as follows:

number of neutrons = mass number − atomic number

Figure 13 shows a popular convention used to represent an individual atom of an element. Consider an atom represented by ${}^{19}_{9}F$. From this information, we can determine that this atom has 9 protons (atomic number, $Z = 9$), 9 electrons (because atoms are neutral), and 10 neutrons ($19 - 9 = 10$). Similarly, an atom of ${}^{40}_{19}K$ would have 19 protons, 19 electrons, and 21 neutrons. The symbol for an atom is often written as C-12 or carbon-12, where 12 is the mass number. You should be able to use all three methods of notation.

mass number (*A*) the sum of the protons and neutrons in the nucleus of an atom

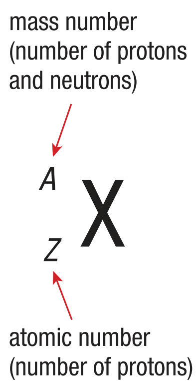

Figure 13 This notation lets us determine the number of protons, electrons, and neutrons in an atom.

The Mass of an Atom

Atoms are extremely small. In the head of a pin there are approximately 8.0×10^{19} iron atoms: 80 000 000 000 000 000 000 atoms. It would be very cumbersome to represent the mass of atoms using the kilogram. To work with the very small mass quantities of atoms and molecules, scientists use a smaller unit agreed upon by IUPAC and SI convention. The accepted unit to measure the mass of atoms is the atomic mass unit, u. The **atomic mass unit** is defined as $\frac{1}{12}$ the mass of a carbon-12 atom. Scientists have determined experimentally that $1\ \text{u} = 1.660\ 540\ 2 \times 10^{-27}$ kg. This unit is a relative one based on the mass of carbon-12. The masses of all other atoms are consequently measured relative to the mass of carbon-12. For example, F-19 has a mass that is $\frac{19}{12}$ times that of carbon-12. An atom of F-19 therefore has a mass of 19 u. Each H-1 atom has $\frac{1}{12}$ the mass of a C-12 atom. Its mass is therefore 1 u.

atomic mass unit a very small unit of mass defined as $\frac{1}{12}$ the mass of a carbon-12 atom; unit symbol u

LEARNING **TIP**

Le Système international d'unités
SI units, or le Système international d'unités, is a set of rules used around the world for scientific communication. The rules specify which units are used for which quantities. Appendix B1 lists many of the SI units commonly used in chemistry.

1.2 Summary

- Atomic theory has evolved through several different models over the past 2000 years.
- The Bohr–Rutherford model shows electrons orbiting the nucleus. This model helps us to predict physical and chemical properties for an element.
- The atomic number of an element is the number of protons in the nucleus. The mass number is the sum of the number of protons and neutrons.
- An atom can be represented in three different ways: $^{12}_{6}\text{C}$, C-12, and carbon-12.
- An atomic mass unit (u) is defined as $\frac{1}{12}$ the mass of one carbon-12 atom.

1.2 Questions

1. Draw Bohr–Rutherford diagrams for the following atoms: K/U C
   (a) Ca (b) C (c) Si
2. How many valence electrons are in each of the atoms that you drew in Question 1? K/U
3. Construct a timeline for the development of atomic theory. Include the major milestones and the names of key scientists and brief descriptions or diagrams of their experiments. (You could make an electronic timeline.) K/U C
4. Reread the proposals of Dalton's atomic theory. Which were still valid after the work of Rutherford and Bohr? Which were no longer valid? Why? K/U T/I
5. Copy and complete **Table 2** in your notebook. K/U T/I
6. Draw a flow chart showing the historical progression of different models of the atom. Include the name of and a sketch for each model. K/U C
7. Why should we learn atomic theory? If it is probably incorrect in some way and will undoubtedly change, why do we spend time learning about it? T/I A
8. Suppose your teacher asked you to build a scale model of an atom on your school's athletic field. Draw a detailed sketch of your model with approximate dimensions. T/I C
9. Most stars are composed primarily of hydrogen and helium. What might a star's atomic spectrum look like? A
10. Imagine that you helped to discover a new element: $A = 302$; $Z = 119$. How many protons, electrons, and neutrons are in each atom of this element? T/I
11. You may have heard about quarks and quark theory. Quarks are elementary particles that make up larger particles, called hadrons. Protons and neutrons are two examples of hadrons. Research quark theory and how quarks might fit into our atomic model. T/I

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Table 2 Atoms of Selected Elements

Name of element	Symbol	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
	U		235			
magnesium					13	
		53	131			
	C				8	
technetium			99			

Ions and the Octet Rule

The noble gases are extremely stable (inert) elements. They do not usually form compounds. One of the noble gases, helium, is less dense than air. Because of its stability and low density, helium is a good choice for filling blimps (**Figure 1**). Helium is not the only gas to be used for lighter-than-air flight, though. An alternative is hydrogen, a very light but highly flammable gas. In 1937 a hydrogen-filled German airship, *Hindenburg*, caught fire and killed 36 people. Helium was already being used for airships in the United States. However, the U.S. government would not share the technology for isolating helium with Germany. Since that tragic event, most airships have been filled with helium.

Argon and krypton are used extensively inside incandescent light bulbs. These noble gases make the filaments in the bulbs last much longer. As well, neon is used in colourful lighting displays and in vacuum and television tubes (**Figure 2**). It is ideal for this use primarily because of its chemical inertness.

Figure 1 Modern blimps get their "lift" from inert helium.

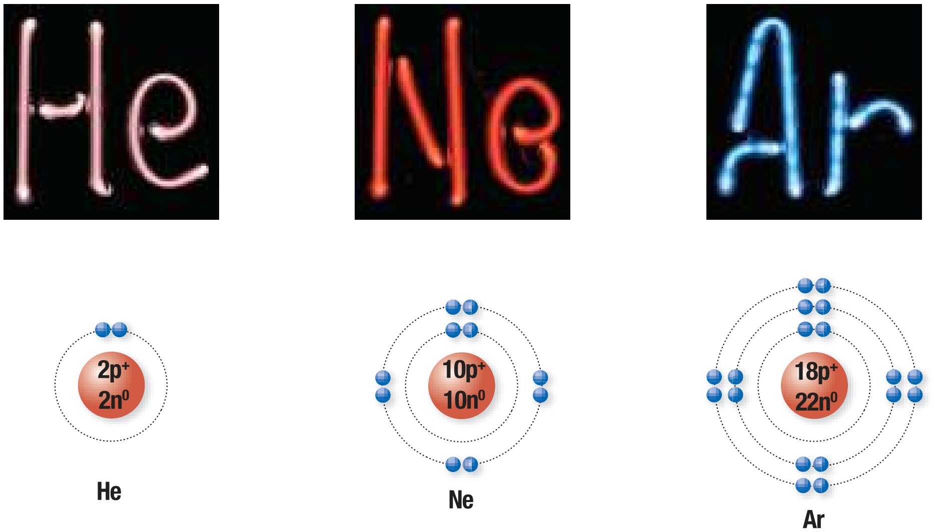


Figure 2 The Bohr–Rutherford diagrams of helium, neon, and argon show their full valence shells.

The Octet Rule

Elements with a full valence shell have a special stability. In the first 18 elements, a full valence shell (except the first shell) contains 8 electrons. Atoms of noble gases each have 8 electrons in their valence shells. This stable arrangement is known as a **full or stable octet**. The exception is helium, which is stable with 2 electrons in its valence shell.

Atoms of other elements do not have a full valence shell. Atoms, when they combine with other atoms, tend to attain this electron arrangement. This generalization is known as the **octet rule**. There are three possible ways in which an atom can achieve this stable arrangement: it can share, lose, or gain electrons. When an atom loses or gains electrons, it forms an **ion:** an entity with a positive or negative charge. (Recall that "entity" is a term that chemists use for an atom, an ion, or a molecule.) Whether an atom loses or gains electrons depends on the number of valence electrons it has.

full or stable octet an electron arrangement where the valence shell is filled with 8 valence electrons (2 for hydrogen and helium)

octet rule a generalization stating that when atoms combine, they tend to achieve 8 valence electrons

ion a charged entity formed when an atom gains or loses one or more electrons

The Formation of Ions

As you know, atoms sometimes lose or gain electrons to form ions. Some elements are more likely to lose electrons and become positive ions; others are more likely to gain electrons and become negative ions.

Positive Ions: Cations

The metals are located to the left of the zigzag staircase line on the periodic table. In the first few groups (columns) on the left side of the periodic table, the metals typically have just a few electrons in their valence shell. In general, metal atoms tend to lose valence electrons in order to achieve a stable electron arrangement.

cation a positively charged ion formed by the removal of one or more electrons from the valence shell of a neutral atom

Consider sodium in Group 1, the alkali metals. Each sodium atom has one valence electron in its outer orbit. In order to achieve a more stable arrangement, a sodium atom tends to lose that one electron (**Figure 3**). When it loses its electron it becomes a positively charged ion. A positively charged ion is known as a **cation**.

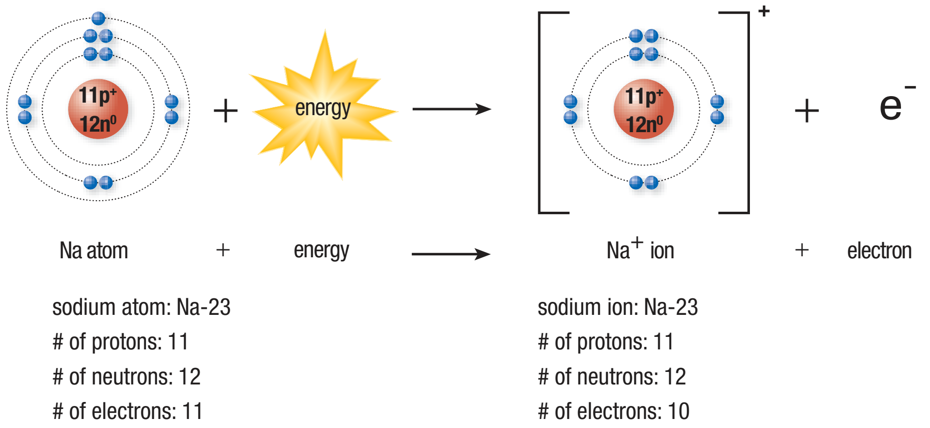


Figure 3 A sodium atom loses its one valence electron to become a sodium cation, Na^+.

valence the charge of an ion; the combining capacity of an atom determined by the number of electrons that it will lose, add, or share when it reacts with other atoms

A sodium ion has a positive charge of +1 because it has 11 positive charges (protons) and 10 negative charges (electrons). The charge of an ion is often called the **valence**. Hence, sodium has a valence of +1.

As another example, an aluminum atom has 3 valence electrons. It will therefore lose 3 electrons to become an Al^{3+} ion. Note that when we represent ions as Bohr–Rutherford diagrams, we enclose them with square brackets and write the charge as a superscript outside the brackets.

Naming cations is very simple. They have the same name as their element. Na^+ ions are simply called sodium ions and Al^{3+} ions are aluminum ions.

Negative Ions: Anions

anion a negatively charged ion formed by the addition of one or more electrons to a neutral atom

The elements on the right side of the periodic table are mostly non-metals. They generally have almost-complete valence shells. Non-metallic atoms tend to gain electrons in order to fill their valence shells. In doing so, they become negatively charged ions. A negatively charged ion is called an **anion**. Consider chlorine, in the halogen family (**Figure 4**). A chlorine atom gains one valence electron to fill its outer shell, forming a Cl^- ion. This ion has a negative charge of –1 because it has 17 positive charges (protons) and 18 negative charges (electrons). Chlorine has a valence of –1.

LEARNING **TIP**

Memory Aid—Cations and Anions
There are several ways to help you remember that cations are positive ions and anions are negative ions. The "t" in cation is like a plus (positive) sign. The first three letters in the word anion might stand for "are negative ions." Finally, the spelling of anion is close to that of "onion," which may have a negative effect on one's breath!

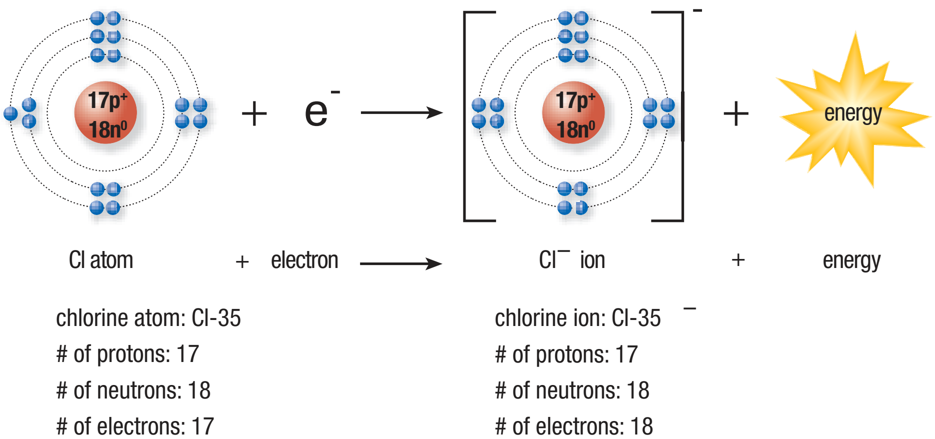


Figure 4 An atom of chlorine gains an electron to become a chloride anion, Cl^-.

Look at the position of sulfur on the periodic table. Each sulfur atom has 6 valence electrons and will therefore gain 2 electrons. In the process it gains a charge of –2 and becomes an S^{2-} ion.

Anions are named a little differently than cations. Non-metal ions are named by replacing the end of the element's name with the suffix *–ide*. For example, the Cl^{-} ion would be called a chloride ion and the S^{2-} ion would be a sulfide ion.

In general, metals lose electrons to become cations with the electron arrangement of the nearest noble gas with a smaller atomic number. Non-metals gain electrons to become anions with the electron arrangement of the nearest noble gas with a larger atomic number. **Figure 5** shows the most common valences for some of the elements on the periodic table that form ions.

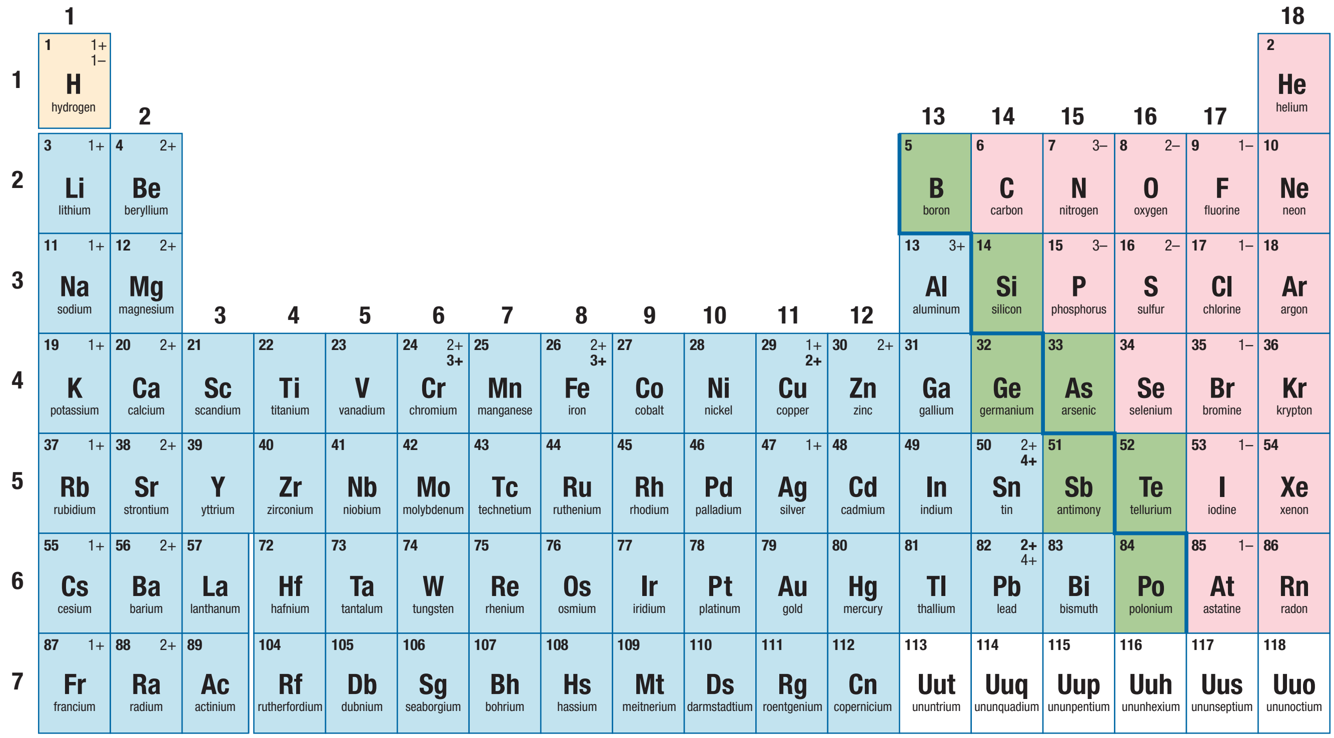


Figure 5 This periodic table shows the common valences for various elements on the periodic table. Look in the top right corner of each cell. Where an element has more than one valence, the most common valence is indicated in bold type. Do you notice any patterns?

Elements with Multiple Ionic Charges

In Figure 5, above, you may notice that some elements can form more than one possible ion. These substances are said to be multivalent. A **multivalent** element is one that can form two or more different stable ions. In fact, most of the transition metals—those in the middle part of the periodic table—can form more than one type of ion. For example, chemists have observed that copper can form both Cu^{+} and Cu^{2+} ions. Many of the multivalent elements are transition metals, so these elements are often called multivalent metals.

multivalent the property of having more than one possible valence

Chemists need a naming system to distinguish between the different ions formed by the same element. A traditional naming system used the Latin name of the element and a suffix of either *–ous* to represent the lower valence or *–ic* for the higher valence. Note that this naming system is useful only for multivalent elements with two different possible valences. This classical system is still widely used, but is gradually being replaced by a system approved by IUPAC. In the IUPAC system, a Roman numeral in the ion's name indicates the charge of the ion. **Table 1** shows some examples of multivalent metals and their classical and IUPAC names.

Table 1 Examples of Multivalent Metals

Metal	Ions	Classical names	IUPAC names
copper, Cu	Cu^{+} Cu^{2+}	cuprous cupric	copper(I) copper(II)
iron, Fe	Fe^{2+} Fe^{3+}	ferrous ferric	iron(II) iron(III)
tin, Sn	Sn^{2+} Sn^{4+}	stannous stannic	tin(II) tin(IV)
lead, Pb	Pb^{2+} Pb^{4+}	plumbous plumbic	lead(II) lead(IV)
manganese, Mn	Mn^{2+} Mn^{3+} Mn^{4+} Mn^{6+} Mn^{7+}	n/a	manganese(II) manganese(III) manganese(IV) manganese(VI) manganese(VII)

Mini Investigation

Using Flame Tests

Skills: Predicting, Performing, Observing, Analyzing, Communicating

SKILLS HANDBOOK A1, B4

Chemists use flame tests and a reference key to identify unknown metallic ions in solution. In this investigation, your task is to observe and create a key for the flame test results when you test different metallic compounds (**Figure 6**). You will then use this key to identify an unknown metallic compound.

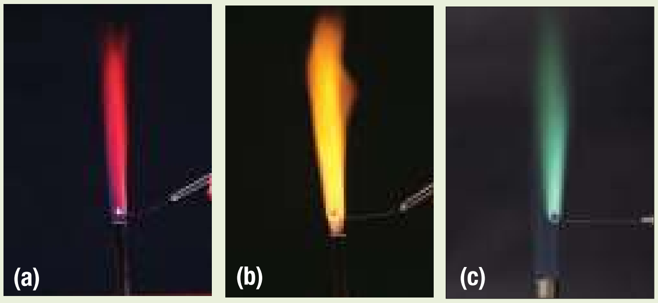


Figure 6 Flame test for (a) lithium, (b) sodium, and (c) a mystery metal compound

Equipment and Materials: chemical safety goggles; lab apron; Bunsen burner clamped to a retort stand; spark lighter; nichrome test wire; small labelled beaker containing 50 mL of dilute hydrochloric acid; small labelled beakers containing dilute 5 mL samples of the following solutions: sodium chloride, calcium chloride, strontium chloride, lithium chloride, potassium chloride, and copper(II) chloride (0.1 mol/L)

This investigation involves the use of open flames. Tie back long hair and secure loose clothing and jewellery.

The hydrochloric acid and sodium hydroxide solution used in this activity are irritants. Wash any spills on skin or clothing immediately with plenty of cool water. Report any spills to your teacher.

1. Put on your chemical safety goggles and lab apron.
2. Light the Bunsen burner with your spark lighter and adjust it so that the flame is blue.
3. Clean the nichrome test wire by dipping it in the acid and then placing it in the flame for 10 s.
4. Dip the clean nichrome wire into one of the solutions. Hold the end of the wire in the flame until a uniquely coloured flame appears. Record your observations.
5. Repeat Steps 3 and 4 to clean the wire and perform the test on the remaining substances.

A. Which, if any, of your samples gave identical results? T/I

B. Suggest an explanation for this observation. T/I

C. Suggest the identity of the mystery substance being tested in Figure 6(c). T/I

D. Suggest a real-life application in which flame tests would be used to identify ions. A

GO TO NELSON SCIENCE

Polyatomic Ions

You may have noticed advertisements for phosphate-free dishwashing detergents, or nitrate-free or nitrite-free meat products (**Figure 7**). Phosphate ions (PO_4^{3-}), nitrate ions (NO_3^-), and nitrite ions (NO_2^-) are ions that consist of more than one atom. An ion that has more than one atom is called a **polyatomic ion**.

polyatomic ion an ion, made up of more than one atom, that acts as a single entity

People sometimes think that substances containing polyatomic ions are dangerous. The compounds have long, complicated-sounding names. If you look at the list of ingredients on food packaging, the polyatomic compounds appear to be "chemicals"—by which we often mean "synthetic compounds" rather than natural substances (**Figure 8**). This is a misconception. Polyatomic ions occur in nature and are essential to our health. For example, calcium phosphate is a major constituent of bones and teeth, and hydrogen carbonate ions help to regulate blood pH.

Figure 7 Some polyatomic ions have negative effects on the environment or human health.

Figure 8 Many compounds containing polyatomic ions are essential for good health.

Polyatomic ions are useful in many industries. For example, sodium nitrate is used in fertilizers. Meat processors add nitrates to their products. These nitrates convert to nitrites due to natural processes within the meat. A manufacturer can claim that no nitrites are added, yet nitrites may still be present.

Polyatomic ions have some negative effects. If phosphates make their way into lakes or ponds, they act as fertilizer and cause blooms of algae. As well, nitrites form cancer-causing agents when they react with substances in the digestive system. For these reasons, the use of phosphates in cleaning products is heavily regulated and we are advised not to consume processed meats too often.

Most polyatomic ions are composed of oxygen, nitrogen, phosphorus, sulfur, chlorine, and carbon. Also, most polyatomic ions are anions (**Table 2**). Each ion gains one or more extra electrons so that each of its atoms reaches a stable arrangement of electrons. The phosphate ion, PO_4^{3-}, has gained 3 extra electrons to reach a stable configuration. A polyatomic ion behaves just like an ion made of only one atom.

Table 2 IUPAC Names and Formulas for Some Common Polyatomic Ions

Name	Formula
acetate	$C_2H_3O_2^-$
bromate	BrO_3^-
carbonate	CO_3^{2-}
hydrogen carbonate	HCO_3^-
hypochlorite	ClO^-
chlorite	ClO_2^-
chlorate	ClO_3^-
perchlorate	ClO_4^-
chromate	CrO_4^{2-}
dichromate	$Cr_2O_7^{2-}$
cyanide	CN^-
hydroxide	OH^-
iodate	IO_3^-
permanganate	MnO_4^-
nitrite	NO_2^-
nitrate	NO_3^-
phosphate	PO_4^{3-}
hydrogen phosphite	HPO_3^{2-}
hydrogen phosphate	HPO_4^{2-}
dihydrogen phosphite	$H_2PO_3^-$
dihydrogen phosphate	$H_2PO_4^-$
sulfite	SO_3^{2-}
sulfate	SO_4^{2-}
hydrogen sulfide	HS^-
hydrogen sulfite	HSO_3^-
hydrogen sulfate	HSO_4^-
thiosulfate	$S_2O_3^{2-}$
ammonium	NH_4^+

Ions in the Human Body

About 99 % of your body is made up of only 6 elements. In order of abundance by mass, these elements are oxygen, carbon, hydrogen, nitrogen, calcium, and phosphorus. You also contain much smaller quantities of sulfur, chlorine, sodium, magnesium, iodine, and iron. Many of these elements exist as ions dissolved in water. As you are probably aware, these ions play key roles in our bodies (**Table 3**).

Table 3 Some Important Ions in the Human Body

Ion	Role	Source
Na^+	important for body fluid control	salt cheese preservatives
K^+	important for body fluid control and cell functions	bananas milk potatoes
Ca^{2+}	a key component of bone and teeth	milk cheese spinach
Fe^{3+}	important in muscle function; an essential part of hemoglobin in blood	kidney beans asparagus pine nuts
Mg^{2+}	crucial for muscle and nerve functions	green plants nuts grains
Cl^-	important for body fluid control	salt
I^-	helps regulate the body's metabolic rate	fish dairy products iodized salt

The human body requires a delicate balance of many ions for good health. Sodium ions, for example, play a critical role in the normal functioning of the human body. Sodium ions help with nerve impulse transmission, muscle contraction, and water balance. If sodium ion levels fall too low, death can result. Our Western diet usually contains plenty of sodium. In fact, many of us consume more than the recommended daily intake of sodium. Excessive sodium intake can, in some people, lead to hypertension and an increased risk of heart disease. There is a great deal of concern in many countries about the high level of sodium in today's foods—especially processed foods and meals from fast-food restaurants.

CAREER **LINK**

Nutritionists advise people on the best foods to eat—or avoid—for optimal health. If this is a career you would like to know more about,

GO TO NELSON SCIENCE

Research This

Tattoo Ink—Decorative Body Art or Toxic Mixture?

Skills: Researching, Analyzing, Defining the Issue, Communicating, Defending a Decision   SKILLS HANDBOOK A5.1, A5.2

Tattoos have long been a form of personal expression. They are created by depositing a pigment into the skin (**Figure 9**). Colours are determined by the chemical composition of the pigment. For example, black is made from carbon and iron(II) oxide, blue is made from copper phthalocyanine, and violet gets its colour from a mixture of aluminum salts.

Figure 9 Are tattoos safe?

1. Research the health concerns related to getting a tattoo. Find resources that consider the concerns in an objective fashion and that cite supporting evidence. Investigate the following questions:
   - Is it safe to get a tattoo?
   - What regulations currently apply to tattoo parlours in Ontario?
   - Should the tattoo industry be more strictly controlled?
   - Canadian Blood Services will not let people donate blood if they have recently received tattoos. Why?

A. Analyze the evidence on the issue of whether or not it is safe to get a tattoo. Formulate arguments for and against the safety of tattoos. T/I A

B. Debate the safety of getting a tattoo. C A

GO TO NELSON SCIENCE

1.3 Summary

- The octet rule states that atoms tend to gain or lose electrons to achieve a stable octet: the electron arrangement of the nearest noble gas.
- Ions are atoms that have gained or lost electrons.
- Metals tend to lose electrons to form cations. Cations are positively charged ions that have fewer electrons than protons.
- Non-metals tend to gain electrons to form anions. Anions are negatively charged ions that have more electrons than protons.
- Multivalent elements can form two or more different ions.
- Polyatomic ions are ions composed of more than one atom.
- Many ions are necessary for good health.

1.3 Questions

1. Draw a Bohr–Rutherford diagram for each of the following ions. Represent the ions correctly with square brackets and charge. T/I C
   (a) K^+ (b) F^- (c) N^{3-} (d) Mg^{2+}
2. For each of the ions in Question 1, name the noble gas with the same electron arrangement. T/I
3. State the octet rule. K/U
4. Write the IUPAC name for each of the following ions: K/U
   (a) O^{2-} (d) SO_4^{2-}
   (b) Cu^+ (e) OH^-
   (c) Sn^{4+} (f) NH_4^+
5. Manganese atoms can form a wide variety of ions, including Mn^{2+}, Mn^{3+} and Mn^{4+}. Propose how we can communicate which ion is present in a compound. K/U
6. Write the formula and charge for each of the following polyatomic ions. K/U
   (a) nitrate (c) acetate
   (b) carbonate (d) permanganate
7. Calcium carbonate, $CaCO_3$, is a critical component of the shells of various aquatic species. Identify the cation and anion in calcium carbonate. T/I
8. Which 4 atoms or ions from the following list have the same electron arrangement? O^{2-} S^{2-} Na^+ Al^{3+} Ne F K/U
9. Give two experimental techniques that might help to identify the presence of metal ions. T/I
10. Anemia is a health condition caused by an iron deficiency. Research common symptoms of anemia and suggest foods that could be added to a diet to provide more iron. T/I A

GO TO NELSON SCIENCE

Isotopes, Radioisotopes, and Atomic Mass

Every year, doctors worldwide request millions of diagnostic medical tests. These tests involve searching for clues within the human body that will help to prevent or treat various diseases (**Figure 1**). Radioactive substances are widely used in these diagnostic procedures. For example, iodine-131 can be used to diagnose and treat thyroid cancer, and technetium-99 is used in a variety of diagnostic tests.

Historically, Canada has played an important role in supplying these critical materials to the world. Chalk River Laboratories, a Canadian nuclear research facility northwest of Ottawa, has provided such materials since its start-up in 1962 (**Figure 2**). However, these nuclear facilities are aging and recently experienced safety problems. These factors have led to facility closures. Canada's leadership role in this industry is now in jeopardy.

What are these radioactive substances? Why are they so important for medical procedures? Why is it so difficult to supply them safely? What are their risks and benefits?

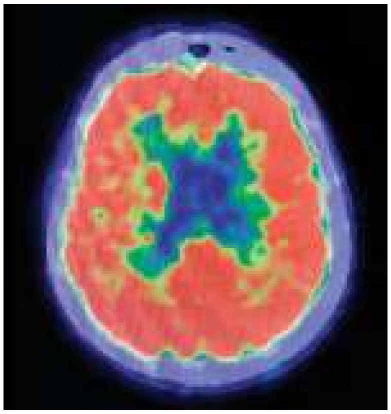


Figure 1 Medical imaging technologies can help doctors diagnose problems.

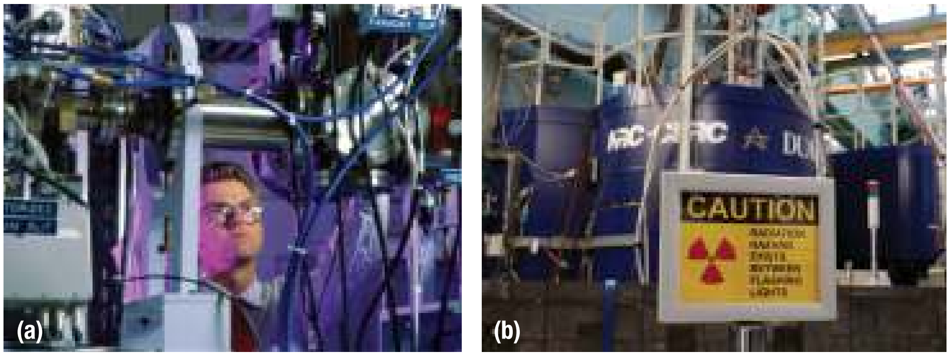


Figure 2 (a) Chalk River Laboratories, Ontario (b) The aging nuclear reactor at Chalk River

CAREER **LINK**

Radiologists are healthcare professionals who use various imaging technologies to diagnose and treat disease. These technologies make use of radioisotopes. To learn more about radiology as a career,

GO TO NELSON SCIENCE

Isotopes and Isotopic Abundance

Isotopes are atoms of the same element that have different numbers of neutrons. The name "isotope" is based on Greek words meaning "at the same place." It was first proposed by Margaret Todd, a Scottish physician. Todd was having a conversation with a colleague of Ernest Rutherford's, Frederick Soddy, about how there seemed to be variations within elements. Soddy adopted the term and went on to prove the existence of isotopes for some elements. He was awarded the Nobel Prize in Chemistry in 1921.

isotope a form of an element in which the atoms have the same number of protons as all other forms of that element, but a different number of neutrons

Scientists use mass number—the total number of protons and neutrons in the nucleus—to distinguish between different isotopes for a given element. For example, carbon-12 and carbon-14 are two different isotopes of carbon. Both isotopes have 6 protons and 6 electrons, but they differ in their number of neutrons: carbon-12 has 6 neutrons and carbon-14 has 8 neutrons. You read in Section 1.2 that the atomic mass unit, u, is defined as $\frac{1}{12}$ the mass of a carbon-12 atom. Scientists had to be this specific because carbon has three naturally existing isotopes, the third one being carbon-13.

Different elements have different numbers of isotopes. These specific isotopes exist in different relative abundances. For example, natural magnesium is a mixture of three isotopes: magnesium-24, magnesium-25, and magnesium-26 (**Figure 3**). On average, a sample of natural magnesium consists of 78.7 % magnesium-24, 10.1 % magnesium-25, and 11.2 % magnesium-26 (**Figure 4**).

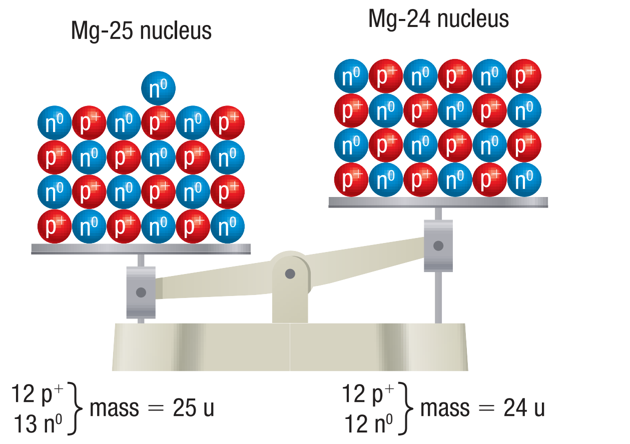
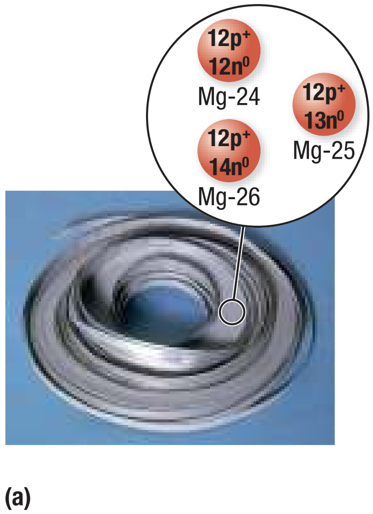


Figure 3 Isotopes have different numbers of neutrons and therefore different masses. This model shows two of the isotopes of magnesium.



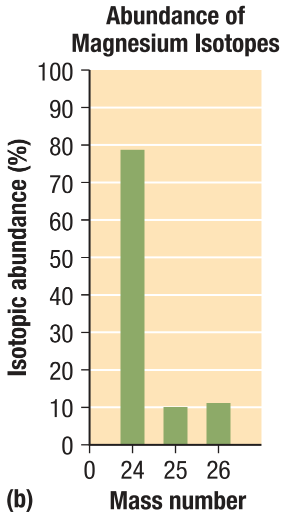


Figure 4 Naturally occurring magnesium contains three isotopes: Mg-24, Mg-25, and Mg-26.

isotopic abundance the percentage of a given isotope in a sample of an element

mass spectrometer a measuring instrument used to determine the mass and abundance of isotopes

The percentage of an isotope in a sample of an element is known as the **isotopic abundance**. As an analogy, consider a class made up of 30 students: 16 girls and 14 boys. The abundance of girls is $\frac{16}{30} = 53\ \%$ and the abundance of boys is $\frac{14}{30} = 47\ \%$.

Scientists use a special instrument known as a **mass spectrometer** to identify isotopes and their respective abundances (**Figure 5**). Mass spectrometers are very useful in chemistry. They can perform a variety of analyses, such as identifying proteins, detecting atmospheric pollutants, and drug testing in sports. Mass spectrometers have even been sent into space.

A simple type of mass spectrometer is composed of three main sections: the ion source, the analyzer, and the detector (**Figure 6**). The sample is injected into the spectrometer and vaporized by heat. The sample is then ionized and accelerated by an electric field. The fast-moving ions next pass through a magnetic field, where they are deflected. The magnetic field deflects smaller isotopes more than larger isotopes. A detector plate senses the relative abundance of each isotope and a computer determines the mass and abundance of each isotope.

Figure 5 A lab technician works at a mass spectrometer.

CAREER LINK

Radioisotope lab technicians work in a variety of settings, operating many types of instruments and equipment. If this career appeals to you,

GO TO NELSON SCIENCE

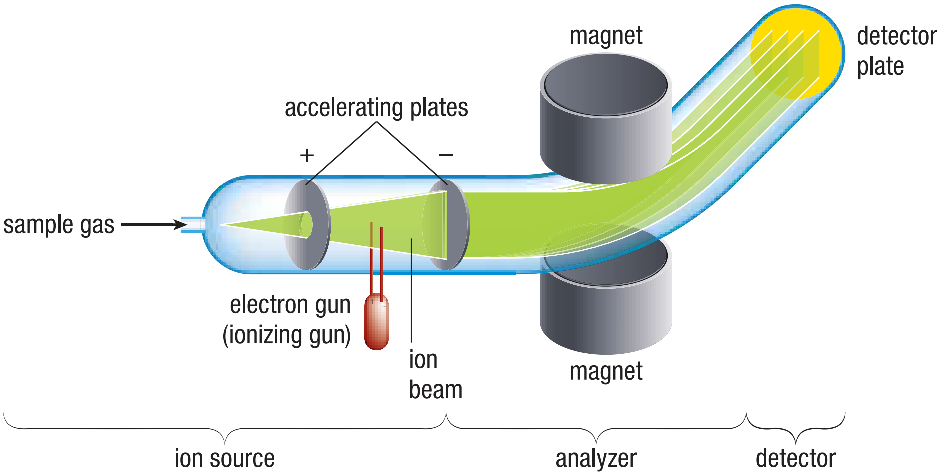


Figure 6 A mass spectrometer

Radiation and Radioisotopes

As we have seen, some elements have more than one isotope. Some isotopes are stable; others break apart easily. The difference in stability is due to the composition of their nuclei. For example, oxygen has three naturally occurring isotopes, O-16, O-17, and O-18, all of which are stable. Scientists have been able to make 10 additional isotopes of oxygen, all of which are unstable and emit nuclear radiation.

Nuclear Radiation

Radioactive substances spontaneously break apart. This disintegration is known as **radioactive decay**. As some isotopes decay, they emit **nuclear radiation** in the form of tiny particles or energy. The three most common types of nuclear radiation are alpha particles, beta particles, and gamma rays. Each type of nuclear radiation has a different ability to penetrate through matter (**Figure 7**). An **alpha particle, α,** has the same structure as the nucleus of a He-4 atom—2 protons and 2 neutrons—and a charge of +2. Alpha particles are blocked by paper. A **beta particle, β,** is a negatively charged electron that can pass through paper but not through aluminum. A **gamma ray, γ,** is a form of high-energy electromagnetic radiation. Gamma rays have no mass and travel at the speed of light. They can penetrate most substances but are blocked by lead.

radioactive decay the spontaneous disintegration of unstable isotopes

nuclear radiation energy or very small particles emitted from the nucleus of a radioisotope as it decays

alpha particle a product of nuclear decay emitted by certain radioisotopes; a positively charged particle with the same structure as the nucleus of a helium atom

beta particle a product of nuclear decay emitted by certain radioisotopes; a negatively charged particle identical to an electron

gamma ray a form of high-energy electromagnetic radiation emitted by certain radioisotopes

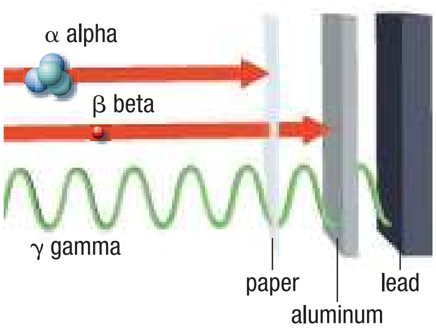


Figure 7 The three types of radiation have different penetrating abilities.

Radioisotopes

Isotopes that decay to produce nuclear radiation are known as **radioisotopes**. All of the isotopes of uranium are radioisotopes; there are no stable uranium isotopes. This is one of the reasons that governments place so many safeguards and restrictions on any proposed uranium mine. All radioisotopes are **radioactive**, which means that they all emit radiation as they decay.

Nuclear chemistry is the branch of chemistry that deals with the nucleus and radioactivity. The isotopes that are used for medical treatments or diagnostic procedures are all radioisotopes. Iodine-131, for example, is a radioisotope of iodine that is used to treat an overactive thyroid gland. Other uses for radioisotopes include smoke detectors (americium-241), food irradiation (cobalt-60), and archaeological dating (carbon-14 and potassium-40).

radioisotope an isotope that spontaneously decays to produce two or more smaller nuclei and radiation

radioactive having the potential to emit nuclear radiation upon decay

CAREER LINK

Nuclear chemistry is a fascinating and rewarding field of study. If you would like to find out more about the education and work of a nuclear chemist,

GO TO NELSON SCIENCE

Determining the Atomic Mass of Elements

Look at the periodic table and find the atomic mass for chlorine. It is usually written below the element's name and, as you will notice, it is not a whole number. The atomic mass for chlorine is 35.45 u. What does this number represent? Why is it not a whole number? How does one go about determining this value? How is this value useful? As we will find out, this value is extremely important as we progress into working with quantities in chemistry.

atomic mass the weighted average of the masses of all the naturally occurring isotopes of an element

The **atomic mass** of an element is determined by calculating the weighted average of the masses of all isotopes of that element. Usually when an average of a set of values is calculated, you simply add up the values and divide by the number of values. This is valid if the values are of all the same significance. A weighted average, however, considers not only the values but also the relative abundance of each value.

Tutorial 1 Calculating Atomic Mass

We can illustrate the difference between non-weighted average and weighted average with a simple example. Imagine that we want to find the average height of a group of eight people: Antonia (who is 150 cm tall), Shirley (170 cm), and Gurpreet (160 cm), plus Ahmed and Ameer (identical twins: 200 cm), and Rachel, Emily, and Elizabeth (identical triplets: 180 cm).

$$\begin{aligned}\text{average height} &= \frac{150\text{ cm} + 170\text{ cm} + 160\text{ cm} + 2(200\text{ cm}) + 3(180\text{ cm})}{8}\\ &= \frac{1420\text{ cm}}{8}\\ &= 178\text{ cm} \quad \text{(rounded to 3 significant figures)}\end{aligned}$$

Notice that the following calculation would be incorrect.

$$\begin{aligned}\text{average height} &= \frac{150\text{ cm} + 170\text{ cm} + 160\text{ cm} + 200\text{ cm} + 180\text{ cm}}{5}\\ &= \frac{860\text{ cm}}{5}\\ &= 172\text{ cm}\end{aligned}$$

Why is this incorrect? Because it does not take into account the fact that there are two people who are 200 cm tall and three people who are 180 cm tall. When you calculate a weighted average, you must account for a greater abundance of people of certain heights.

Consider the actual abundance (as a percentage) of people with different heights. $\frac{1}{8}$ (or 12.5 %) of the people are 150, 160, and 170 cm tall; $\frac{2}{8}$ (or 25 %) of the people are 200 cm tall; and $\frac{3}{8}$ (37.5 %) of the people are 180 cm tall. The average could also be determined by the following calculation:

$$\begin{aligned}\text{average height} &= 12.5\ \%\ (150) + 12.5\ \%\ (160) + 12.5\ \%\ (170) + 25\ \%\ (200)\\ &\quad + 37.5\ \%\ (180)\\ &= 178\text{ cm}\end{aligned}$$

FINDING THE ATOMIC MASS GIVEN ISOTOPIC ABUNDANCE

We can use the process above to calculate the weighted average of all the isotopes for a certain element. We must consider both the mass and the abundance of each of the different isotopes of an element. This calculation gives the atomic mass.

When solving problems related to atomic mass, use the GRASS problem-solving format and the general equation for atomic mass:

$$\begin{aligned}\text{atomic mass} = {} &\%\text{ abundance of isotope 1 (mass of isotope 1)} + \\ &\%\text{ abundance of isotope 2 (mass of isotope 2)} + \ldots\end{aligned}$$

Investigation 1.4.1

The Nuts and Bolts of Atomic Mass **(p. 42)**
You will use a model to represent the isotopes of an element, and calculate its atomic mass.

LEARNING TIP

Working with Weighted Averages
Whether you realize it or not, you work with weighted averages all the time. Your final mark in a course is usually determined by calculating a weighted average of your semester work (70 %) and your final summative assessment and/or culminating activity (30 %).

Sample Problem 1: Calculating Atomic Mass For 3 Isotopes

Calculate the atomic mass of magnesium. Magnesium-24, magnesium-25, and magnesium-26 have isotopic abundances of 78.7 %, 10.1 %, and 11.2 % respectively.

Given: atomic mass and abundance of the 3 isotopes of magnesium

Required: atomic mass of magnesium

Analysis:

$$\begin{aligned}\text{atomic mass} &= \text{\% abundance of isotope 1 (mass of isotope 1)} + \\ &\quad \text{\% abundance of isotope 2 (mass of isotope 2)} + \ldots\end{aligned}$$

Solution:

$$\begin{aligned}\text{atomic mass} &= 78.7\ \%\ (24\ \text{u}) + 10.1\ \%\ (25\ \text{u}) + 11.2\ \%\ (26\ \text{u}) \\ &= 24.3\ \text{u}\end{aligned}$$

Think about the answer obtained. Confirm that the answer makes sense. Round it to the appropriate number of digits.

Statement: The atomic mass of magnesium is 24.3 u.

Sample Problem 2: Calculating Atomic Mass For 5 Isotopes

Germanium has the following isotopic composition. Calculate the atomic mass of germanium.

20.5 %	Ge-70
27.4 %	Ge-72
7.8 %	Ge-73
36.5 %	Ge-74
7.8 %	Ge-76

Given: atomic mass and abundance of the 5 isotopes of germanium

Required: atomic mass of germanium

Analysis:

$$\begin{aligned}\text{atomic mass} &= \text{\% abundance of isotope 1 (mass of isotope 1)} + \\ &\quad \text{\% abundance of isotope 2 (mass of isotope 2)} + \ldots\end{aligned}$$

Solution:

$$\begin{aligned}\text{atomic mass} &= 20.5\ \%\ (70\ \text{u}) + 27.4\ \%\ (72\ \text{u}) + 7.8\ \%\ (73\ \text{u}) + 36.5\ \%\ (74\ \text{u}) + \\ &\qquad 7.8\ \%\ (76\ \text{u}) \\ &= 14.35 + 19.73 + 5.69 + 27.01 + 5.93\ \text{[extra digits carried]} \\ &= 72.71\ \text{u}\end{aligned}$$

Think about the answer obtained. Confirm that the answer makes sense. Round it to the appropriate number of digits.

Statement: The atomic mass of germanium is 73 u.

Practice

1. Calculate the atomic mass of each of the following elements, given these naturally occurring isotopes and abundances: K/U T/I
   (a) Neon: Ne-20 (90.5 %), Ne-21 (0.3 %), Ne-22 (9.2 %) [ans: 20 u]
   (b) Titanium: Ti-46 (7.9 %), Ti-47 (7.3 %), Ti-48 (73.9 %), Ti-49 (5.5 %), Ti-50 (5.4 %) [ans: 48 u]

Research This

Radon in the Home

Skills: Researching, Analyzing, Evaluating, Defending a Decision, Communicating SKILLS HANDBOOK A5.1

Radon is a colourless, odourless, radioactive noble gas. It is produced when naturally occurring uranium undergoes radioactive decay. Radon can collect in confined areas of the home and contribute to our daily dose of radiation (**Figure 8**). This radiation increases a person's risk for lung cancer. Home inspectors will test homes for radon gas.

1. Research the potential sources of radon in a home.
2. Learn more about the geology of your local area and whether the presence of uranium might mean high levels of radon.

A. Do you think there should be more public awareness about radon in homes? If so, suggest methods of promoting this. T/I A

B. In your opinion, should home inspectors test for radon levels in houses on the market? Support your position. T/I A

Figure 8 How radon enters a house

GO TO NELSON SCIENCE

CAREER **LINK**

Home inspectors visit homes and check them for safety and soundness. To find out more about this career,

GO TO NELSON SCIENCE

1.4 Summary

- Elements can exist as a variety of isotopes.
- Isotopes are atoms of the same element that have different numbers of neutrons.
- Radioisotopes are unstable isotopes. They spontaneously undergo radioactive decay. During decay they produce other elements, radiation, and energy.
- There are three types of radiation produced during radioactive decay: alpha radiation, beta radiation, and gamma radiation.
- The atomic mass for an element is the weighted average of the masses of all its naturally occurring isotopes. The atomic mass is generally listed in the periodic table.

1.4 Questions

1. There are three isotopes of hydrogen: hydrogen-1, hydrogen-2 (known as deuterium), and hydrogen-3 (known as tritium). Create a table listing the number of protons and neutrons in an atom of each of hydrogen's isotopes. T/I C
2. The atomic mass of chlorine is 35.45 u. Is it possible for any single atom of chlorine to have a mass number of exactly 35.45? Explain. K/U
3. Silver exists in nature as two isotopes: Ag-107 and Ag-109. If the average atomic mass of silver is 107.9 u, which isotope is more abundant? Explain your answer K/U
4. Silicon naturally exists as three isotopes (**Table 1**). Determine the atomic mass of silicon. T/I

Table 1 Percentage Abundance of Three Silicon Isotopes

Isotope	% abundance
Si-28	92.23
Si-29	4.67
Si-30	3.10

5. Imagine an element, X, that has two naturally occurring isotopes. If you know the mass and the percentage abundance of one of the isotopes, how would you determine the percent abundance of the other isotope? Describe your problem-solving process. T/I C
6. The atomic mass of carbon is 12.0107 u. It exists naturally as three isotopes: C-12, C-13, and C-14. Based on your understanding of isotopes and atomic mass, determine which isotope would have the greatest abundance. Explain your choice. T/I C
7. Naturally occurring chlorine consists primarily of two isotopes: Cl-35 and Cl-37. Determine the number of protons, electrons, and neutrons for an atom of each isotope. T/I
8. Distinguish between an isotope and a radioisotope. K/U
9. Potassium naturally consists of 93.10 % K-39 and 6.90 % K-41. Calculate the atomic mass for potassium. T/I
10. Research three different careers in which people handle radioisotopes (**Figure 9**). List safety precautions they take to ensure minimal exposure to radiation. T/I A

Figure 9 Many medical careers involve radioisotopes.

11. In 1951, Canadian scientists developed a revolutionary new medical procedure known as the Cobalt bomb. Research this technology and compare its risks and benefits. T/I A
12. A Geiger counter is a device that is used to detect radiation from radioisotopes. Research how a Geiger counter works and list three specific circumstances in which a Geiger counter would be useful. T/I A
13. The town of Port Hope in Ontario is home to Cameco (previously named Eldorado), a major producer of uranium. In the 1940s, Eldorado supplied weapons-grade uranium to the United States. The radioactive waste from the mine was used as backfill for ravine properties in Port Hope. The town is still dealing with the radioactive contamination today. Investigate the present situation in Port Hope and what Cameco has done to help the situation. T/I A

GO TO NELSON SCIENCE

1.5 The Periodic Table and Periodic Law

You are watching a movie. The camera pans around and you catch sight of a periodic table on the wall. You instantly know that the action is in a science lab or classroom. The periodic table has become an icon of science. It is also very a useful tool.

Review of the Periodic Table

The periodic table contains all the known elements. Independently, these elements exhibit many characteristic properties. When the elements are arranged according to patterns of properties, they form the familiar periodic table. These patterns help us predict the physical properties of the elements and how elements will combine to form various types of compounds.

The periodic table is like an alphabet for chemistry. The English language contains 26 letters, and from these we can make up any word. We can compare letters to elements and words to compounds—although we would need a much longer song to remember all the elements. Fortunately, the periodic table provides an organized collection of all the elements and their properties.

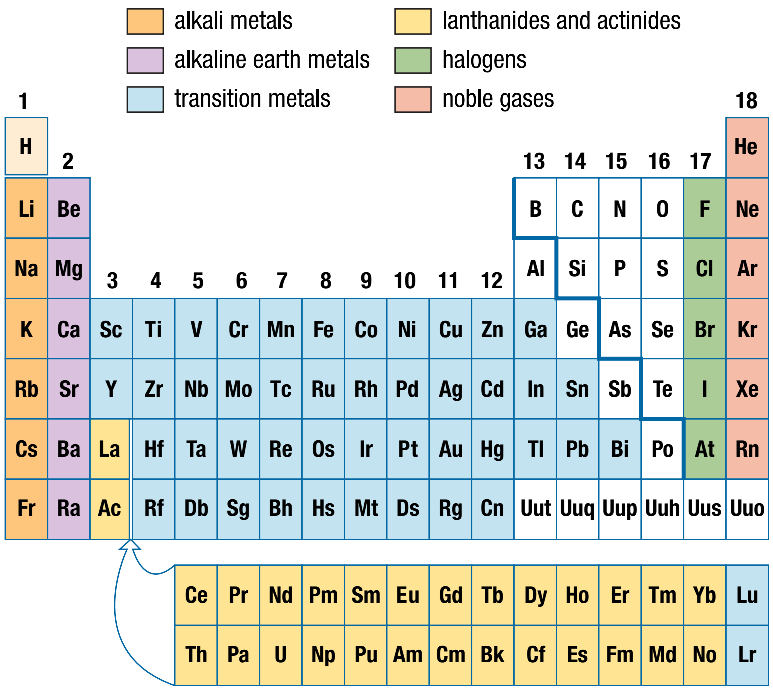
WEB LINK

Tom Lehrer's 1959 song, *The Elements*, lists the 102 elements known at that date. To read the lyrics or listen to the song,

GO TO NELSON SCIENCE

Organization of the Periodic Table

The modern periodic table has many important features and characteristics. Elements are organized in order of increasing atomic number, yet spaced so that elements with similar physical and chemical properties are aligned in columns (**Figure 1**).



hydrogen: a unique element with many physical properties of non-metals but chemical properties of metals

alkali metals: soft, silver-coloured elements; solids at room temperature; exhibit metallic properties; react violently with water to liberate hydrogen gas; react with halogens to form compounds such as sodium chloride, NaCl(s); stored under oil or in a vacuum to prevent reaction with air

alkaline earth metals: light, very reactive metals; solids at room temperature; exhibit metallic properties; react with oxygen to form oxides with the general chemical formula, MO(s); all except beryllium will react with hydrogen to form hydrides; react with water to release hydrogen

transition metals: exhibit a range of chemical and physical properties; strong, hard metals with high melting points; good conductors of electricity; variable reactivity; form multivalent ions; many react with oxygen to form oxides; some react with acids to release hydrogen gas

lanthanides: elements with atomic numbers 57 to 70

actinides: elements with atomic numbers 89 to 102

transuranic elements: synthetic (not naturally occurring) elements with atomic number 93 or greater (beyond uranium)

noble gases: gases at room temperature; low melting and boiling points; extremely unreactive, making them especially interesting to chemists; krypton, xenon, and radon reluctantly form compounds with fluorine; radon is radioactive

halogens: may be solids, liquids, or gases at room temperature; exhibit non-metallic properties—not lustrous and non-conductors of electricity; extremely reactive, especially fluorine; react readily with hydrogen and metals

representative elements: both metals and non-metals from Groups 1, 2, and 13 to 18; may be solids, liquids, or gases at room temperature; many form colourful compounds

Figure 1 Characteristic properties of elements in the periodic table

Metals and non-metals are separated on the table by a zigzag "staircase." The elements right next to this line, on either side of it, are **metalloids**. Most periodic tables indicate the state (gas, liquid, or solid) at which an element exists at 25 °C and a pressure of 100 kPa. (See the periodic table on the inside back cover of this textbook.) Some periodic tables also indicate whether an element occurs naturally or is synthetic (artificially made).

metalloid an element that has properties of both metals and non-metals

WEB **LINK**

There are many interactive periodic tables on the Internet providing a wealth of information. To try out some of them,

GO TO NELSON SCIENCE

Each vertical column on the periodic table is called a **group** (or family) of elements. Groups are numbered from 1 to 18 and some are given special names. On the far left side of the table, Group 1 is known as the alkali metal family and Group 2 is the alkaline earth metal family. On the right side of the table, elements in Group 17 are known as halogens and Group 18 elements are known as noble gases or inert gases.

group a column of elements in the periodic table; sometimes referred to as a family

Groups 1, 2, and 13 to 18 are sometimes called the representative elements.

The elements in the middle region of the periodic table—including silver, Ag, gold, Au, manganese, Mn, and iron, Fe—are the transition metals. All these elements have the characteristics of metals. However, they also display a wide range of properties. For example, platinum and gold are quite unreactive, whereas manganese and iron form a variety of compounds.

A horizontal row on the periodic table is referred to as a **period**. As you look down the periodic table, the number of elements in a period tends to increase. The word "period" signifies repeating cycles, just like your school periods that occur at regular intervals. As this term implies, we observe trends or patterns in the properties of the elements as we move across a period. Such trends are referred to as periodic trends. The **periodic law** states that we can observe periodic trends when the elements are arranged in order of increasing atomic number.

period a row in the periodic table

periodic law a rule, developed from many observations, stating that when the elements are arranged in order of increasing atomic number, their properties show a periodic recurrence and gradual change

The lanthanides (rare earth metals) and actinides are actually parts of Periods 6 and 7.

Investigation **1.5.1**

The Search for Patterns (p. 43)
By testing the reactivities of various metals, you will be able to start identifying trends within groups and periods of the periodic table.

Periodic Trends and Bohr–Rutherford Diagrams

You may recall from earlier investigations in chemistry that elements in the same family have similar physical and chemical properties. For example, the alkali metals lithium, Li, sodium, Na, and potassium, K, are all shiny, soft, silvery metals. These elements are highly reactive with water (**Figure 2**). In fact, the alkali metals, when they exist in their pure form, must be stored in mineral oil in order to prevent a reaction with the moisture in the air. Hydrogen is usually placed above the alkali metals, yet some versions of the periodic table place it above the noble gases or up in a location of its own. Hydrogen behaves more like a non-metal, but chemists suggest that, at extremely high pressures and low temperatures, hydrogen becomes a metal. The cores of large gas planets, such as Jupiter, may contain hydrogen metal.

Elements in Group 2, the alkaline earth metals, also exhibit characteristic metallic properties. These elements are not as soft or as reactive as the alkali metals.

Figure 2 The alkali metals (lithium, sodium, and potassium) are all reactive when mixed with water.

The Group 17 halogens, consisting of the elements chlorine, Cl, fluorine, F, iodine, I, and bromine, Br, also share similar properties. They all exist as diatomic molecules (for example, Cl_2), and they are quite reactive.

The elements in Group 18, the noble gases, also share common physical and chemical properties. Helium, He, argon, Ar, neon, Ne, and krypton, Kr, are all inert.

Chemists try to connect the observed properties of the elements with current atomic models in order to make further predictions (**Figure 3**).

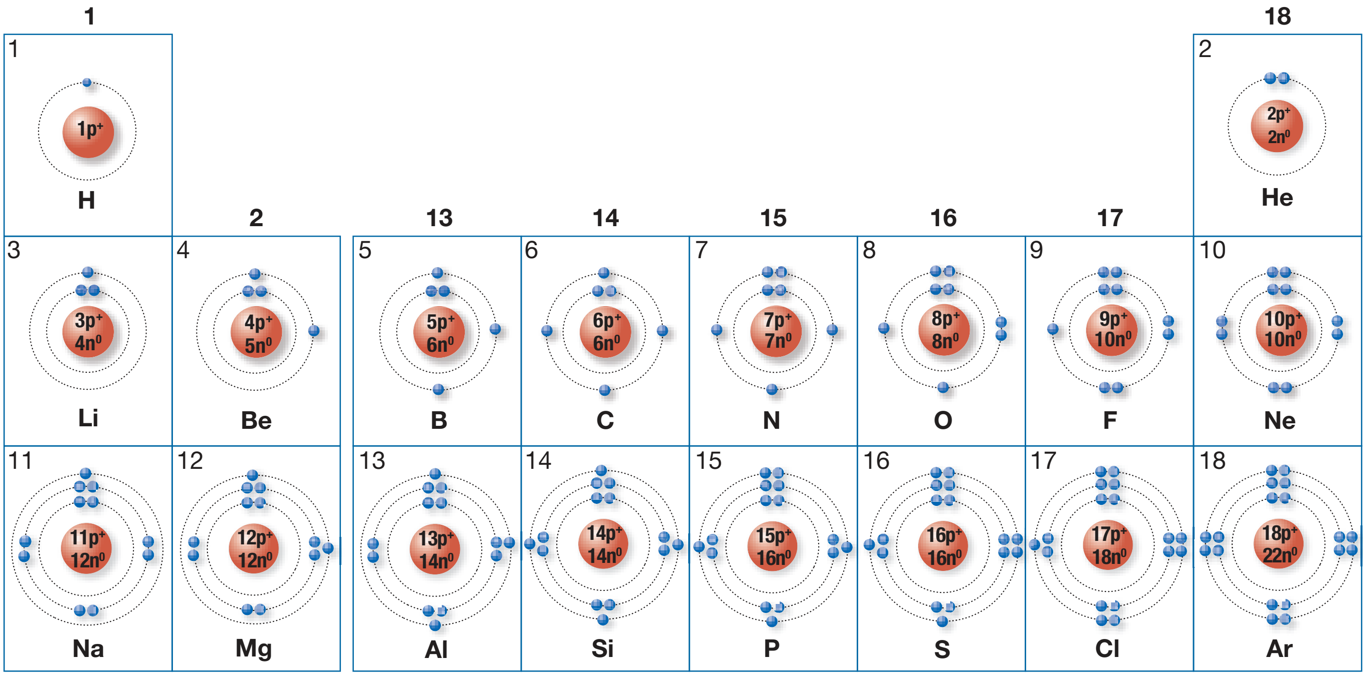


Figure 3 Bohr–Rutherford diagrams of the first 18 elements. Note that the transition metals are omitted to save space.

Notice that, in the Bohr–Rutherford diagrams for the alkali metals, each atom has a single valence electron. The alkaline earth metals each have 2 valence electrons. The halogens each have 7 valence electrons and the noble gases each have 8 valence electrons or a full orbit. We can connect the number of valence electrons with the chemical properties of elements in a family. As we move across a period from left to right, we observe trends in reactivity. Specifically, we see that elements in Group 1 are highly reactive, then reactivity decreases as we move to the right. Reactivity starts to increase again after Group 14, rising as we reach the halogens and then dropping rapidly at the noble gases. We can use these trends to predict elements' properties.

Lewis Symbols

Bohr–Rutherford diagrams are useful for representing the number of electrons in an atom. However, they are quite tedious to draw. G.N. Lewis, an American chemist, devised a shorthand system for representing elements. A **Lewis symbol** shows only the chemical symbol and dots representing the valence electrons around the chemical symbol. (This explains an alternative name: electron dot diagram.) For example, the Lewis symbol for lithium, Li•, indicates that the atom has just one valence electron.

The Lewis symbols for the representative elements are shown in **Figure 4**. Compare **Figures 3** and **4**.

Lewis symbol a representation of an element consisting of the chemical symbol and dots to represent the valence electrons; electron dot diagram

LEARNING **TIP**

Many Names; Same Diagram
Lewis symbols have several other names, including electron dot diagrams, Lewis dot structures, electron dot symbols, or almost any combination of these terms.

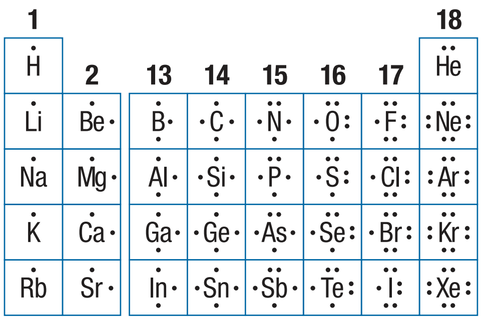


Figure 4 Lewis symbols for the representative elements

There are several ways to position the electrons in Lewis symbols. The convention used in this textbook is the same as that used for Bohr–Rutherford diagrams: we place a single electron at the top, then on the right, at the bottom, and on the left to a total of 4. Only then do we add the second electron, in the same order, to make electron pairs.

You will use Lewis symbols as we move forward into chemical bonding and molecular structures.

1.5 Summary

- The periodic table is an arrangement of the elements that exhibits periodic trends.
- Vertical columns are called groups. Horizontal rows are known as periods.
- All elements in the same group have the same number of valence electrons, resulting in similar chemical properties.
- Elements can be represented by Bohr–Rutherford diagrams or by Lewis symbols. Both show the number of valence electrons.

1.5 Questions

1. From the periodic table, give two examples of each of the following: T/I K/U
   (a) a metal
   (b) a metalloid
   (c) an element that is a gas at room temperature
   (d) a halogen
   (e) an alkali metal
   (f) a synthetic element
   (g) an element that is a liquid at room temperature
   (h) a member of the actinide series
2. Refer to the periodic table to identify each of the following elements: K/U
   (a) a member of the halogen family in the third period
   (b) a member of the noble gases in the fourth period
   (c) a metalloid in the third period
   (d) a member of the alkali metal family in the fifth period
3. Suggest how similar electron arrangements result in similar chemical properties. Refer to elements in the noble gas family in your explanation. T/I
4. Draw the Lewis symbol for the following elements: T/I C
   (a) Mg (d) F
   (b) Al (e) Ar
   (c) N
5. (a) Draw the Bohr–Rutherford diagrams and Lewis symbols for sodium, Na, beryllium, Be, and boron, B.
   (b) Explain how to draw Lewis symbols once you have drawn the Bohr–Rutherford diagrams. K/U C
6. How many valence electrons are there in an atom of each of the following elements? K/U
   (a) chlorine (e) neon
   (b) sulfur (f) carbon
   (c) magnesium (g) aluminum
   (d) sodium
7. Based on their position on the periodic table, predict the relative reactivities of the following elements: T/I
   (a) Cs versus Ba (c) Na versus Ar
   (b) C versus F (d) Mg versus Si
8. The calendar year repeats itself again and again, thus displaying periodic behavior. Name three other periodic phenomena. A
9. (a) Investigate the origins of the names of five elements.
   (b) Organize your findings into a format to share with your classmates. T/I C
10. The periodic table that you are used to is not the only arrangement of the elements. There are quite a few alternative proposals (**Figure 5**). Research a different periodic table of the elements arrangement and explain the differences and similarities to the one used in this textbook. A

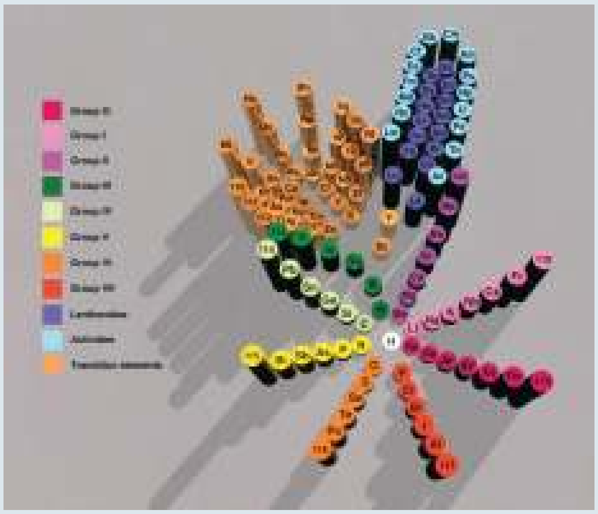


Figure 5 The radiating spokes of this periodic table represent the groups of elements.

GO TO NELSON SCIENCE

A Not-So-Elementary Task

SKILLS HANDBOOK A4

ABSTRACT

The periodic table of the elements represents one of science's greatest achievements. Many scientists have contributed, over hundreds of years, to this iconic arrangement. Dmitri Mendeleev is credited with inspiring the current version, but his work was based on the efforts of many other chemists (**Figure 1**). Today, chemists are still discovering new elements. New arrangements of the elements challenge the present structure of the periodic table.

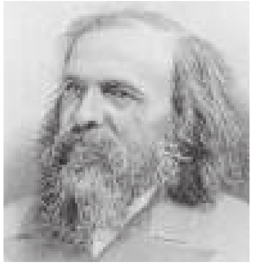


Figure 1

We have known about elements such as gold, silver, and copper for thousands of years. In the seventeenth century, the hunt for new elements began in earnest.

Early Efforts at Organizing the Elements

In 1649, the scientist Hennig Brand isolated phosphorus by distilling human urine. Since then, scientists have discovered and isolated more and more elements. As the list of elements grew, so did the need to classify and organize them.

In 1789 Antoine-Laurent de Lavoisier published one of the earliest chemistry books: *Elementary Treatise of Chemistry*. In this book, Lavoisier included a list of 33 different elements. While he mistakenly included "light" and "caloric," he did distinguish between metals and non-metals.

Johann Dobereiner (1780–1849), a German chemist, noticed that there were trends within groups of elements. He arranged "triads" of elements with similar chemical properties. For example, he recognized that lithium, sodium, and potassium have properties in common.

By the 1860s, scientists were starting to arrange the elements in order of increasing atomic mass. In England, John Newlands was the first to do this. Following Dobereiner's ideas of groups of elements, Newlands arranged the known elements into columns of seven. Newlands, a music lover, proposed a "law of octaves" (**Figure 2**). In a letter to the editor of *Chemical News*, Newlands wrote: "The eighth element, starting from a given one, is a kind of repetition of the first, like the eighth note of an octave in music." Newlands' law appeared to be useful only up to the 20th element, calcium. Because of this weakness, and because the similarity to the musical scale appeared whimsical, other scientists dismissed Newlands' ideas.

H	F	Cl	Co/Ni	Br	Pd	I	Pt/Ir
Li	Na	K	Cu	Rb	Ag	Cs	Tl
G	Mg	Ca	Zn	Sr	Cd	Ba/V	Pb
Bo	Al	Cr	Y	Ce/La	U	Ta	Th
C	Si	Ti	In	Zn	Sn	W	Hg
N	P	Mn	As	Di/Mo	Sb	Nb	Bi
O	S	Fe	Se	Ro/Ru	Te	Au	Os

Figure 2 Newlands' law of octaves

The Work of Dmitri Mendeleev

The famous Russian chemistry professor, Dmitri Mendeleev (1834–1907), gets the credit for developing the periodic table that led to our current version. Mendeleev, born in Tobolsk, Siberia, was the youngest of 17 children. Overcoming many hardships, Mendeleev became a professor of chemistry in St. Petersburg. He published his table of elements in 1869 (**Figure 3**).

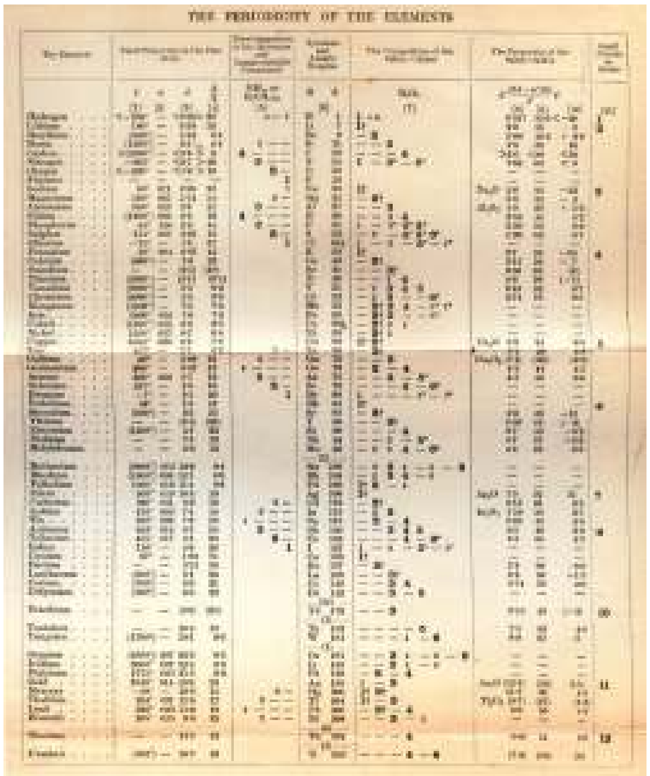

THE PERIODICITY OF THE ELEMENTS

Figure 3 The first English publication of Mendeleev's proposal was in his book, *Principles of Chemistry*, published in 1891.

Like Newlands, Mendeleev arranged the elements according to increasing atomic mass and observed the periodic repeating of properties. Julius Lothar Meyer independently created a very similar periodic table in 1864, yet is not given much credit for his contribution. Lothar Meyer's proposed table, which consisted of only 28 elements, did not leave room for undiscovered elements and was arranged by valence, instead of by atomic mass.

In *Principles of Chemistry*, Mendeleev wrote: "I began to look about and write down the elements with their atomic weights [masses] and typical properties, analogous elements and like atomic weights on separate cards, and this soon convinced me that the properties are in periodic dependence upon their atomic weights." These words clearly show his insight into the periodic nature of the elements.

Mendeleev's contribution is extraordinary because he observed "holes" in his arrangement of elements. He predicted that there were undiscovered elements. He even predicted their properties. He also proposed the presence of new row of elements: the actinide series.

Mendeleev's achievements earned him a great place in the science history books. To honour his contribution, element 101 will forever be known as mendelevium.

Discovering Yet More Elements

The twentieth century brought many additions to the periodic table. Henry Moseley, a British physicist, experimentally confirmed Bohr's model of the atom. Moseley's work also showed that there were additional gaps in the periodic table. The elements technetium and promethium were eventually discovered, and fitted neatly into the spaces left for them. Tragically, Moseley was killed in the Battle of Gallipoli, in Turkey, in 1915. His death was seen as a serious blow to the scientific community.

The American chemist Glenn Seaborg helped to discover an additional 10 elements: plutonium, americium, curium, berkelium, californium, einsteinium, fermium, mendelevium, nobelium, and element 106—which was named seaborgium in his honour. Seaborg's list of achievements is truly amazing. He served as a nuclear chemist on the Manhattan Project (the development of the atomic bomb during World War II), taught at the University of California, and headed the United States Atomic Energy Commission in the 1960s.

Over the years, the international (IUPAC) protocol for naming elements has led to names honouring scientists, countries, laboratories, and even deities from around the world. More recently, IUPAC has given temporary names to newly discovered elements until establishing their permanent names. For example, element 113 is temporarily named ununtrium.

So how many elements are there? The quest for new elements continues as scientists are predicting the characteristics of element 119 and beyond.

Alternatives to the Periodic Table

As more and more elements are discovered, scientists are proposing some interesting alternatives to our familiar periodic table. **Figure 4** shows the "Chemical Galaxy," an intriguing arrangement developed by Phillip Stewart in 2004. In Stewart's arrangement, the periodic trends are represented in a circular or spiral fashion. Other scientists have proposed different arrangements. What will the future bring for the periodic table of the elements?

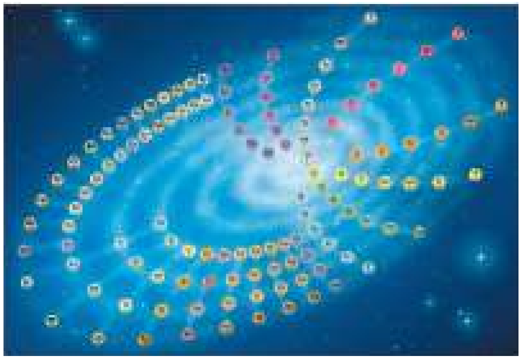


Figure 4 In the galactic model, elements with similar properties are aligned along curving spokes leading out from the centre.

Scientific progress works in strange ways: sometimes steady and unwavering, sometimes sudden and daring, and, occasionally, backwards or unpredictable. The periodic table has been through many changes as new evidence appears. In these days of electronic networking, each element has its own social networking page! What will the future bring for the periodic table of the elements?

Further Reading

Asimov, I. (1982). *Biographical encyclopedia of science and technology*. New York: Doubleday and Co.

Levi, P. (1985). *The periodic table*. New York: Knopf Publishing Group.

GO TO NELSON SCIENCE

1.6 Questions

1. Draw a timeline outlining the main steps taken in the formation of the periodic table. K/U C
2. Do you feel that Mendeleev should receive the main credit for the structure of today's periodic table? Defend your position in a paragraph. T/I
3. Identify factors that help to speed up scientific progress. Identify factors that slow down scientific progress. A
4. (a) Suggest an advantage that the chemical galaxy has over the traditional periodic table.
   (b) Research an alternative arrangement of the elements. Why was this arrangement proposed? T/I

1.7 Periodic Trends in Atomic Properties

The periodic table contains a great deal of information and presents numerous trends and patterns. In this section, you will explore atoms more closely. You will look at additional atomic properties and see how they are all connected. You will also look for trends among the properties.

atomic radius a measurement of the size of an atom, usually expressed in picometres (pm); the distance from the centre of an atom to the outermost electrons

Atomic Radius

An individual atom is incredibly small. Its radius is expressed in picometres (one picometre = 1 pm = 1×10^{-12} m).

Measuring the radius of a tennis ball would be a fairly straightforward task. We could easily measure the diameter with a ruler and divide by 2 to obtain the radius. The ball has definite boundaries and our measuring device is effective for the task.

How might we define and measure the size of something as tiny as an atom? This is a difficult task for a few reasons. The boundaries of an atom are not as definite and it is difficult to determine where the electron orbits end for an atom. Also, it is so incredibly small that we would need a very tiny ruler! In fact, a direct measurement is impossible, but scientists are able to determine the atomic radius indirectly.

Chemists define the radius of an atom in a few different ways. Simply, the **atomic radius** of an atom is defined as the distance from the nucleus to just beyond the outermost electrons (valence electrons). In a diatomic molecule (such as nitrogen, N_2, or oxygen, O_2), the atomic radius is defined as the distance between the two nuclei, divided by 2 (**Figure 1**).

Figure 2 shows the atomic radii of the elements in the periodic table. Look at the relative sizes of the atoms and examine the periodic trends associated with this property.

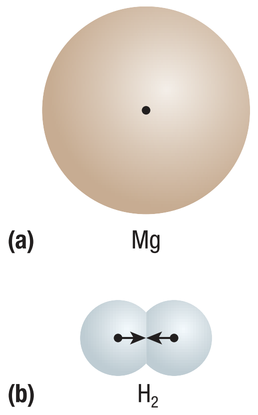


Figure 1 The atomic radii of (a) magnesium, an unbonded atom, and (b) hydrogen, a pair of bonded atoms

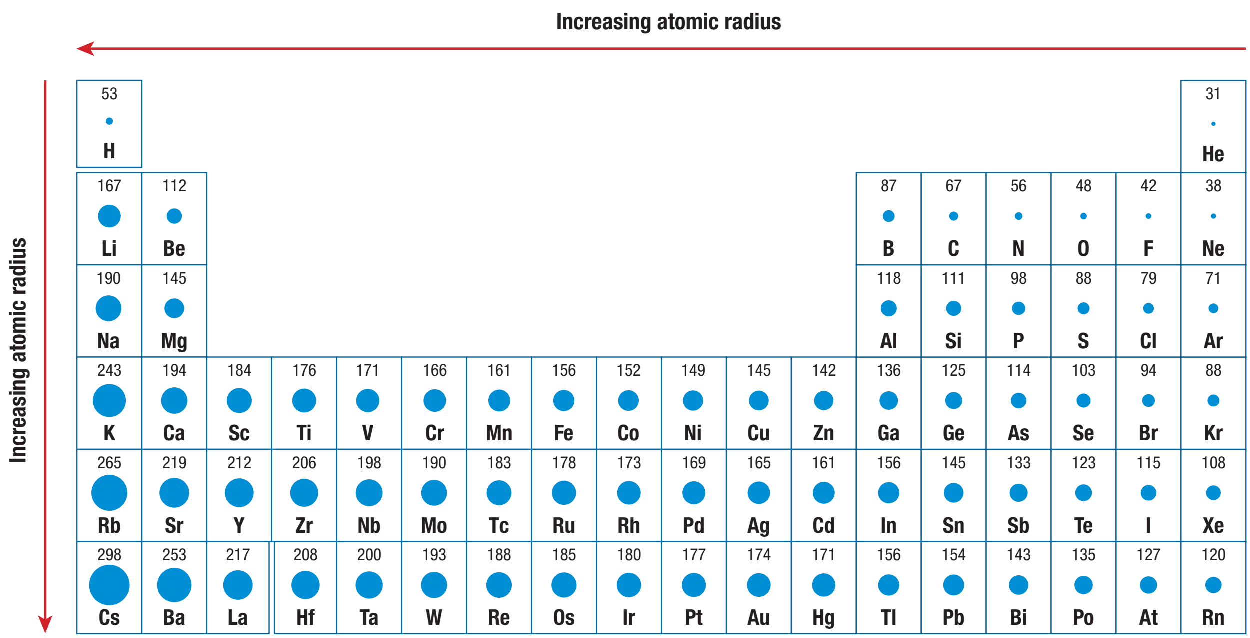


Figure 2 Atomic radius is a periodic trend. What trends can you observe and explain in the first six periods?

Every electron in an atom experiences a force of attraction toward the nucleus, known as the **effective nuclear charge.** The ideas of effective nuclear charge and energy levels help us to explain trends in atomic radius. As you move down a group, notice that the atomic radius increases. This is a result of adding another energy level as you progress from one period to the next. Each additional energy level is a greater distance from the nucleus. This occurs because the electrons of the inner levels shield the outer electrons from the full charge of the nucleus. As a result, the outer electrons are not as strongly attracted by the nucleus, resulting in a larger atomic radius.

effective nuclear charge the net force experienced by an electron in an atom due to the positively charged nucleus

Atomic radii decrease as you move from left to right across a period. This trend is a result of the increasing positive charge of the nucleus. As you move across a period, each element has one more proton and one more electron than the element before it. Since the additional electrons are being added to the same energy level, no screening of the nucleus occurs. Therefore, as the nuclear charge increases moving across a period, the attraction for electrons also increases. The increased attraction pulls the electrons closer to the nucleus resulting in a smaller atomic radius. As result of this trend, the noble gases are the smallest atoms in their respective periods.

Ionic Radius

When an atom gains or loses one or more electrons, it forms an ion. When sodium's one valence electron is removed, the remaining entity is a positively charged sodium ion, Na^+. When chlorine gains one more valence electron, it becomes a negatively charged chloride ion, Cl^-.

The size of an ion is described by the **ionic radius**. Like atomic radius, it is measured in picometres (pm). Is the ionic radius larger or smaller than the atomic radius of the same element? To answer this, we once again need to consider the forces acting on the valence electrons.

ionic radius a measurement of the size of an ion, usually expressed in picometres (pm); the distance from the centre of an ion to the outermost electrons

First, let us consider what happens when an atom of sodium becomes a sodium ion. With the removal of the only valence electron, there is one fewer electron orbit. A sodium ion has an ionic radius of 95 pm compared to the atomic radius of 190 pm for a neutral sodium atom (**Figure 3(a)**). Positive ions (cations) are always smaller than their original neutral atoms. We can explain this by referring to the force of attraction exerted by the nucleus. The force is now shared among fewer electrons, so is slightly stronger on each one. The net result is that the nucleus pulls each electron a little closer.

Negative ions (anions), on the other hand, are always bigger than their original neutral atoms. For example, a chloride ion has an ionic radius of 181 pm, whereas a chlorine atom has an atomic radius of 79 pm (**Figure 3(b)**). When an atom gains an electron, repulsion among the electrons increases while the nuclear charge remains the same. This results in a larger ionic radius. The theoretical explanation for this is that the effective nuclear charge is now shared among more electrons. The force of attraction exerted by the nucleus is therefore slightly weaker on each electron. The nucleus cannot hold each electron quite so close to itself.

WEB **LINK**

To see a comparison of atomic radii and ionic radii for all the elements,

GO TO NELSON SCIENCE

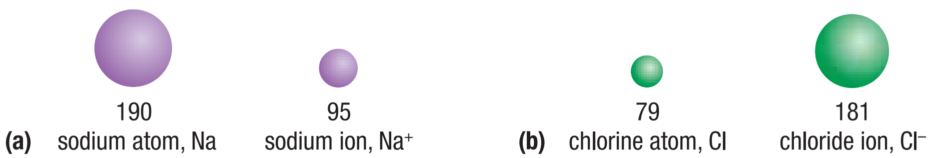


Figure 3 Notice how a positive ion is larger than its atom, whereas a negative ion is smaller than its atom. Suggest a reason for this observation.

Mini Investigation

Organizing Aliens

Skills: Performing, Observing, Analyzing, Communicating

Equipment and Materials: a set of cards with drawings of aliens with various characteristics

1. Your teacher will distribute an almost-complete set of cards to each group of students. These cards have drawings of aliens with various features (**Figure 4**).
2. In your group, organize the aliens in a meaningful arrangement or table.
3. Predict what the missing alien(s) would look like, and where it/they would fit in your table.
4. Obtain the missing card(s) from your teacher and compare your prediction with the actual card(s).

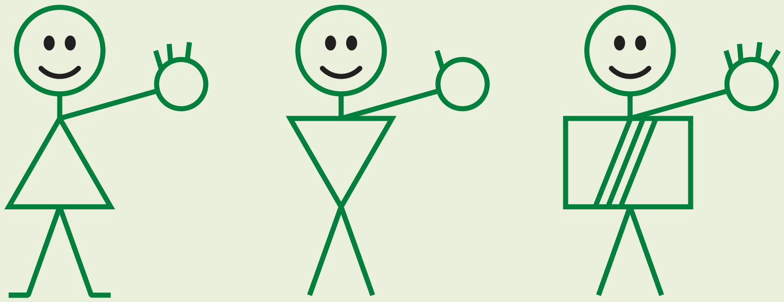


Figure 4 Classify the aliens according to trends in their characteristics.

A. While you were arranging the aliens, what patterns did you notice first? T/I

B. How many groups did your alien table have? T/I

C. How many periods did your alien table have? T/I

D. How did your prediction compare with the missing card(s)? T/I

E. How is your alien table analogous to the formation and usefulness of the periodic table? T/I

F. Think of another example in which items are organized by their properties into some type of a table. A

Ionization Energy and Electron Affinity

Valence electrons are bound to an atom by their attractive force to the nucleus. Removing electrons requires energy. Gaining electrons often releases energy.

Ionization Energy

Have you ever accidentally placed a metal spoon or a sheet of aluminum foil in a microwave oven? The resulting sparks are dramatic and potentially dangerous. Yet you can heat water molecules in a plastic or ceramic cup without incident. The sparks that we see are a result of electrons leaving the metal atoms. This is evidence that metal has a weaker hold on its electrons than do the hydrogen and oxygen atoms in water. The hold that an atom has on its electrons is an important periodic property known as ionization energy.

ionization energy the quantity of energy required to remove an electron from an atom or ion in the gaseous state

Ionization energy is the quantity of energy required to remove a single valence electron from an atom or ion in the gaseous state. If an electron is removed from any metal atom, a positive ion (cation) is formed. For example, it takes 520 kJ/mol of energy to ionize lithium, Li.

General ionization equation:

$X(g) + \text{energy} \rightarrow X^{+}(g) + e^{-}$

Specific example:

$Li(g) + 520\ \text{kJ/mol} \rightarrow Li^{+}(g) + e^{-}$

The unit for ionization energy is kJ/mol, or kilojoules per mole. The kilojoule (kJ) is a unit of energy (1 kJ = 1000 J) and the mole (mol) is a standard quantity of a substance. You will learn much more about the mole in Unit 3.

As there may be more than one electron that can possibly be removed, we usually specify which electron we are dealing with. The first ionization energy is the energy required to remove the most loosely held electron from an atom or ion (**Figure 5**). The second ionization energy is the energy required to remove the next most loosely bound electron, and so on.

Investigation 1.7.1

Graphing Periodic Trends (p. 45)
In this investigation you will create a graph of first ionization energy and atomic radius against atomic number, using provided data, to look for periodic trends.

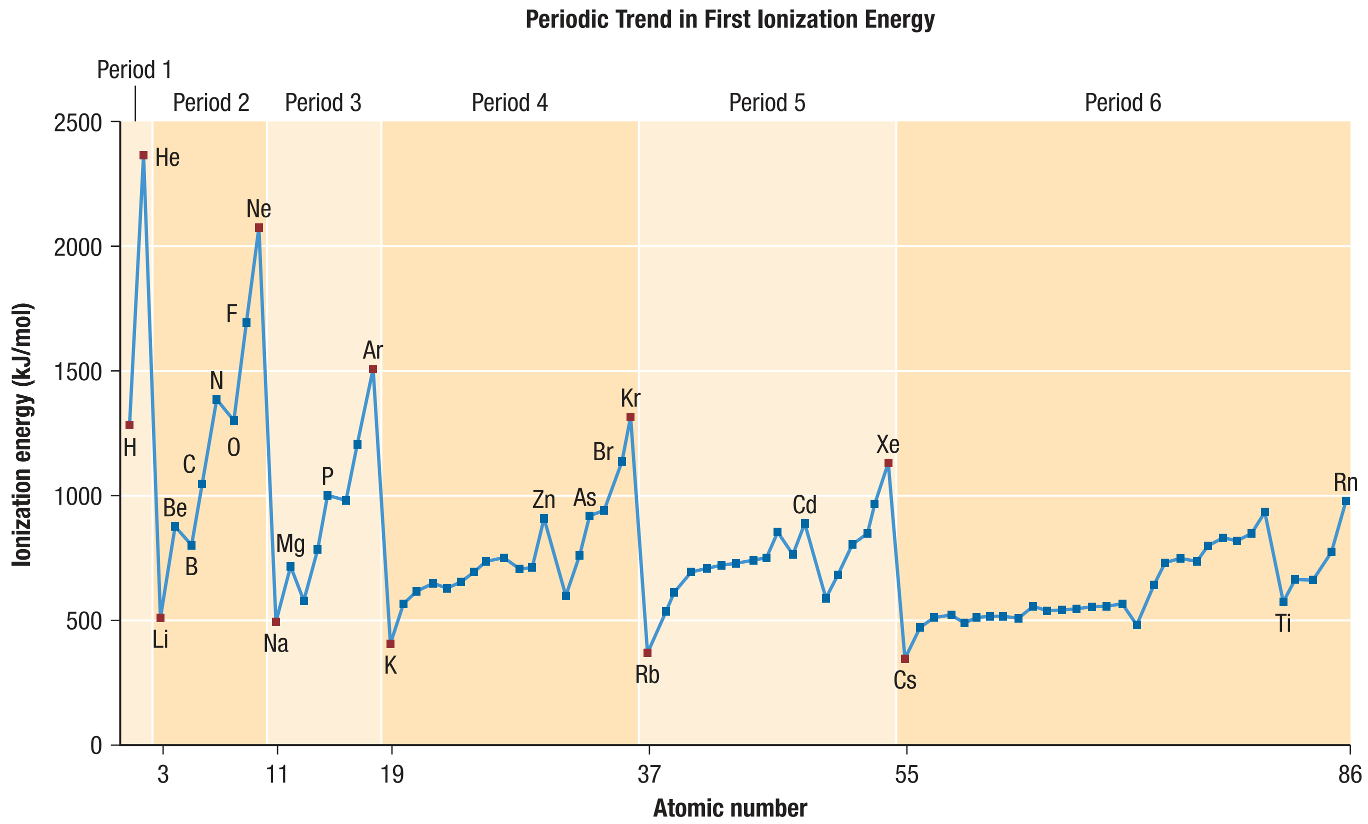


Figure 5 First ionization energy shows periodic trends. What trends can you observe and explain?

Figure 5 shows that, as we move down a group, the ionization energies tend to decrease. Hydrogen, H, has an ionization energy of 1312 kJ/mol; lithium, Li, has an ionization energy of 520 kJ/mol; and sodium, Na, has an ionization energy of 496 kJ/mol. Scientists reason that less energy is required to remove an electron as we move down a group because the force of attraction between the electron and the nucleus decreases as the atomic radius increases. Simply put, the farther away an electron is from the nucleus, the easier it is to remove it.

Look from left to right across a period. Notice that the ionization energy tends to increase. For example, the ionization energy of calcium (atomic number 20) is 590 kJ/mol; the ionization energy of arsenic, As, is 947 kJ/mol; and the ionization energy of bromine, Br, is 1140 kJ/mol. This trend is supported by our explanation that, as atomic radius decreases, the pull on the outermost electrons increases and thus it is harder to remove them. (Remember that atoms are generally smaller on the right side of the periodic table than on the left side.) Following this pattern, the ionization energies for the noble gases are extremely large. It is very difficult to remove any of their electrons.

Electron Affinity

electron affinity the energy change that occurs when an electron is added to a neutral atom in the gaseous state

If a neutral atom gains one or more electrons, it becomes negatively charged. The **electron affinity** for an element is defined as the energy change that occurs when an atom in the gaseous state gains an electron. Electron affinity, like ionization energy, is measured in kJ/mol.

When considering electron affinity, chemists need to be specific about the number of electrons that are acquired. The energy released when an atom acquires one electron is referred to as the first electron affinity. The energy released when a second electron is acquired is the second electron affinity, and so on. For example, 349 kJ/mol of energy is released when an atom of chlorine gas gains an electron.

General electron affinity equation:

$X(g) + e^- \rightarrow X^- (g) + \text{energy}$

Specific example:

$Cl(g) + e^- \rightarrow Cl^- (g) + 349 \text{ kJ/mol}$

Table 1 shows the electron affinities for the first 20 elements. What trends can you observe? The electron affinities of helium, beryllium, nitrogen, magnesium, and argon are all less than zero. This means that energy is not released when these elements gain an electron. On the contrary, energy is needed. How can we explain this? When an electron is added to an atom, it is attracted by the nucleus and also repelled by the atom's electrons. If the attractive force is greater, the electron is received, an anion is formed, and energy is released. If, however, the repulsive force is greater, energy is required to add the electron.

Table 1 Electron Affinities and Atomic Radii of the First 20 Elements

Element	Electron affinity (kJ/mol)	Atomic radius (pm)	Element	Electron affinity (kJ/mol)	Atomic radius (pm)
H	73	53	Na	53	190
He	0	31	Mg	0	145
Li	60	167	Al	43	118
Be	0	112	Si	134	111
B	27	87	P	72	98
C	154	67	S	200	88
N	7	56	Cl	349	79
O	141	48	Ar	0	71
F	328	42	K	48	243
Ne	0	38	Ca	2	194

As you investigate the patterns in the periodic table, you might notice that the electron affinity for elements decreases as you move down a group and it increases as you move across a period. Electron affinity shows the same trends as ionization energy. What is the theoretical explanation for the trend in electron affinity values? As you move to the right across the periodic table, the number of valence electrons increases and the atomic radius decreases. The force of attraction between the nucleus and the valence electrons increases, so more energy is released when a new electron is acquired.

LEARNING **TIP**

Using Analogies in Chemistry

An analogy might help you understand ionization energy and electron affinity better. Imagine paper clips (electrons) attached to a strong magnet (nucleus). You have to use energy to pry the paper clips away. In the same way, it takes energy—ionization energy—to remove the electrons from the atom. When you bring a paper clip close to the magnet, it jumps onto the magnet with a snap. The snap represents electron affinity. The louder the snap, the greater the electron affinity.

Interpreting Data

In this section, data are provided in several formats: as a figure with representative drawings (Figures 2 and 3), as a graph (Figure 5), and in a table (Table 1). You may find one format easier to use than the others. If you have difficulty interpreting data in the future, you could convert it into your preferred format.

1.7 Summary

- Atomic radius, ionic radius, ionization energy, and electron affinity are all atomic properties that exhibit periodic trends (**Figure 6**).
- The trends for atomic radius can be explained using the concept of effective nuclear charge, the attractive force of a positively charged nucleus on the outermost electrons in an atom.
- The periodic trends for ionization energy and electron affinity can be explained using the concept of effective nuclear charge and the trends in atomic radius.

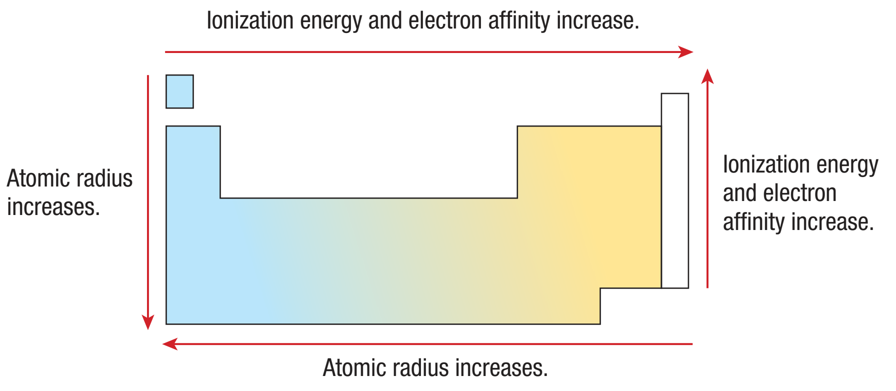


Figure 6 Periodic trends

1.7 Questions

1. Refer to Figure 2 to identify which element in each of the following groups has the largest atomic radius: K/U
   (a) halogens
   (b) alkali metals
   (c) noble gases
2. Arrange the following atoms in order of increasing atomic radius (smallest to largest): K, Rb, Cs, F, Li, C T/I
3. Predict which has the larger radius: a magnesium atom or a magnesium ion. Explain. T/I
4. Compare and contrast ionization energy and electron affinity. K/U
5. For each of the following pairs of atoms, state which would have
   (a) a larger size
   (b) a higher ionization energy
   (c) a lower electron affinity

   Support each of your predictions with an explanation. T/I
   (i) Ca, K (ii) Al, P (iii) I, F (iv) Li, Ra
6. For each of the following pairs of entities, state which entity is larger. Support each of your predictions with an explanation. T/I
   (a) S, S^{2-}
   (b) Ca^{2+}, Ca
   (c) F^-, Li^+
   (d) Cl^-, Br^-
7. Consider two isotopes of chlorine, Cl-35 and Cl-37. Would you expect the first ionization energies for these two isotopes to be the same or different? Explain your prediction. A
8. An anomaly is a deviation or departure from the normal. Using the data in this section, find an anomaly for
   (a) the periodic trend for atomic radius
   (b) the periodic trend for ionization energy

   Propose why such an anomaly might exist. T/I
9. The periodic table in Figure 2 shows the atomic radii for the first six periods. Write a hypothesis predicting the atomic radii for the following elements in Period 7: T/I A
   (a) francium
   (b) unumseptium
10. A ceramic plate has a shiny rim. The ceramic plate is mostly composed of silicon and oxygen atoms. The shiny rim is mostly silver. When the plate is heated in a microwave oven, sparks are observed. Compare the ionization energy values of silver to those of the silicon and oxygen atoms. Explain what is happening in the oven. Why are microwaveable containers usually non-metallic? Why are most pots and pans metallic? K/U A
11. Scientists are trying to synthesize element 119. Based on your knowledge of the trends in the periodic table, predict three physical properties and one chemical property for element 119. T/I A
12. Predict the atomic radii of the first three elements in Period 7. Explain your predictions. T/I A

Investigation 1.4.1 OBSERVATIONAL STUDY

SKILLS MENU

• Questioning	• Planning	• Observing
• Researching	• Controlling Variables	• Analyzing
• Hypothesizing	• Performing	• Evaluating
• Predicting		• Communicating

The Nuts and Bolts of Atomic Mass

The atomic mass recorded on the periodic table is the weighted average of all the naturally occurring isotopes for that element. One of the challenges when working with atomic models and concepts is that we cannot see individual atoms—let alone measure their mass. To help you to understand the concept of atomic mass, you will use nuts and bolts to build models of the different isotopes of the element boltium (an element very similar to hydrogen!). In this model, the bolts will represent protons and the nuts will represent neutrons (**Figure 1**). You will find the mass and abundance of these different isotopes and you will use this data to determine the atomic mass for boltium.

Figure 1 Nuts and bolts represent neutrons and protons of the element boltium.

Purpose

To gain a better understanding of the concept of atomic mass

Equipment and Materials

- 20 bolts
- 40 nuts
- plastic cup
- electronic balance

Procedure

SKILLS HANDBOOK A2.1, A2.4

1. Create a table in which to record your observations.
2. Obtain a sample of boltium from your teacher. Your sample will contain a mixture of different boltium isotopes: boltium-1, boltium-2, and boltium-3.
3. Place your sample of boltium in a plastic cup. Measure and record the total mass of the sample in the cup.
4. Carefully empty the contents onto a large flat surface.
5. Measure and record the mass of the empty cup.
6. Sort the different isotopes of boltium. Count and record the total number of atoms of each isotope.
7. Measure and record the mass of all of the boltium-1 atoms.
8. Repeat Step 7 for the other two isotopes.

Analyze and Evaluate

(a) From your measurements in Steps 3 and 4, calculate the total mass of your sample of boltium. T/I

(b) Calculate the isotopic abundance for each isotope of boltium. T/I

$$\%\text{ abundance of boltium-1} = \frac{\#\text{ of atoms of boltium-1}}{\text{total }\#\text{ of atoms}} \times 100\%$$

(c) Calculate the mass for your entire sample by multiplying your atomic mass by the total number of atoms. T/I

(d) Compare the measured mass of your sample to the calculated mass of your sample determined in Question (c). Explain your answer. T/I

Apply and Extend

(e) Determine the atomic mass for one atom of boltium using the mass of each isotope and the isotopic abundance for each isotope. T/I

(f) Describe the difference between the isotopes of neon-19, neon-20, and neon-22. K/U

(g) Define the terms isotope, isotopic abundance, and atomic mass. K/U

(h) This activity is intended to provide an analogy or comparison to isotopes and atomic mass. Are nuts and bolts good analogies for isotopes of atoms? Explain the benefits and drawbacks of using this analogy. T/I

Investigation 1.5.1 OBSERVATIONAL STUDY

SKILLS MENU

- Questioning
- Researching
- Hypothesizing
- Predicting
- Planning
- Controlling Variables
- Performing
- Observing
- Analyzing
- Evaluating
- Communicating

The Search for Patterns

As you examine the physical and chemical properties of elements in groups and periods, you will be able to connect these properties to the atomic structure of the elements.

This two-part investigation allows you to explore the patterns in reactivity in the periodic table. In Part A, you will observe what happens when five metals (calcium, magnesium, copper, sodium, and lithium) are placed in water. In Part B, you will test the reaction of four metals (magnesium, iron, zinc, and aluminum) with dilute hydrochloric acid.

Purpose

To observe relationships between the reactivity of elements and their location on the periodic table

Equipment and Materials

SKILLS HANDBOOK A1.1, A1.2

Part A: Reactivity of Metals in Water

- chemical safety goggles
- lab apron
- safety gloves
- 250 mL beaker
- test tube (18 mm × 150 mm)
- test-tube rack
- scoopula
- spark lighter
- Bunsen burner clamped to a retort stand
- test-tube clamp
- red and blue litmus paper or pH paper
- paper towel
- small samples of
  - calcium, Ca(s)
  - magnesium, Mg(s)
  - copper, Cu(s)
- wooden splint
- small samples of
  - sodium, Na(s) (to be handled by your teacher only)
  - lithium, Li(s) (to be handled by your teacher only)

This investigation involves the use of open flames. Tie back long hair and secure loose clothing.

Calcium, sodium, and lithium are highly reactive and may burn in water. Do not touch any of these metals with your hands. Only your teacher should handle sodium and lithium.

Part B: Reactivity of Metals in Hydrochloric Acid

- chemical safety goggles
- lab apron
- safety gloves
- 20 mL graduated cylinder
- 4 test tubes (18 mm × 150 mm)
- test-tube rack
- scoopula
- small samples of
  - magnesium, Mg(s)
  - iron, Fe(s)
  - zinc, Zn(s)
  - aluminum, Al(s)
- dilute hydrochloric acid (0.5 mol/L)
- spark lighter
- Bunsen burner clamped to a retort stand
- paper towel
- wooden splint
- masking tape or test-tube stopper

Hydrochloric acid is an irritant. Wash any spills on skin or clothing immediately with plenty of cold water. Report any spills to your teacher.

This investigation involves the use of open flames. Tie back long hair and secure loose clothing.

Procedure

SKILLS HANDBOOK A2.1, A2.4

1. Read the entire procedure before conducting the lab. Prepare a list of safety concerns. Discuss how these concerns should be addressed in the lab.
2. Prepare appropriate data tables in which to record your observations. Write titles for all data tables.
3. Write a procedure, based on **Figure 1** (on the next page), to collect any gas produced when a metal reacts with water. Include any necessary safety precautions. Ask your teacher to approve your procedure.

Figure 1 Collecting gas produced by the reaction of a metal in water

Part A: Reactivity of Metals in Water

4. Put on your chemical safety goggles, lab apron, and safety gloves.
5. Half-fill a clean 250 mL beaker with tap water.
6. Test the water with either litmus paper or pH paper. Record your results.
7. Using a folded, square piece of paper towel, obtain a piece of calcium from your teacher. Observe and record its physical properties.
8. Stand at arm's length from the beaker. Using the scoopula, carefully place the calcium in the water. Record your observations.
9. When the reaction is complete, test the resulting solution with either litmus paper or pH paper. Record your results.
10. Dispose of the waste materials according to your teacher's instructions. If possible, recycle your materials.
11. Repeat this procedure using samples of magnesium and copper. Record all your results.
12. Select a metal that produced a gas when reacting with water. Use your approved procedure from Step 3 to collect a test-tubeful of the gas.
13. Set the test tube, mouth down, in the test-tube rack.
14. Use the spark lighter to light your Bunsen burner. Light a wooden splint from the flame. Turn off the burner.
15. Using a test-tube clamp, lift the test tube and stand it, with its opening at the top, in the test-tube rack.
16. Insert the burning splint inside the mouth of the test tube. Record your observations.
17. Extinguish the splint with water from the tap. Dispose of any waste according to your teacher's instructions.
18. Watch your teacher perform the same tests on sodium and lithium. Observe and record the results.

Part B: Reactivity of Metals in Hydrochloric Acid

19. Put on your chemical safety goggles, lab apron, and safety gloves.
20. Measure out 20.0 mL of dilute hydrochloric acid. Pour 5.0 mL into each of four clean test tubes in your test-tube rack.
21. On folded pieces of paper towel, obtain small quantities of magnesium, iron, zinc, and aluminum from your teacher.
22. Record the physical properties of each of the metals.
23. One at a time, carefully drop each of the metals into a different test tube of hydrochloric acid.
24. Observe over a 20 min period. Occasionally feel the test tubes with your hand to find out if they get warmer as any reaction proceeds. Record your observations.
25. If you observe bubbles, collect a sample of the gas by covering the test tube loosely with a piece of masking tape or a rubber stopper. While the test tube is in the test-tube rack, test the resulting gas by removing the stopper and quickly holding a glowing splint just inside the mouth of the test tube. (Light the splint using the procedure in Step 14.)
26. Recycle or dispose of any waste according to your teacher's instructions.

Analyze and Evaluate

(a) Rank the elements tested in Part A from least reactive to most reactive. T/I C

(b) Sort the elements that you tested into groups. State the apparent order of reactivity as one proceeds down a group. Does reactivity increase or decrease? T/I

(c) Sort the elements that you tested into periods. State the apparent order of reactivity as one proceeds across a period. Does reactivity increase or decrease? T/I

(d) Is the solution that is produced when a metal reacts with water acidic or basic? T/I

(e) Based on your gas test observation in Step 16, what gas is produced? T/I

(f) Rank the elements tested in Part B from least reactive to most reactive. T/I

(g) Does the reactivity increase or decrease as one moves across a period of elements? T/I

(h) Based on your gas test results, what is the gas produced in these reactions (Steps 16 and 25)? T/I

(i) Does the reactivity increase or decrease as one moves down a group of elements? T/I

(j) Evaluate this investigation. Did the design enable you to collect enough evidence to answer the questions? How could it have been improved? Would your suggested improvements raise any safety concerns? T/I

Apply and Extend

(k) Predict what might happen if you were to drop a piece of potassium into a beaker of water. T/I

(l) Connect the trends observed in Parts A and B to atomic radius, ionization energy, and electron affinity. Write a paragraph to explain the trends. T/I C

Investigation 1.7.1 CORRELATIONAL STUDY

Graphing Periodic Trends

We can identify several periodic trends when we explore the arrangement of elements in the periodic table. In this investigation you will graph provided data for the atomic radius and the first ionization energy of the first 20 elements. You can use graph paper or graphing software to create your graph. You will search for the trends and attempt to explain the relationship among these properties.

SKILLS MENU

• Questioning	• Planning	• Observing
• Researching	• Controlling Variables	• Analyzing
• Hypothesizing	• Performing	• Evaluating
• Predicting		• Communicating

Purpose

To discover how two periodic properties, atomic radius, and first ionization energy, vary with the atomic number for the first 20 elements

Variables

What variables are involved in this investigation?

Equipment and Materials

- graph paper or graphing software

Procedure

SKILLS HANDBOOK A2.3

1. Find the data for atomic radii and first ionization energy in Section 1.7, Figures 2 and 5.
2. Select a graphing program or obtain a piece of graph paper.
3. Label the horizontal axis "Atomic number." It should go from 1 to 20, representing the elements H to Ca. Choose an appropriate scale.
4. Label the vertical axis on the right "Atomic radius." Choose an appropriate scale and include the appropriate units.
5. Label the vertical axis on the left "First ionization energy." Choose an appropriate scale and include the appropriate units.
6. Plot the values of atomic radius versus atomic number. Join the points using straight lines.
7. Plot the values of first ionization energy versus atomic number. Join the points using a different colour.
8. Write a title and provide a legend to indicate the two different trend lines.

Analyze and Evaluate

(a) Describe any trends for atomic radius versus atomic number within a period and within a group. T/I

(b) In which group of elements do the atoms have the largest radii as you move across a period? T/I

(c) In which group of elements do the atoms have the smallest radii as you move across a period? T/I

(d) Describe any trends for first ionization energy versus atomic number within a period and within a group. T/I

(e) In which group of elements do the atoms have the greatest first ionization energy as you move across a period? T/I

(f) In which group of elements do the atoms have the smallest first ionization energy as you move across a period? T/I

(g) Compare the trend for atomic radius and the trend for ionization energy. How do these properties relate to one another? T/I

(h) Using your graphed data, predict the atomic radius and the ionization energy for scandium, Sc, and titanium, Ti. Research the accepted values. Compare your predictions to the actual values. How close were your predictions? T/I

Apply and Extend

(i) Predict the relationship between electron affinity and atomic number. Sketch a graph of these two variables. Label the axes properly. T/I C

GO TO NELSON SCIENCE

Summary Questions

1. Create a study guide based on the points listed in the margin on page 6. For each point, create three or four sub-points that provide relevant examples, explanatory diagrams, or general equations.
2. Look back at the Starting Points questions on page 6. Answer these questions using what you have learned in this chapter. Compare your latest answers with those that you wrote at the beginning of the chapter. Note how your answers have changed.

Vocabulary

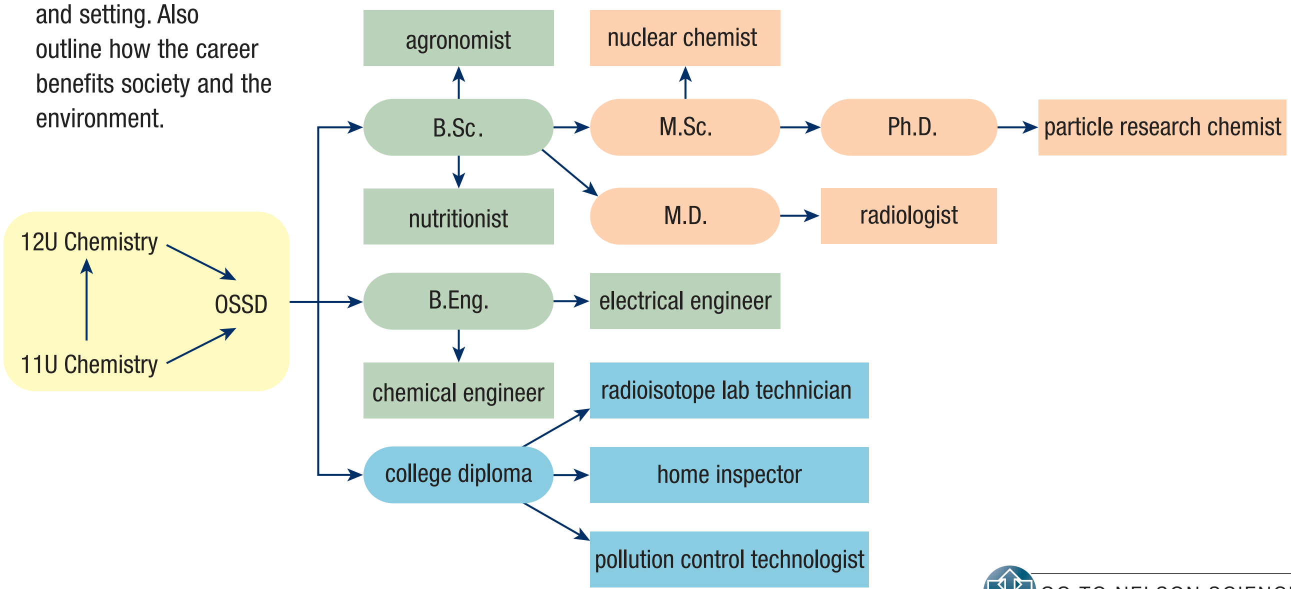
matter (p. 8)
empirical knowledge (p. 9)
theoretical knowledge (p. 9)
theory (p. 9)
atom (p. 11)
electron (p. 12)
proton (p. 13)
neutron (p. 13)
energy level (p. 14)
valence shell (p. 14)
valence electron (p. 14)
atomic number (*Z*) (p. 15)
mass number (*A*) (p. 15)
atomic mass unit (p. 16)
full or stable octet (p. 17)
octet rule (p. 17)
ion (p. 17)
cation (p. 18)
valence (p. 18)
anion (p. 18)
multivalent (p. 19)
polyatomic ion (p. 20)
isotope (p. 23)
isotopic abundance (p. 24)
mass spectrometer (p. 24)
radioactive decay (p. 25)
nuclear radiation (p. 25)
alpha particle (p. 25)
beta particle (p. 25)
gamma ray (p. 25)
radioisotope (p. 25)
radioactive (p. 25)
atomic mass (p. 26)
metalloid (p. 31)
group (p. 31)
period (p. 31)
periodic law (p. 31)
Lewis symbol (p. 32)
atomic radius (p. 36)
effective nuclear charge (p. 37)
ionic radius (p. 37)
ionization energy (p. 38)
electron affinity (p. 40)

CAREER PATHWAYS

Grade 11 Chemistry can lead to a wide range of careers. Some require a college diploma or a B.Sc. degree. Others require specialized or postgraduate degrees. Here are just a few pathways to careers mentioned in this chapter.

SKILLS HANDBOOK A7

1. Select two careers, related to atomic structure, that you find interesting. Research the educational pathways that you would need to follow to pursue these careers. What is involved in the required educational programs? Prepare a brief report of your findings.
2. For one of the two careers that you chose above, describe the career, main duties and responsibilities, working conditions, and setting. Also outline how the career benefits society and the environment.



GO TO NELSON SCIENCE

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

For each question, select the best answer from the four alternatives.

1. What is included in the study of chemistry? (1.1) K/U
   (a) motion and forces
   (b) composition and interaction of matter
   (c) the economics of manufacturing
   (d) forces and technology
2. Who first theorized that atoms of the same element are alike, and atoms of different elements are different? (1.2) K/U
   (a) Dalton
   (b) Democritus
   (c) Rutherford
   (d) Thomson
3. Which term applies to the number of protons in an atom? (1.2) K/U
   (a) mass number
   (b) atomic mass
   (c) atomic number
   (d) atomic mass unit
4. According to the octet rule, how many valence electrons are in the valence shells of most atoms that are stable? (1.3) K/U
   (a) 2
   (b) 8
   (c) 18
   (d) 36
5. Which statement is true about an alpha particle? (1.4) K/U
   (a) It contains electrons.
   (b) It has a negative charge.
   (c) It is a type of electromagnetic radiation.
   (d) It has the same structure as a helium-4 nucleus.
6. Three isotopes exist for potassium. Approximately 93 % of all potassium atoms have a mass of 39 u. 7 % of potassium atoms have a mass of 41 u. A very small proportion of potassium atoms have a mass of 40 u. What is the atomic mass of potassium? (1.4) K/U T/I
   (a) 33.40 u
   (b) 39.14 u
   (c) 40.07 u
   (d) 40.86 u
7. Which statement is true about magnesium, Mg, and strontium, Sr? (1.5) K/U
   (a) They are in the same group.
   (b) They are in the same period.
   (c) They are both metalloids.
   (d) They are both gases.
8. Which British scientist predicted the properties of elements that would eventually fill gaps in the periodic table? (1.6) K/U
   (a) Dobereiner
   (b) Seaborg
   (c) Newlands
   (d) Moseley
9. Based on its position on the periodic table, which of the following elements has the largest atomic radius? (1.7) K/U
   (a) iron, Fe
   (b) germanium, Ge
   (c) potassium, K
   (d) krypton, Kr

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

10. Mass is anything that has both mass and volume. (1.1) K/U
11. Empirical knowledge depends on theories. (1.1) K/U
12. Bohr's model of the atom includes neutrons located in energy levels. (1.2) K/U
13. An anion forms when an atom gains one or more electrons. (1.3) K/U
14. Another name for cupric oxide is copper(I) oxide. (1.3) K/U
15. All isotopes are radioactive. (1.4) K/U
16. The noble gases exist as individual atoms and are quite reactive. (1.5) K/U
17. IUPAC developed the rules for naming newly discovered elements. (1.6) K/U
18. Moving down a group, ionization energies tend to decrease. (1.7) K/U

To do an online self-quiz,
GO TO NELSON SCIENCE

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

Knowledge

For each question, select the best answer from the four alternatives.

1. Which of the following terms best describes the study of the composition and interaction of matter? (1.1) K/U
   (a) physics
   (b) science
   (c) chemistry
   (d) technology
2. Which term applies to the macroscopic world? (1.1) K/U
   (a) experimenting
   (b) imagining
   (c) conceptualization
   (d) theorizing
3. Who first theorized that the atom contains a central core called a nucleus? (1.2) K/U
   (a) Dalton
   (b) Thomson
   (c) Democritus
   (d) Rutherford
4. Which of the following ions always contain more than one atom? (1.3) K/U
   (a) polyatomic ions
   (b) cations
   (c) multivalent ions
   (d) anions
5. Different isotopes of the same element contain different numbers of which of these particles? (1.4) K/U
   (a) electron
   (b) proton
   (c) neutron
   (d) nucleus
6. Which statement is true about the elements sulfur, S, and chlorine, Cl? (1.5) K/U
   (a) They are in the same group.
   (b) They are in the same period.
   (c) They are both metalloids.
   (d) They are both halogens.
7. Element X is a light, very reactive metal that forms an oxide with the formula XO. This element also reacts with water to release hydrogen. The element belongs to which of the following groups? (1.5) K/U
   (a) alkali metals
   (b) alkaline earth metals
   (c) transition metals
   (d) rare earth metals
8. Which scientist is given credit for developing the periodic table that led to the modern version of the periodic table? (1.6) K/U
   (a) Meyer
   (b) Dobereiner
   (c) Newlands
   (d) Mendeleev
9. What is the energy change that occurs when an electron is gained by an atom in the gaseous state? (1.7) K/U
   (a) ionization energy
   (b) electron affinity
   (c) ionization
   (d) electronegativity

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

10. Air is an example of matter. (1.1) K/U
11. The mass number of an atom is the number of protons it contains. (1.2) K/U
12. Bohr's work explained why each element has a unique spectrum. (1.2) K/U
13. The nitrate ion and the sulfate ion are examples of polyatomic cations. (1.3) K/U
14. You can determine the atomic mass of an element by calculating the weighted average of the masses of all of its ions. (1.4) K/U
15. A Lewis symbol of an element shows only its symbol and its valence electrons. (1.5) K/U
16. Newlands arranged groups of three atoms that have similar properties. (1.6) K/U
17. Atomic radius is generally greater on the left side of the periodic table than on the right side. (1.7) K/U
18. A metal ion is larger than the atom from which it formed. (1.7) K/U

Match each term on the left with the most appropriate description on the right.

19.

(a) Dalton	(i) discovered the neutron
(b) Rutherford	(ii) discovered the electron
(c) Bohr	(iii) proposed the nuclear model of the atom
(d) Thomson	(iv) suggested that electrons exist in energy levels
(e) Chadwick	(v) proposed the existence of atoms
(f) Democritus	(vi) proposed that atoms are not created or destroyed in chemical reactions (1.2) K/U

Write a short answer to each question.

20. List five different branches of chemistry. (1.1) K/U
21. Compare and contrast science and technology. (1.1) K/U T/I
22. How do theoretical knowledge and empirical knowledge differ? (1.1) K/U T/I
23. What did Thomson observe and conclude about electrons by using a cathode ray tube? (1.2) T/I
24. An aluminum atom contains 13 protons and 14 neutrons. (1.2) K/U T/I
    (a) What is the atomic number of aluminum?
    (b) What is the mass number of this aluminum atom?
25. Neon, argon, and krypton are used in light fixtures. Helium is used in balloons and in protecting valuable paper documents from damage. These, and other, uses of noble gases are based on the chemical inertness of the gases. What is meant by the term "chemical inertness"? (1.3) K/U
26. State two ways in which polyatomic ions are important to human health. (1.3) K/U
27. What are three possible ways in which an atom can achieve a complete octet? (1.3) K/U
28. What is a polyatomic ion? (1.3) K/U
29. How are nitrate ions and nitrite ions similar and how are they different? (1.3) K/U
30. (a) What is an isotope?
    (b) How does a radioisotope differ from an isotope that is not radioactive? (1.4) K/U T/I
31. Suppose that an isotope with a mass number of 260 u undergoes radioactive decay. The decay produces nuclear radiation and an isotope with a mass number of 256 u. (1.4) K/U T/I
    (a) What type of radiation did the isotope emit when it decayed?
    (b) Explain your answer.
32. Draw Bohr–Rutherford diagrams for atoms of the following elements: (1.5) K/U
    (a) aluminum-27
    (b) oxygen-16
    (c) hydrogen-1
33. Draw Lewis diagrams for atoms of the following elements: (1.5) K/U
    (a) silicon
    (b) chlorine
    (c) potassium
    (d) krypton
34. (a) What is a periodic trend?
    (b) How do periodic trends relate to periodic law? (1.5) K/U
35. What is true about the valence electrons within a group on the periodic table? (1.5) K/U
36. (a) Which elements are considered to be representative elements?
    (b) Why are these elements called representative elements? (1.5) K/U
37. (a) Where are the metals located on the periodic table?
    (b) Where are the non-metals?
    (c) Where are the metalloids? (1.5) K/U
38. Some scientists made major contributions to the organization of elements. However, even small contributions added to what was known about the elements. (1.6) K/U
    (a) What contribution to the organization of elements was made by Hennig Brand?
    (b) How did Antoine Lavoisier contribute to the knowledge of elements?
39. What unit is usually used when communicating the radius of an atom? (1.7) K/U
40. (a) What energy change usually occurs when an atom gains an electron?
    (b) What energy change occurs when an atom loses an electron? (1.7) K/U
41. In your own words, describe what is meant by electron affinity. (1.7) K/U

Understanding

42. Many of the central ideas in chemistry come from the atomic theory. (1.1, 1.2) T/I
    (a) Why is knowledge of the atom considered to be theoretical knowledge?
    (b) Predict how atomic theory 30 years from now might compare to current atomic theory.
43. Mass is usually measured in grams or kilograms. (1.2) K/U T/I
    (a) Instead of grams or kilograms, what unit is most commonly used to measure the mass of an individual atom?
    (b) Explain why this unit is used instead of grams or kilograms.
44. (a) What did Rutherford observe in his gold foil experiment?
    (b) Why did his observations lead him to theorize that atoms have a positively charged nucleus? (1.2) K/U T/I
45. Many elements form more than one polyatomic ion with oxygen. (1.3) K/U T/I
    (a) Give an example of two or more polyatomic ions that contain the same element and oxygen.
    (b) Explain why you think this element forms more than one polyatomic ion with oxygen.

46. Nitrites in meat slow down the growth of bacteria and give the meat a desirable red colour. However, nitrites can form compounds that are carcinogenic. Even if no nitrites are added to meat, they might be present. How can nitrites be present in meat if only nitrates were added to the meat? (1.3) K/U

47. Create a table and complete it with the following information for the elements listed below:
    - the number of electrons in the atom
    - the number of valence electrons
    - whether it will gain or lose electrons (or neither gain nor lose) to become stable
    - whether it will form an anion or a cation, or neither
    - the symbols of the ion it forms (1.3) K/U T/I

    (a) neon, Ne
    (b) oxygen, O
    (c) magnesium, Mg

48. (a) List the three main parts of a mass spectrometer.
    (b) Describe what happens in each part during the analysis of a sample. (1.4) K/U

49. There are three forms of nuclear radiation. Gamma rays are used in radiation therapy to kill certain kinds of cancer cells. Why do you think this form of radiation is chosen, rather than either of the other forms? (1.4) T/I A

50. A radioisotope is sometimes called "parent material." As the parent material decays and emits nuclear radiation, it forms a new substance that is often referred to as "daughter material." If the parent material emits alpha rays as it decays, predict how atomic number and mass number of the daughter material will be different from those of the parent material. Explain your answer. (1.4) T/I

51. Examine the Lewis symbols in Figure 4 of Section 1.5. The elements in Groups 1 and 2 tend to lose electrons to another atom to become stable. Those in Groups 16 and 17 tend to gain electrons from other atoms. (1.3, 1.5) K/U T/I A
    (a) Name one element that might gain the one valence electron from a sodium atom and become stable.
    (b) Name one element that might gain the two valence electrons from a calcium atom and become stable.

52. Early periodic tables listed elements by increasing atomic mass from left to right, arranging elements with similar properties in the same column. How does this periodic table compare to the one we currently use? (1.5) K/U T/I

53. (a) Describe how you would write the Lewis symbol for sulfur.
    (b) How does the Lewis symbol for sulfur differ from the Bohr–Rutherford diagram for sulfur? (1.5) K/U T/I C

54. The periodic table currently in use is the result of scientists studying the elements for hundreds of years. Summarize the contributions the following scientists made in developing a system of organizing elements. (1.5, 1.6) K/U T/I C
    (a) Dmitri Mendeleev
    (b) Johann Dobereiner
    (c) John Newlands

55. The radius of an individual chlorine atom, Cl, is likely to be different from the radius of a chlorine atom in a chlorine molecule, Cl_2. Explain why these radii are different. (1.7) K/U

56. (a) In general, how does atomic radius change from top to bottom in a group on the periodic table?
    (b) Explain why this pattern in atomic radius occurs. (1.7) K/U T/I

57. (a) As you move from top to bottom in a group on the periodic table, how does the ionization energy change?
    (b) Why does ionization energy follow this pattern in a group?
    (c) As you move from left to right across a period, how does the ionization energy change?
    (d) Explain why ionization energy follows this pattern in a period. (1.7) K/U

Analysis and Application

58. A student performed an experiment in which she warmed three different liquids to 80 °C, and then measured their temperatures every 5 minutes as they cooled. She discovered that the liquids cooled at different rates. Was the knowledge from this experiment theoretical knowledge or empirical knowledge? Explain your answer. (1.1) T/I A

59. Classify each of the following as either "matter" or "not matter." Explain each classification. (1.1) K/U T/I A
    (a) thoughts
    (b) water
    (c) air
    (d) bricks
    (e) ideas

60. Alchemists did not discover any parts of the atom or otherwise develop any part of the atomic theory as it currently exists. Why do you think the work of alchemists was important to atomic theory? (1.2) K/U T/I

61. Magnesium is a strong, light metal that is primarily used in strong, low-density alloys. A magnesium atom contains 12 protons, 12 electrons, and 12 neutrons. (1.2) K/U T/I C
    (a) How many valence electrons does a magnesium atom have?
    (b) Explain how you determined the number of valence electrons in a magnesium atom.
62. Too much table salt, sodium chloride, NaCl, is bad for our health. From the effect each of these ions has on the human body, infer why too much salt can cause a person to have high blood pressure. (1.3) K/U T/I A
63. When iron metal reacts with oxygen, each iron atom first loses 2 electrons to an oxygen atom. The iron ion formed then loses another electron, forming a different ion. (1.3) K/U T/I A
    (a) Write the formula for the first iron ion formed. Name the ion, using both its classical and its IUPAC name.
    (b) Write the formula for the second iron ion formed. Name the ion, using both its classical and its IUPAC name.
64. Examine the system used for naming the polyatomic ions formed from chlorine and oxygen and those formed from nitrogen and oxygen, as shown in Table 2 in Section 1.3. (1.3) K/U T/I A
    (a) What is the difference between an *-ate* ion and an *-ite* ion?
    (b) What is the difference between the chlorate ion and the perchlorate ion?
    (c) What would be the formula for the periodate ion? Explain.
65. During nuclear decay, emission of alpha particles changes both the identity and the mass of the atom. Beta emission changes the identity of the atom but not its mass. Gamma emission changes neither the identity of the atom nor its mass. When a C-14 atom undergoes radioactive decay, an N-14 atom forms, along with a particle of radiation. (1.4) K/U T/I A
    (a) What particle of radiation forms during this radioactive decay?
    (b) Explain your answer.
66. A beam of radiation is emitted by a radioisotope. (1.4) K/U T/I A
    (a) How might passing the beam between electrically charged plates help you to identify what type of radiation was emitted?
    (b) How might you identify what type of radiation was emitted by focusing the beam on sheets of different materials?
67. Chromium is found in small quantities in the environment throughout Canada. This element occurs naturally as four different isotopes, as shown in **Table 1.**

Table 1 Isotopes of Chromium

Isotope	Abundance (%)
Cr-50	4.35
Cr-52	83.79
Cr-53	9.50
Cr-54	2.36

    What is the atomic mass of chromium? (1.4) T/I
68. Potassium has 21 isotopes. However, almost all of the potassium in a sample consists of the isotopes K-39, K-40, and K-41. The atomic mass of potassium is 39.1. (1.4) T/I A
    (a) Which isotope of potassium is present in greatest abundance?
    (b) Explain your answer.
69. Small quantities of boron do not negatively affect human health, but large quantities are dangerous. For this reason, Health Canada has set a maximum level of 5 mg/L of boron in drinking water. The two naturally occurring isotopes of boron are boron-10 and boron-11. The abundance of B-10 is 19.8 %. The remainder of the atoms are B-11. What is the atomic mass of boron? (1.4) K/U T/I A
70. Cobalt, Co, was once mined along with silver near Sudbury, Ontario. Cobalt is now mostly a by-product of copper and nickel mining. The element fluorine, F, has also been found in coal deposits in Western Canada. (1.5) K/U A
    (a) Look at the position of these two elements on the periodic table. Which of the elements is a representative element?
    (b) Based on your answer to (a), predict which element most closely follows the periodic law.
71. Most naturally occurring elements were discovered before the existence of the noble gases was known. Why do you think the noble gases were discovered so late? (1.5, 1.6) T/I
72. (a) Julius Lothar Meyer proposed a periodic arrangement of known elements. How did he arrange the elements in his table?
    (b) Suggest a drawback of Lothar Meyer's periodic table. (1.6) K/U T/I C

73. (a) Use what you know about effective nuclear charge to explain why the radius of rubidium, Rb, is larger than the radius of silver, Ag.
(b) Explain the atomic radius of iodine compared to the atomic radii of rubidium and silver. (1.7) K/U T/I A

74. An atom of magnesium, Mg, loses 2 electrons to achieve a stable octet. The first ionization energy for magnesium is 738 kJ/mol, and the second ionization energy for magnesium is 1451 kJ/mol. Why do you think the second ionization energy is greater than the first ionization energy? (1.7) K/U T/I A

Evaluation

75. Compare and contrast chemistry today to what you think chemistry would be like if the International Union of Pure and Applied Chemistry (IUPAC) had not been established. (1.1) K/U T/I C

76. (a) Briefly describe how the work of Niels Bohr contributed to modern atomic theory.
(b) How did the work of Ernest Rutherford contribute to modern atomic theory?
(c) How did the experimental results obtained by both Bohr and Rutherford contribute to the Bohr–Rutherford model of the atom? Include a diagram in your explanation. (1.2) K/U T/I C

77. Suppose 48.2 % of an element is an isotope with a mass of 107 u. Another 32.1 % is an isotope with a mass of 105 u. The remaining proportion of the element is an isotope with a mass of 106. (1.4) K/U T/I A
(a) A student calculates the percentage of the remaining isotope and then calculates the atomic mass as follows:

$$\frac{107 + 105 + 106}{3} = 106$$

Is this an appropriate method of determining the atomic mass of the element? Explain your answer.
(b) Use the procedure outlined in the chapter to calculate the atomic mass of the element.
(c) Explain which is the better procedure, and why.

78. From what you know about the periodic table, would you describe this proposed arrangement of all possible elements as a theory or a law? Explain. (1.5, 1.6) T/I

79. Use what you know about valence shells and ion formation to compare the radii of the following pairs of entities. Explain your answers. (1.7) K/U T/I
(a) a calcium ion, Ca^{2+}, and a calcium atom, Ca
(b) a sulfide ion, S^{2-}, and a sulfur atom, S

80. Locate beryllium, Be, nitrogen, N, and arsenic, As, on the periodic table. For the following pairs of elements, compare the atomic radius, ionic radius, ionization energy, and electron affinity of the two elements. (1.7) T/I
(a) Be and N
(b) N and As

Reflect on Your Learning

81. The term "theory" describes certain concepts in chemistry, such as the atomic theory. The term "law" describes other concepts, such as the law of conservation of mass. Based on what you have learned in this chapter, what are some differences between theories and laws? K/U T/I

82. As a young child, your concept of education was not the same as your current concept of education. Explain how the development of your current concept of education is similar to the development of the atomic theory. K/U T/I C

83. For each of the following descriptions, write the formula of an ion that meets both requirements: K/U
(a) a polyatomic ion and an anion
(b) a polyatomic ion and a cation
(c) an ion with a classical name ending in *-ous* and a valence of +2
(d) an ion with a classical name ending in *-ic* and a valence of +4

84. (a) Should everyone be concerned about radon in the home? Explain your answer.
(b) Do you ever check your home for radon? If so, how do you check it?
(c) Has anything been done to prevent radon from entering your home? If yes, explain what was done. T/I C A

85. We need to know the properties of different elements to determine their uses. Suppose you are designing a building and need to choose metals for structural support in the walls and floors, for electrical wiring, and for plumbing. K/U T/I A
(a) What category of metals is best suited for use in your construction project?
(b) Justify your answer.

86. Element 113 is temporarily named ununtrium, and element 114 is temporarily named ununquadium. The names are derived from the atomic numbers. For example, *un-* stands for "1," *tri-* stands for "3," and *quad-*" stands for "4." If *pent-* means "5," what is the temporary name of element 115? K/U T/I A

87. Imagine that you are clasping a hockey puck in your hands. Anyone who wants to take that puck from you will have to exert energy to do so. K/U T/I A
   (a) How is this situation analogous to removing a valence electron from an atom?
   (b) What is the term applied to the quantity of energy needed to remove a valence electron from an atom?

Research

GO TO NELSON SCIENCE

88. The different forms, or allotropes, of carbon are important to Canada. K/U T/I A
   (a) The Diavik Diamond Mine is located approximately 300 km north of Yellowknife, Northwest Territories. What properties of diamonds make them a desirable form of carbon?
   (b) Extracting iron from its ore requires elemental carbon. Charcoal was widely used for this process in the nineteenth century and into the twentieth century in Canada. Suggest two reasons why charcoal was chosen for this purpose.
   (c) Which province in Canada produces graphite in a commercial quantity?
89. (a) The largest natural deposit of fullerenes on Earth is in Canada. Where is this deposit of fullerenes located?
   (b) Jan Cami, an astronomer at Western University in London, Ontario, found buckminsterfullerene in a rather surprising place. Where did he locate this form of carbon? T/I
90. As atomic theory has developed, scientists have discovered that particles even smaller than protons, neutrons, and electrons exist. One of these particles is called a quark. K/U C
   (a) What is a quark?
   (b) The different types of quarks are called "flavours." What names are assigned to the different flavours of quarks?
   (c) How do quarks relate to protons, neutrons, and electrons?
91. Schools in Canada must control certain compounds used in their laboratories because of safety factors. Investigate each of these compounds. Which ion in each compound might be harmful to students? Why might it be harmful? T/I A
   (a) sodium cyanide, NaCN
   (b) magnesium chromate, MgCrO4
92. Radon in the home is hazardous to human health. T/I C A
   (a) In a building, what is the maximum concentration of radon that the Canadian government considers to be safe?
   (b) What type of radiation does radon emit?
   (c) How can the quantity of radon in a building be limited?
93. One major change between Mendeleev's periodic table and the modern periodic table is the ordering of elements by increasing atomic number instead of increasing atomic mass. K/U T/I
   (a) Who proposed this change to the table?
   (b) What elements were out of order on a table based on atomic mass?
   (c) What led scientists to suspect that the change from atomic mass to atomic number needed to be made?

CHAPTER 2

Chemical Compounds and Bonding

KEY CONCEPTS

After completing this chapter you will be able to

- analyze the properties and effects of useful compounds, evaluate their risks and benefits, and propose ways to minimize any negative effects on societies and the environment
- explore and compare the properties of ionic and molecular compounds, and use these properties to classify compounds as either ionic or molecular
- describe and explain the differences in bonds in ionic compounds and molecular compounds
- use electronegativity values to predict the nature of a bond
- represent ionic and molecular compounds using Lewis structures, molecular models, and structural formulas
- write chemical formulas and IUPAC names for ionic and molecular compounds

How Do Chemical Bonds Determine a Compound's Properties?

Most of the pure substances around us are not elements—they are compounds.

While many compounds occur naturally, many more are synthetic. Teflon is an example of an important synthetic compound. Teflon, made by DuPont, is a brand name for polytetrafluoroethylene (PTFE). It is white, waxy, extremely slippery, and one of the few substances that a gecko cannot stick to!

PTFE and a group of similar compounds have a wide variety of uses, from the non-stick coating on frying pans to non-reactive thread seal tape. A substance similar to PTFE was even used in the roofing materials of the Chinese Olympic stadium (the Bird's Nest) in Beijing. These chemical compounds are known as perfluorochemicals. These compounds are relatively unreactive and do not mix with water or oil.

The use of PTFE as a non-stick coating on cookware means that we can use less fat and oil in our cooking. PTFE can therefore be seen as a beneficial substance.

Technology, however, is sometimes a double-edged sword. Recent studies suggest that, if a Teflon-coated pan is heated to above 350 °C, the coating deteriorates and emits a substance known as perfluorooctanoic acid (PFOA). The United States Environmental Protection Agency describes PFOA as "likely to be carcinogenic to humans." Health Canada states that there is no risk of exposure to PFOA from using cooking utensils and equipment with non-stick coatings. However, non-stick coatings are a risk if an empty pan is left heating on a stove. PFOA is used in the manufacturing process for PTFE. Because PFOA is relatively unreactive, it tends to stay in the human body for a long time. Small quantities of the substance have been detected in humans.

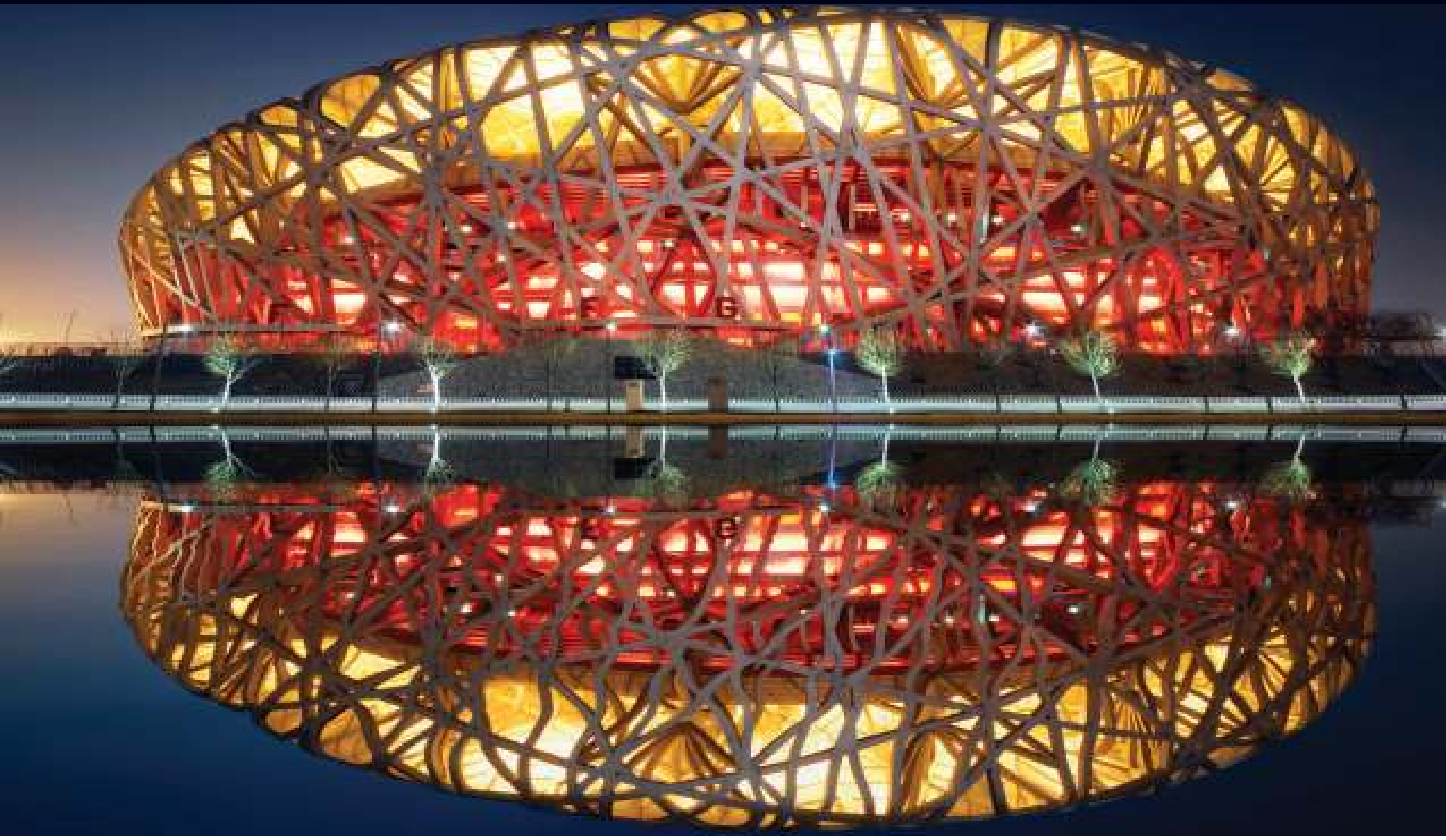
DuPont agrees with Health Canada that Teflon is safe when used properly. The company plans to modify its manufacturing processes, phasing out the use of certain compounds by 2015. According to Health Canada, workers are not exposed to PFOA levels high enough to cause concern. Other organizations, however, disagree. They cite DuPont's plans as evidence that PTFE and PFOA do present a serious health risk.

In this chapter, you will continue to explore the principles of green chemistry and the effects of useful chemicals. You will learn about the bonds within compounds, and how these bonds affect physical and chemical properties.

STARTING **POINTS**

Answer the following questions using your current knowledge. You will have a chance to revisit these questions later, applying concepts and skills from the chapter.

1. What physical or chemical properties of PTFE make it useful or potentially dangerous? Can you make an informed decision on whether to use PTFE products?
2. How might an understanding of the chemical structure of a compound allow scientists to predict its properties?
3. What policies are in place to ensure that consumer products are safe? Who regulates safety standards?
4. Do you think that a company should be obliged to disclose safety concerns? Explain.
5. We are less aware of environmental impacts when the effects occur far away. Do you think that a company should be obliged to disclose the environmental effects of its production facilities in other countries? Explain.



Mini Investigation

Fun with Teflon Tape

Skills: Questioning, Performing, Observing, Analyzing, Communicating

SKILLS HANDBOOK A1.2

In this activity you will examine the physical properties of a product known as "thread seal tape." It is typically made of PTFE and is used to seal joints in plumbing fixtures.

Equipment and Materials: MSDS for thread seal tape; 2 small beakers; round-tipped scissors; pen; ruler; permanent marker; 30 cm length of thread seal tape; 10 mL vegetable oil

1. Before conducting the investigation, read the Material Safety Data Sheet for PTFE thread seal tape.
2. Cut three 5 cm samples of tape. Record any observable physical properties.
3. Pour about 10 mL of vegetable oil into one beaker and 10 mL of water into the other. Drop one of the samples of tape into each beaker. Record your observations.
4. Take the third small sample of the tape and test out additional physical properties. Fold and unfold it. Stretch it widthwise and lengthwise. Record your observations.
5. Place the remaining (15 cm) piece of tape flat on your desk. Using a permanent marker, write a short message or draw a picture on the tape. Your message should be four or five words long, and will be sent to another group.
6. Stretch the tape widthwise until the message is no longer legible. Be careful not to tear the tape.
7. Exchange pieces of tape with another group. Pull the tape lengthwise and read the message.

A. According to the MSDS, what happens to PTFE when it is heated to very high temperatures? T/I

B. What did you observe when you placed the tape in water and oil? State your observations as a property of PTFE. T/I

C. What observed property makes PTFE so useful? T/I

D. Suggest additional uses for this product. A

E. Do you think that this substance would degrade naturally? What implications does this have for disposal? T/I A

2.1 Ionic Compounds

How many compounds do you think there are? Scientists estimate that over 10 million different compounds have been identified—and more are being made every day! In your household alone there is a dazzling variety of compounds (**Figure 1**).

Figure 1 Take a closer look in your cupboards and cabinets and on the shelves at the grocery store. You will find many different types of chemical compounds. Some are safe to consume; others are not.

CAREER **LINK**

Our kidneys remove unwanted substances from our blood and maintain the body's water balance. A nephrologist is a doctor who specializes in the functions and diseases of the kidney. To find out more about a career as a nephrologist,

GO TO NELSON SCIENCE

North Americans consume millions of bottles of sports drinks every day. These drinks consist primarily of water, sugar, and salt, with additives to give a particular taste. They were first developed to help football players gain energy and replenish ions lost in perspiration. Why do athletes, recovering patients, and other thirsty people choose these drinks? The characteristics of a product are a direct result of the properties of the substances it contains.

Some compounds, such as table salt, sugar, and Aspirin (acetylsalicylic acid), are consumable. Of course, we have to be careful how much we consume. Other compounds, such as the ammonia in window cleaners, are corrosive and even toxic.

chemical bond the force of attraction between two atoms or ions

The force of attraction holding two atoms or ions together in a compound is known as a **chemical bond**. Chemical bonds are the invisible "glue" that holds atoms together. How do these bonds form and how do they affect the properties of a compound?

Investigation **2.1.1**

Comparing Ionic and Molecular Compounds (p. 84)
You will have an opportunity to write your own question and answer it by designing and performing your own experiment.

We can divide chemical compounds into two types: ionic compounds and molecular compounds. Each type has distinctive physical and chemical properties. In previous science courses, you have probably learned that an ionic compound is composed of a metal and a non-metal whereas a molecular compound is made up of two or more non-metallic elements. In this chapter we will dig deeper into the chemical bonding within ionic and molecular compounds. We will also explore the effects of these bonds.

The Properties of Ionic Compounds

What do table salt, NaCl(s), limestone, $CaCO_3$(s), and cinnabar, HgS(s), all have in common (**Figure 2**)? They are all ionic compounds. If we investigate the chemical formulas of these substances we see some similarities. They are all composed of a metallic element combined with one or more non-metallic elements. These compounds are all classified as ionic compounds because they share similar chemical structures and exhibit similar properties.

Figure 2 (a) Sodium chloride, (b) calcium carbonate, and (c) mercury(II) sulfide (cinnabar) are all ionic compounds. What properties do they share?

LEARNING **TIP**

Pass the Salt
In chemistry, a salt is described as an ionic compound produced from the neutralization reaction between an acid and a base. Quite often, we confuse this use of "salt" with the common name for a specific chemical compound: sodium chloride, NaCl. Be alert for these two uses of the same name.

What similarities can you observe among the ionic compounds in Figure 2? To start, all three substances are solids at ambient (normal) temperatures and they all have definite geometries. Ionic compounds have several other properties in common:

- They are hard and brittle.
- They have relatively high melting and boiling points. For example, sodium chloride melts at 801 °C.
- They conduct electricity as molten liquids but not as solids.
- They conduct electricity when dissolved in water.

Pure water is a poor conductor of electricity (**Figure 3(a)**). If we add table salt, however, we notice that the salt dissolves fairly easily, forming a salt solution (**Figure 3(b)**). This solution now conducts electricity (**Figure 3(c)**). A substance that forms a solution that conducts electricity is called an **electrolyte**. Sodium chloride is therefore an electrolyte. Sports drinks contain electrolytes as well as sugar and flavouring. Electrolytes are important in the body because they help carry electrical impulses throughout the body. As you exercise and sweat, you lose electrolytes.

How can we explain our observations of the properties of ionic compounds? What theories and models will help us to account for the properties? For a theory to be successful, it must explain as many observations as possible and predict future observations.

electrolyte a compound that dissolves in water, producing a solution that conducts electricity

WEB **LINK**

Gatorade was invented in 1965 by Dr. J. Robert Cade, a doctor who specialized in the function of the kidneys. To learn more about the invention of electrolyte drinks,

GO TO NELSON SCIENCE

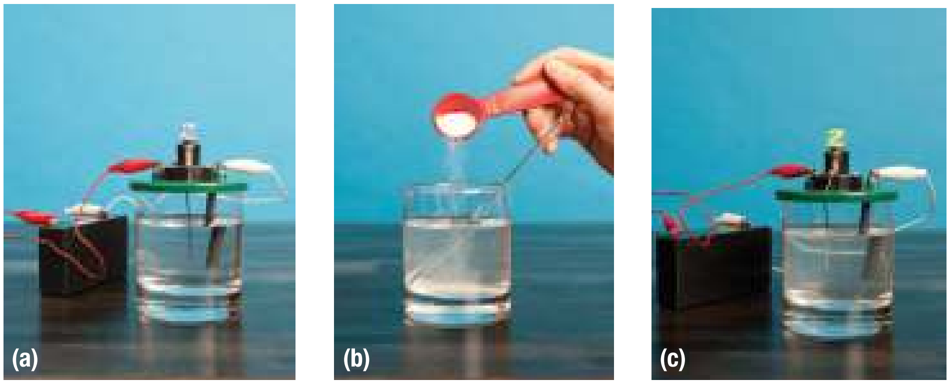


Figure 3 (a) Pure water is a poor conductor of electricity. (b) Sodium chloride dissolves in water to form a solution. (c) The resulting solution can conduct electricity extremely well. Sodium chloride is therefore an electrolyte.

The Formation of Ionic Compounds

We will start by considering the formation of one of the most familiar ionic compounds: sodium chloride (**Figure 4**).

Figure 4 Sodium metal and chlorine gas react violently to form sodium chloride (table salt).

In Chapter 1 you learned that neutral atoms can gain or lose electrons to become ions. Ions have either positive or negative charges. When a neutral atom loses one or more electrons it becomes a cation: a positively charged ion. Metals, such as sodium and potassium, tend to lose electrons to form cations. We can represent the formation of a sodium ion, Na^+, using Lewis symbols as follows:

$$\dot{\mathrm{Na}} \rightarrow [\mathrm{Na}]^+ + e^-$$

If a neutral atom gains one or more electrons, it becomes an anion: a negatively charged ion. Non-metals, such as fluorine and oxygen, tend to gain electrons to form anions. We represent the formation of a chloride ion, Cl^-, with Lewis symbols as follows:

$$\cdot\underset{\cdot\cdot}{\ddot{\mathrm{Cl}}}: + e^- \rightarrow [:\underset{\cdot\cdot}{\ddot{\mathrm{Cl}}}:]^-$$

An ionic bond forms when the non-metal atom removes an electron from the metal atom:

$$\dot{\mathrm{Na}} + \cdot\underset{\cdot\cdot}{\ddot{\mathrm{Cl}}}: \rightarrow [\mathrm{Na}]^+[:\underset{\cdot\cdot}{\ddot{\mathrm{Cl}}}:]^-$$

LEARNING **TIP**

Energy in Chemical Equations
You might notice that sometimes chemical equations include the term "energy" on one side or the other. At other times this term is omitted. Generally, "energy" is only included in the equation if it is relevant in the context. Here, to keep the focus on the formation of ions, we have not shown the energy term.

Electrons are transferred from the sodium atoms to the chlorine atoms, forming positive sodium ions and negative chlorine ions. You may recall the concept of electrostatic force from a previous science course. Electrostatic force is the force that attracts a positively charged object to a negatively charged one. This force holds together the charged sodium and chloride ions. The bond between a positive ion and a negative ion is an **ionic bond**.

ionic bond the electrostatic force of attraction between a positive ion and a negative ion; a type of chemical bond

What happens when a different metal, such as magnesium, forms an ionic compound with chlorine? Each magnesium atom loses 2 electrons to form a magnesium ion, Mg^{2+}, with a complete valence shell. These 2 electrons are taken by 2 different chlorine atoms that, in turn, become 2 chloride ions, Cl^-. The resulting compound, magnesium chloride, has the chemical formula $MgCl_2$.

$$\dot{\mathrm{Mg}}\cdot + 2\,\cdot\ddot{\underset{\cdot\cdot}{\mathrm{Cl}}}\colon \rightarrow [\colon\ddot{\underset{\cdot\cdot}{\mathrm{Cl}}}\colon]^- \,[\mathrm{Mg}]^{2+}\,[\colon\ddot{\underset{\cdot\cdot}{\mathrm{Cl}}}\colon]^-$$

Tutorial 1 Representing the Formation of Ionic Compounds Using Lewis Symbols

The following sample problem provides useful strategies for representing the formation of bonds within ionic compounds.

Sample Problem 1: The Formation of Ionic Compounds

Draw the Lewis symbols to show the formation of bonds within sodium oxide. Write the chemical formula for the compound produced.

Step 1. Write a Lewis symbol for each atom of the combining elements.

$$\dot{\mathrm{Na}} \qquad \cdot\ddot{\underset{\cdot}{\mathrm{O}}}\colon$$

Step 2. Consider how many electrons the atoms are likely to gain or lose to achieve full valence shells and decide on the ratio of ions required to produce a neutral ionic compound.

Two Na atoms are needed for every O atom.

Step 3. Draw each ion as a Lewis symbol. Place the ions in square brackets with their respective charges as superscripts. Include full octets for anions.

$$\dot{\mathrm{Na}} + \cdot\ddot{\underset{\cdot}{\mathrm{O}}}\colon + \dot{\mathrm{Na}} \rightarrow [\mathrm{Na}]^+[\colon\ddot{\underset{\cdot\cdot}{\mathrm{O}}}\colon]^{2-}[\mathrm{Na}]^+$$

Step 4. Write the chemical formula, incorporating the ions you have drawn.

The chemical formula for sodium oxide is Na_2O.

Practice

1. Draw Lewis symbols to show the formation of bonds within the following ionic compounds. In addition, write the chemical formula for the compound produced. K/U C
   (a) potassium chloride
   (b) calcium chloride
   (c) calcium oxide
   (d) potassium oxide

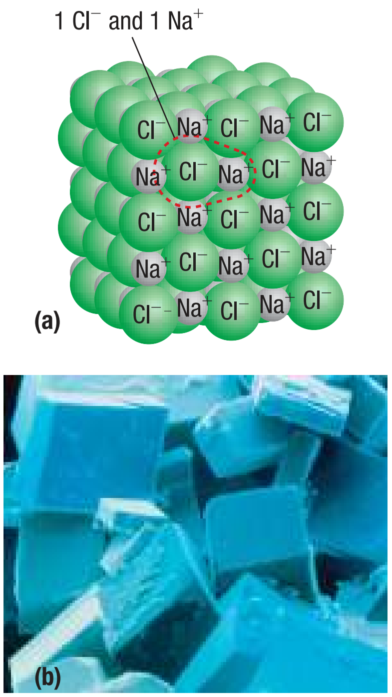


Figure 5 (a) Sodium chloride, NaCl(s), is made up of equal numbers of positive and negative ions. (b) Sodium chloride crystals are cubic in shape, as this image, taken with a scanning electron microscope, shows.

The Structure of Ionic Compounds

When sodium chloride forms, large numbers of sodium ions and chloride ions arrange themselves into crystals. The rigid arrangement of ions is called a crystal lattice structure. **Figure 5(a)** shows a model of the three-dimensional structure of sodium chloride. Notice that it is cube-shaped. This model is supported by our observation that grains of sodium chloride (table salt) are little cubes (**Figure 5(b)**).

Not all ionic crystals exist as cubic structures. Crystals of cinnabar (mercury(II) sulfide), for example, do not contain right angles. Why are crystals different shapes? We could hypothesize that it is because of the different sizes of the ions, and the strength with which they attract each other. As an analogy, imagine packing apples and oranges into a cubic box. You can easily pack the similar-sized fruit into a cube. In the same way, ions of similar size can be packed in a cubic shape. Now imagine packing grapefruits and limes into a box. Due to their different sizes, you might need a box of a different shape. Different-sized ions pack to form different crystal shapes.

In a tiny grain of sodium chloride, there are billions and billions of sodium ions and chloride ions bonded together. It is important to realize that there is no single unit made up of a single sodium ion and a single chloride ion: there are no molecules of sodium chloride, NaCl. We do, however, define the smallest repeating unit in an ionic crystal structure as the **formula unit** (**Figure 6**). The formula unit for sodium chloride is NaCl; the formula unit for mercury(II) sulfide is HgS; the formula unit for calcium carbonate is $CaCO_3$. An **ionic compound**, therefore, is not made up of one positive ion and one negative ion. Rather, an ionic compound is composed of huge numbers of positive and negative ions in a fixed ratio.

formula unit the smallest repeating unit in an ionic crystal

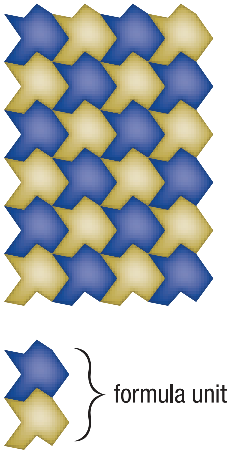


Figure 6 The formula unit is the smallest repeating unit for a repeating pattern.

ionic compound a pure substance composed of positively charged ions and negatively charged ions in a fixed ratio

Explaining the Properties of Ionic Compounds

We can use our model of the structure of ionic compounds to explain their properties.

- Ionic compounds have relatively high melting points because their ions are held together by strong electrostatic forces (ionic bonds).
- Ionic compounds are hard because their bonds resist being "stretched."
- A piece of sodium chloride is easily cracked or fractured. This is because, when an outside force strikes the crystal, the crystal lattice structure is offset. Suddenly, positively charged ions are side by side with other positively charged ions. A repulsive force quickly develops between the like charges, and the crystal breaks.
- Ionic compounds are electrolytes because, when an ionic crystal is placed in water, water molecules surround each ion and separate it from the crystal. The crystal breaks up or dissolves, releasing free-floating ions into the solution. The ions are able to move, and thus to carry electric charges, through the water (**Figure 7**). This is what happens when electricity passes through a solution.

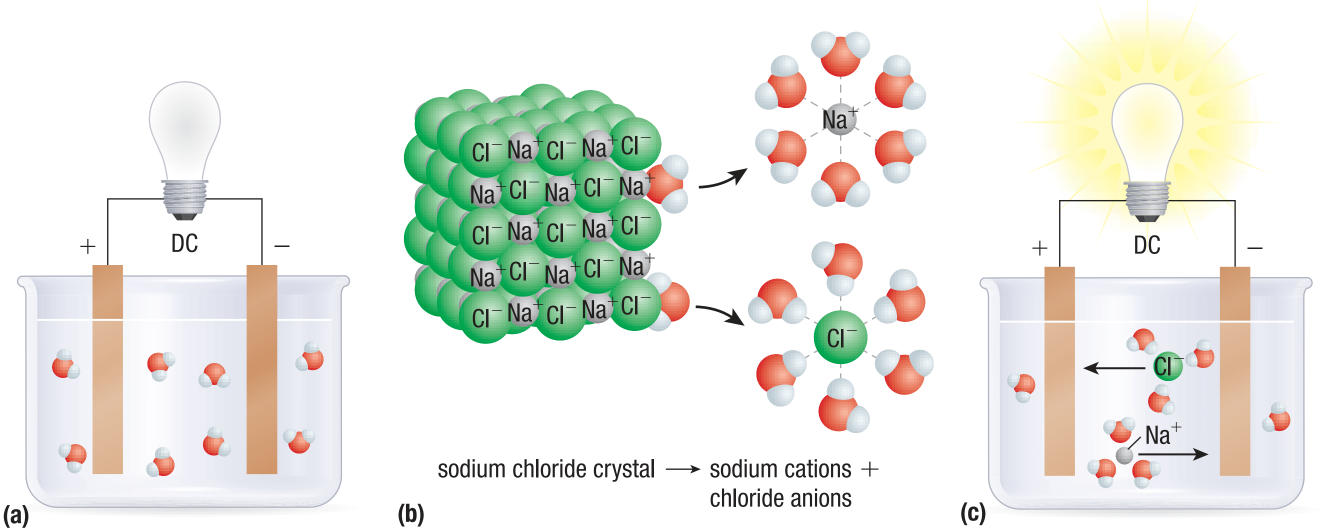


Figure 7 (a) Neutral water molecules (shown here as red and white models) do not carry charge. Therefore, they are poor conductors of electricity. (b) Sodium chloride dissolves in water, releasing sodium ions, Na^+, and chloride ions, Cl^-. (c) The ions allow electricity to flow through the salt solution.

LEARNING TIP

Improving Definitions

You now have a definition for ionic compounds that is slightly different from the one you were given in earlier grades. As you learn more about a concept, you will continue to enhance and improve your definition.

Remember that pure water is a poor conductor of electricity. Tap water, lake water, and seawater, however, are good conductors because they contain ions from a variety of sources, such as minerals. (Remember that water is not an ionic substance, as we will explore in Section 3.5.)

2.1 Summary

- There are two main types of chemical compounds: ionic and molecular.
- Ionic compounds are made up of positive ions (cations) and negative ions (anions) held together with ionic bonds.
- Ionic compounds have distinguishing physical properties. These properties can be explained by the nature of ionic bonds.
- The formation of ionic compounds comes about by the transfer of electrons to form ions, and the subsequent attraction of the oppositely charged ions.
- Lewis symbols can be used to represent the formation of ionic bonds.

2.1 Questions

1. List three distinctive physical properties of ionic compounds. K/U
2. Define the following terms: electrolyte, ionic bond, and formula unit. K/U
3. Draw the Lewis symbols to show the formation of bonds within magnesium oxide. Write the chemical formula for the compound produced. K/U T/I C
4. Pure water is a poor conductor of electricity. Tap water, however, is a fair conductor of electricity. Explain this observation. K/U
5. The melting points for four different ionic compounds are given in **Table 1**. Use this information to compare the strength of the ionic bond in each of the compounds. Suggest a reason why some ionic bonds might be stronger than others. T/I

Table 1 Melting Points of Selected Ionic Compounds

Chemical name	Chemical formula	Melting point (°C)
sodium chloride	NaCl	801
lithium chloride	LiCl	605
potassium chloride	KCl	770
magnesium oxide	MgO	2852

6. In general, ionic compounds are hard solids at ambient temperatures but they shatter fairly easily. How do these two properties support the theory of ionic bonding? K/U
7. Use Lewis symbols to illustrate the formation of the following ionic compounds: K/U
   (a) lithium chloride
   (b) magnesium chloride
   (c) sodium sulfide
   (d) aluminum oxide
8. Element A has 2 valence electrons. Element B has 7 valence electrons. K/U T/I C
   (a) Classify each of these elements as a metal or a non-metal.
   (b) Use Lewis symbols to show how A and B would form a compound.
   (c) What kind of compound is formed?
   (d) Write the chemical formula for this compound.
9. What do ionic compounds, houndstooth check fabrics, and the work of M.C. Escher have in common? Research to find out. T/I A
10. As you are probably aware, various ionic compounds (mostly sodium chloride and calcium chloride) are used as road salt in countries that have cold winters. These compounds lower the freezing point of water to slow down ice formation. This helps make roads safer in the winter. These salts, however, promote corrosion of cars and contribute to various environmental problems. Research these two road salts and possible alternatives that are less harmful. T/I A

GO TO NELSON SCIENCE

Molecular Elements and Compounds

2.2

You and a candle have something in common. You both consume oxygen and fuel. The candle burns paraffin wax and your cells "burn" sugar. Also, you both produce water, $H_2O(l)$, and carbon dioxide, $CO_2(g)$. These are all molecular substances. Oxygen is a molecular element; the rest are molecular compounds.

A **molecular element** consists of molecules containing atoms of only one type of element. Nitrogen, $N_2(g)$, and bromine, $Br_2(l)$, are **diatomic** elements: each molecule is made up of 2 atoms. Ozone, $O_3(g)$, and sulfur, $S_8(s)$, are other examples of molecular elements. All the molecules in a particular molecular element are identical.

A **molecular compound** is also made up of molecules, each with the same arrangement of specific atoms. Each molecule in a molecular compound is made up of at least 2 different elements—sometimes many more. All the molecules in a particular compound are identical, but different from the molecules in another compound. Molecular compounds generally consist of only non-metallic elements, although some important molecular compounds involve metals and metalloids.

molecular element a pure substance composed of molecules made up of two or more atoms of the same element

diatomic made up of two atoms

molecular compound a pure substance composed of molecules made up of two or more non-metallic elements

What do substances look like at the atomic level? A model would enable us to visualize how compounds are "constructed" at this scale. Indeed, understanding the shape and structure of a molecule is critical as we explore its properties. For example, carvone is a molecular compound found in the oils of many plants. Carvone has a molecular formula of $C_{10}H_{14}O$, but has two different possible structures that happen to be mirror images of each other. One form of carvone is found in spearmint plants while its mirror image is found in caraway plants (**Figure 1**). These two molecules exhibit different physical properties—especially their smell!

WEB LINK

To see the two different structures of carvone in 3-D animations,

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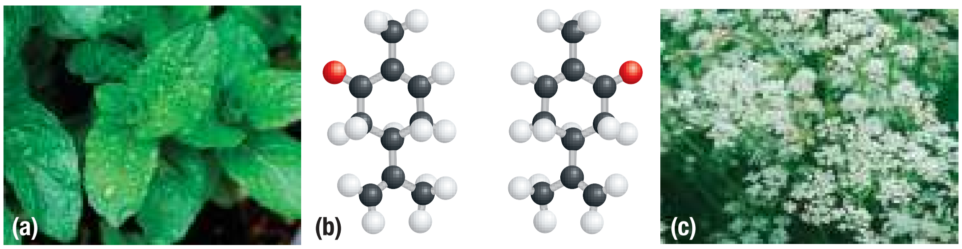

Figure 1 Molecular structure plays a key role in determining compounds' properties. (a) A spearmint plant (b) Two structures of carvone, $C_{10}H_{14}O$ (c) A caraway plant

How can two molecules that are composed of the same atoms have such different odours? Biochemists propose that, for us to detect a specific smell, individual scent molecules have to fit into a scent receptor in the nose like a key fits into a lock. A mirror-image key cannot fit into the same lock. Look at a key and you will see why! According to this theory, the two different carvone molecules fit into different scent receptors because of their slight differences in structure. Hence we detect different smells.

CAREER LINK

Biochemists study the formation and reactions of substances within organisms. To find out more about the work of a biochemist,

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As we investigate the properties of molecular compounds, it is important to be able to represent their structures clearly to get an idea of their shapes. Their structures will, in turn, tell us a great deal about their properties.

The majority of known compounds are molecular compounds. **Table 1** lists some examples of simple and complex molecular compounds. What properties do these molecular compounds have in common? What distinguishes molecular compounds from ionic compounds? As we investigate the properties of molecular compounds, we will propose a model for how these compounds form. What keeps the atoms together in each molecule? How can diagrams and structures help us visualize individual molecules?

Table 1 Some Molecular Compounds

Name	Molecular formula
Simple molecules	
methane	$CH_4(g)$
ammonia	$NH_3(g)$
hydrogen peroxide	$H_2O_2(l)$
Complex molecules	
sucrose (a sugar)	$C_{12}H_{22}O_{11}(s)$
oleic acid (in corn oil)	$C_{18}H_{34}O_2(l)$

The Properties of Molecular Substances

Molecular substances exist as gases, liquids, and solids at ambient temperatures (**Figure 2**). These substances can have a variety of textures: paraffin wax is waxy and soft while sugar is hard. Wax is quite flexible and malleable while sugar is brittle. Molecular elements are usually gases, due to their low boiling points. Molecular compounds also have relatively low melting and boiling points. As well, many molecular substances do not dissolve readily in water.

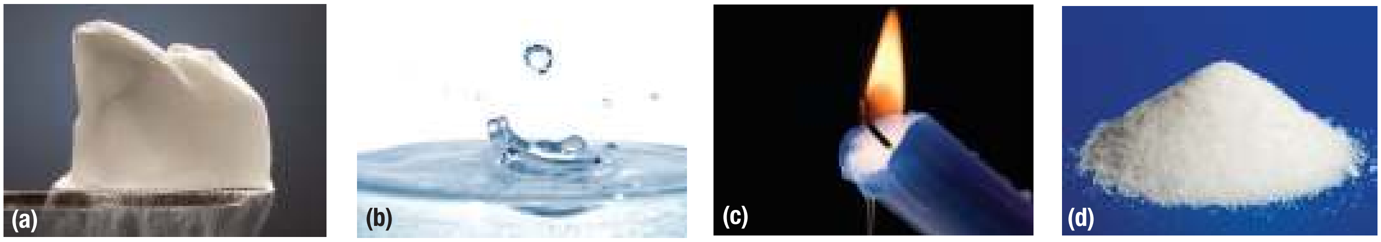

Figure 2 (a) Carbon dioxide, (b) water, (c) paraffin wax, and (d) sugar are examples of molecular compounds.

Covalent Bonding within Molecules

The physical properties of molecular substances depend on their structure. We will now look at the structures of molecules, and how they are formed.

MOLECULAR ELEMENTS

We will start with a quick review of the formation of the simplest of these molecules: hydrogen, H_2. A hydrogen atom consists of just 1 proton and 1 electron. When 2 hydrogen atoms collide, the proton of each atom attracts the electron of the other atom. This is a bit like a tug-of-war competition for electrons that neither atom wins. As a result, the 2 atoms share their electrons and form a hydrogen molecule, H_2. The attractive force or bond that results from the sharing of an electron pair is called a **covalent bond** (**Figure 3**). In this way, each hydrogen atom achieves a full valence shell. As you know from Chapter 1, a full valence shell is a stable electron arrangement similar to that of the nearest noble gas. The nearest noble gas to hydrogen on the periodic table is helium, which has 2 electrons in its valence shell.

covalent bond the bond that results from the sharing of a pair of electrons by two atoms

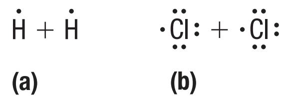

Figure 3 A covalent bond forms when 2 atoms share a pair of valence electrons.

How does another diatomic element, chlorine, form molecules? When 2 chlorine atoms join, they form a similar bond (Figure 3(b)). The protons of each chlorine atom attract the valence electrons of the other and a single covalent bond is formed. The result is a Cl_2 molecule. In this manner, each chlorine atom achieves the stable electron arrangement of the noble gas argon.

Look again at Figure 3. Notice that each unbonded hydrogen and chlorine atom has 1 unpaired electron in its valence shell. This unpaired electron is sometimes called a **bonding electron**. It is available to form a covalent bond with another atom.

bonding electron an electron, in the valence shell of an atom, that is available to form a covalent bond with another atom

The number of bonding electrons in an atom affects the number of bonds it can form. This number is known as the **bonding capacity**. **Table 2** shows the bonding capacities of some common elements. These elements are representative of the other elements in these groups.

bonding capacity the number of covalent bonds that an atom can form

LEARNING **TIP**

Co-valent
Note that the prefix *co-* typically means "with" or "shared" in words such as cooperate and cohabitate. The term "covalent" implies "sharing its valence" or "sharing electrons."

Table 2 Bonding Capacities and Lewis Symbols of Atoms Commonly Forming Covalent Bonds

Group	1	14	15	16	17
Number of valence electrons	1	4	5	6	7
Bonding capacity	1	4	3	2	1
Lewis symbols	H·	·C·	·N·	·O:	·F:

When we use Lewis symbols to represent the bonds within a molecule, we produce a **Lewis structure**. Lewis himself used only dots to represent all the electrons. However, chemists now typically use a single dash to replace each pair of bonding electrons.

Lewis structure a representation of covalent bonding based on Lewis symbols; a model in which shared electron pairs are shown as lines and unshared electrons are shown as dots

A single dash represents a single covalent bond. A double dash represents a double covalent bond. A double covalent bond is formed when two atoms share two pairs of electrons. A triple dash represents a triple covalent bond, which is formed when two atoms share three pairs of electrons. Each pair of electrons that is not involved in bonding is called a **lone pair**. Remember that Lewis structures, while useful, are two-dimensional representations of three-dimensional objects.

When two oxygen atoms combine to form molecular oxygen, O_2, the oxygen atoms share two pairs of valence electrons in order to satisfy the octet rule. When atoms share two pairs of electrons, they form a double covalent bond (**Figure 4(a)**).

When two nitrogen atoms combine to form molecular nitrogen, N_2, the nitrogen atoms share three pairs of valence electrons in order to satisfy the octet rule. When atoms share three pairs of electrons, they form a triple covalent bond (**Figure 4(b)**).

Figure 4 Lewis structures show lone pairs of electrons as dots and each shared pair as a line. (a) The double line in an oxygen molecule represents two shared pairs: a double covalent bond. Each atom still has two unshared pairs of valence electrons: two lone pairs. (b) The triple line in a nitrogen molecule represents three shared pairs: a triple covalent bond. Each atom has one lone pair.

So far we have considered diatomic molecules of the same element. As we look at molecular compounds with different types of atoms and with a variety of numbers of atoms, we need to devise some tips and strategies for drawing Lewis structures.

MOLECULAR COMPOUNDS

Molecular compounds form in much the same way as molecular elements. The difference is that the atoms are not from the same element. **Figure 5** shows how a molecule of hydrogen chloride forms from an atom of hydrogen and an atom of chlorine. Note how they share a pair of electrons, forming a single covalent bond. In this way, both atoms fulfill the octet rule.

Figure 5 A single covalent bond forms when two different atoms share a pair of valence electrons.

Molecules are not limited to two atoms each. You are already very familiar with some three-atom molecules: water, H_2O, and carbon dioxide, CO_2 (**Figure 6**). The atoms in these molecules arrange themselves so they each have a complete valence shell. Each molecule has two bonds: water forms two single covalent bonds and carbon dioxide forms two double covalent bonds.

Figure 6 Lewis structures of (a) water and (b) carbon dioxide show how different elements can share valence electrons to form stable molecules.

You now know how to represent a molecular substance using a Lewis structure. This type of diagram shows both the lone pairs of electrons surrounding each atom and the covalent bonds (single, double, or triple) between the atoms.

Chemists often simplify their diagrams even further by only including the covalent bonds in their diagrams. This representation, without the lone pairs, is called a **structural formula** (**Figure 7**).

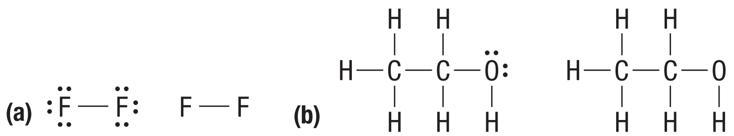

Figure 7 Compare the Lewis structures and the structural formulas of (a) a molecule of fluorine, F_2, and (b) a molecule of ethanol, C_2H_5OH.

lone pair a pair of electrons that is not involved in covalent bonding

WEB LINK

To find out more about the International Union of Physical and Applied Chemistry and its importance to the scientific world,

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LEARNING TIP

Lewis Symbols and Lewis Structures

Both Lewis symbols and Lewis structures show the chemical symbols of elements and the valence electrons of each atom. So what is the difference? In Section 1.5 you learned that a Lewis symbol represents an atom or ion of an element. A Lewis structure, however, shows how bonds form between two or more atoms to form a molecule. The lines represent covalent bonds.

structural formula a representation of the number, types, and arrangement of atoms in a molecule, with dashes representing covalent bonds

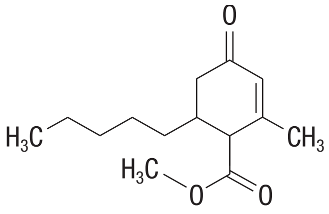

Figure 8 Many of the carbon and hydrogen symbols have been omitted to make this structure easier to interpret.

WEB LINK

Chemists use powerful computer programs to draw structures for complex molecules. Some of these programs are available online. To try a demonstration copy of a molecular modelling program,

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Figure 9 This ball-and-stick model of methanal, CH_2O, shows the carbon atom as a black sphere, the oxygen atom as a red sphere, and the hydrogen atoms as white spheres. The connecting sticks represent single or double covalent bonds.

Water and carbon dioxide are relatively simple molecular compounds. Compounds such as paraffin wax, sugar, and lactic acid, however, are made of much larger and more complex molecules. Biological compounds are often hugely complex. One feature they all have in common is that they contain a lot of carbon atoms. Representing them by Lewis structures or even structural formulas would be time-consuming and cluttered. Biochemists have developed yet another style of diagram: they leave out all the carbon and hydrogen atoms and show just the bonds. Because of the bonding pattern of carbon, the bonds almost always change direction at a carbon atom (**Figure 8**). You can therefore assume that there is a carbon atom at each angle or intersection of the bonds.

Large molecular compounds are usually liquids or solids. Chemists have developed many molecular compounds with very useful properties, such as plastics, textiles, explosives, medicines, fats, and coatings.

Methanal, CH_2O, is a very important building block for the synthesis of larger molecular substances (**Figure 9**). In 2004, methanal was declared a known cancer-causing agent (a carcinogen). There are significant health concerns associated with exposure to methanal in the workplace. As well, there are concerns about the release of methanal gas from many synthetic products such as carpets and furniture.

Representing Molecular Structures

You have seen a few Lewis structures and structural formulas. Now you will learn how to draw these structures.

Tutorial 1 Drawing Lewis Structures

CASE 1: MOLECULAR COMPOUNDS

We can draw Lewis structures for molecular compounds by considering bonding capacities and following some general rules. The following sample problems illustrate these rules. Note that the central atom in the structure is generally the element for which there is a single atom. The central atom also tends to be (but is not always) the element with the highest bonding capacity.

Sample Problem 1: Drawing a Lewis Structure for a Molecular Compound

Draw a Lewis structure and a structural formula for a molecule of methanal, CH_2O.

Step 1. Arrange the symbols of the elements in the compound so that the element with the largest bonding capacity is in the centre, surrounded by the elements with smaller bonding capacities. Spread the surrounding atoms as far apart as possible around the central atom.

```
   O

   C
H     H
```

Step 2. Count all the valence electrons for all the atoms in the molecule. This total represents the total number of electrons (dots) you can use in your Lewis structure. Record this number.

C atom: 4 e^-

O atom: 6 e^-

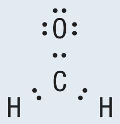
2 H atoms: 2 (1 e^-) = 2 e^-

Total: 12 e^-

Step 3. Place 2 electrons (a single covalent bond) between the central atom and each of the surrounding atoms.

```
   O
   ··
  :C:
H     H
```

Step 4. Place lone pairs of electrons around each of the surrounding atoms (except hydrogen) to satisfy the octet rule for these atoms.



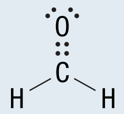
Step 5. Determine how many electrons are still available by subtracting the number of electrons you have used so far from the total number of valence electrons. Record your new value.

All 12 electrons have been used.

Step 6. Place the remaining electrons on the central atom(s) in pairs.

None are remaining.

Step 7. If the central atom does not have a full octet, move lone pairs from the surrounding atoms into a bonding position between the central and surrounding atoms.

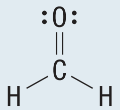


The central atom has a full octet.

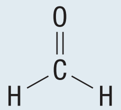
Step 8. If the surrounding atoms have complete octets and there are electrons remaining, add these electrons as lone pairs onto the central atom. Check the finished structure. All atoms (except hydrogen) should have a complete octet, counting shared and lone pairs.

All octets are complete.

Step 9. Replace the shared pairs of electrons with dashes to represent the covalent bonds. Use double or triple dashes for double or triple bonds.



Step 10. To convert the Lewis structure to a structural formula, remove the dots representing lone pairs.



Practice

1. Draw a Lewis structure and structural formula for each of the following molecules: T/I C

(a) F_2
(b) N_2
(c) PH_3
(d) CO_2
(e) SiH_4
(f) C_2H_6
(g) C_2H_4
(h) CO
(i) HCN
(j) HNO_3

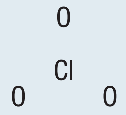
CASE 2: POLYATOMIC IONS

The process for drawing a Lewis structure for a polyatomic ion is similar to that for drawing a molecule. Use the same strategies introduced for molecular compounds in Case 1. This time, you also have to consider the overall charge on the structure. In Step 2, after adding up all the valence electrons around the various atoms, add an extra electron for each unit of negative charge in the polyatomic ion or subtract one electron for each unit of positive charge in the ion. Draw square brackets around the structure of the polyatomic ion and indicate the charge of the ion outside the brackets.

Sample Problem 2: Drawing a Lewis Structure for a Polyatomic Ion

Draw the Lewis structure and structural formula for the polyatomic chlorate ion, ClO_3^-.

Step 1. Arrange the symbols of the elements in the compound so that there is a single atom in the centre, surrounded by the other atoms. Spread the surrounding atoms as far apart as possible around the central atom.



Step 2. Count all the valence electrons for all the atoms in the molecule. This total represents the total number of electrons (dots) you can use in your Lewis structure. Record this number.

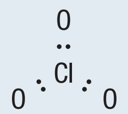
Cl atom: 7 e^-

3 O atoms: $3 \times 6\ e^- = 18\ e^-$

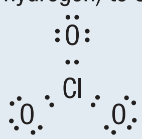
-1 charge of ClO_3^-: 1 e^-

Total: 26 e^-

Step 3. Place two electrons (a single covalent bond) between the central atom and each of the surrounding atoms.



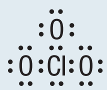
Step 4. Place lone pairs of electrons around each of the surrounding atoms (except hydrogen) to satisfy the octet rule for these atoms.



Step 5. Determine how many electrons are still available by subtracting the number of electrons you have used so far from the total number of valence electrons. Record your new value.

$(26\ e^-) - (24\ e^-) = 2\ e^-$

Step 6. Place the remaining electrons on the central atom in pairs.



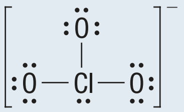
Step 7. If the central atom does not have a full octet, move lone pairs from the surrounding atoms into a bonding position between the central and surrounding atoms, forming double or triple bonds.

In this case, all atoms have full octets, so no lone pairs have to be moved. There are no double or triple bonds in this structure.

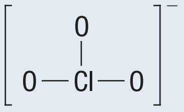
Step 8. If the surrounding atoms have complete octets and there are electrons remaining, add these electrons as lone pairs onto the central atom. Check the finished structure. All atoms (except hydrogen) should have a complete octet, counting shared and lone pairs.

In this case, all octets are complete.

Step 9. Replace the shared pairs of electrons with dashes to represent covalent bonds. Put square brackets around the structure and write the charge outside the brackets.



Step 10. To convert the Lewis structure to a structural formula, remove the dots representing lone pairs.



Practice

2. Draw the Lewis structure and structural formula for the following polyatomic ions:
   (a) OH^- (b) CN^- (c) NH_4^+ (d) PO_4^{3-}

Mini Investigation

Building Molecular Models

Skills: Performing, Observing, Communicating

SKILLS HANDBOOK A2.4

Structural diagrams are very useful in helping us understand the two-dimensional structure of a molecule. Molecular model kits allow us to extend structural diagrams into three dimensions. Each sphere in a molecular model kit represents a specific type of atom. The number of connection sites represents that atom's bonding capacity. For example, hydrogen atoms are usually represented by white spheres with one connection site. Carbon atoms are usually represented by black spheres with four connection sites. In this activity, you will build models of molecules to help you understand how some elements form molecular compounds.

Equipment and Materials: molecular model kit

1. Copy **Table 3** into your notebook, leaving enough space for your observations.
2. Record the colour code for the atoms in your kit. For example, the colour code might be white for hydrogen, red for oxygen, green for the halogens, and black for carbon.
3. Draw structural formulas for the substances listed in the first column.
4. Select the appropriate "atoms" for each molecule.
5. Build the molecules using the connectors to represent covalent bonds.

A. Complete your table. K/U T/I C

B. Explain how building molecular models might help you envision the shape and formation of molecules. T/I

C. Suggest examples of molecules that would have similar structural formulas and molecular shapes to the substances in your table. T/I

D. Propanone (sometimes called acetone) and propenol have the same chemical formula—C_3H_6O—but their molecular structures are quite different. As a result, the physical and chemical properties of these two substances are quite different. Research the structures, properties, and uses of these two substances. T/I A

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Table 3 Selected Molecular Substances

Chemical name	Chemical formula	Structural formula of molecule	Number of single covalent bonds	Number of double or triple covalent bonds	Number of lone pairs around each atom	Sketch of molecular shape
nitrogen	N_2					
water	H_2O					
carbon dioxide	CO_2					
propane	C_3H_8					
ammonia	NH_3					
silicon tetrachloride	$SiCl_4$					
ethanol	C_2H_5OH					

Benzene's Mysterious Structure

Benzene is a widely used but carcinogenic compound. It was first discovered in 1825 by the famous English scientist Michael Faraday. Although scientists knew the chemical formula of the compound, C_6H_6, its structure remained a mystery for many years. The story of its discovery is fascinating and somewhat controversial.

In one version, Friedrich August Kekule, a German chemist, came upon the idea for a ring structure from a dream about a snake holding its own tail (**Figure 10**). As he wrote, "My mental eye . . . could now distinguish larger structures of manifold conformations; long rows, sometimes more closely fitted together; all twisting and turning in snake-like motion. But look! What was that? One of the snakes had seized hold of its own tail. . . ." A less memorable version credits one of Kekule's professors, Johann Loschmidt, for originally proposing the ring structure.

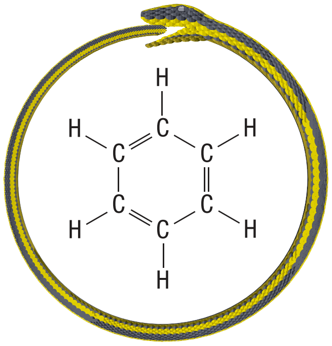

Figure 10 Kekule suggested that his dream of the snake biting its tail gave him the idea of a carbon ring.

Exceptions to the Octet Rule

Nitrogen dioxide, NO_2, is one of several nitrogen oxides. It is a toxic, reddish-brown gas that has a distinctive odour. The industrial manufacture of nitric acid produces large quantities of nitrogen dioxide. If it escapes into the atmosphere it is a serious air pollutant. Its presence contributes to respiratory problems caused by smog (**Figure 11**). What is the structure of nitrogen dioxide? No matter how you draw it, the structure always has 1 extra electron. That extra electron is what makes nitrogen dioxide so toxic.

LEARNING **TIP**

Exceptions to the Rule

As we develop a better understanding of chemical concepts, our models and rules improve. The octet rule is very helpful as we start to draw and visualize molecules. There are exceptions, however, that we cannot explain at this stage. Indeed, there are oddities and exceptions even with the best present-day models. Scientific knowledge is always evolving and there is definitely room for future discoveries.

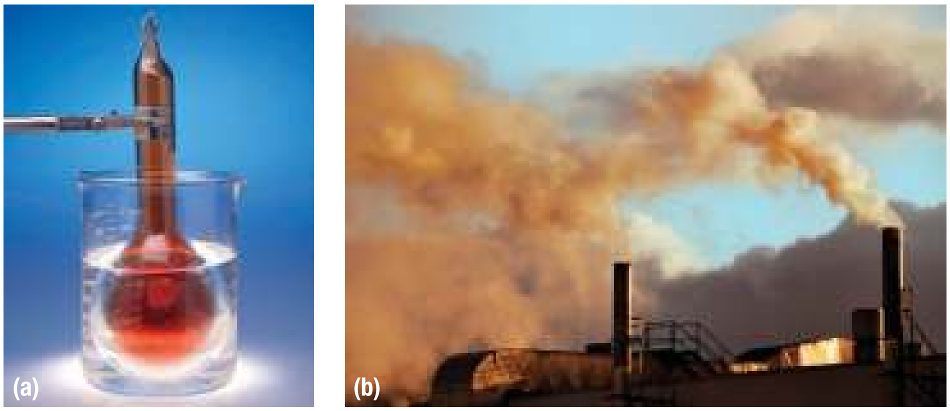

Figure 11 (a) Nitrogen dioxide gas causes (b) the familiar brownish colour of visible air pollution.

The nitrogen dioxide example indicates that the octet rule is not perfect: there are exceptions to the rule. It seems that some elements can form molecular compounds with an overfilled or underfilled valence shell. Subsequent chemistry courses will provide an explanation of these exceptions.

Explaining the Properties of Molecular Compounds

How does our model of molecular bonding help to explain the physical and chemical properties of molecular compounds? The physical properties of molecular compounds depend on the forces of attraction between individual molecules. These forces, in turn, depend on the covalent bonds between the atoms and how they are arranged. The rest of this chapter and unit will help you understand how the nature of the covalent bond and the shapes of molecules affect the properties of molecular compounds.

UNIT TASK **BOOKMARK**

You will want to consider the properties of the substances you choose for your green product as you work on your Unit Task (page 134).

Summary of Ionic versus Molecular Compounds

Table 4 summarizes the physical properties of ionic and molecular compounds. Consider how the models for ionic and molecular compounds explain these properties.

Table 4 Comparison of Physical Properties of Ionic and Molecular Compounds

Property	Ionic compound	Molecular compound
State at ambient temperature	crystalline solid	solid, liquid, or gas
Physical properties	hard, brittle	solids can be soft, waxy, flexible, or crystalline
Relative melting point/boiling point	high	low
Electrical conductivity when dissolved in water	good	poor (for most molecular compounds)
Electrical conductivity in the liquid state	good	poor
Examples	sodium chloride, calcium carbonate	water, carbon dioxide, methane, sucrose

2.2 Summary

- There are two main types of chemical compounds: ionic and molecular.
- Molecular compounds have distinguishing physical properties. These properties can be explained by the nature of their chemical bonds.
- Molecular compounds are made up of molecules consisting of two or more atoms (usually non-metallic) held together with covalent bonds.
- A single covalent bond forms when two atoms share a pair of valence electrons. A double covalent bond forms when two atoms share two pairs of valence electrons.
- Atoms within molecules are held together by single, double, or triple covalent bonds. Some molecules contain more than one type of covalent bond.
- Lewis structures can represent molecular elements and compounds and polyatomic ions. The polyatomic ion charge must be taken into consideration.
- A structural formula is a simpler version of a Lewis structure, in which the lone pairs of electrons are not shown.

2.2 Questions

1. List three distinctive physical properties of molecular compounds. K/U
2. Draw the Lewis structure and the structural formula for each of the following molecular compounds: K/U C
   (a) $I_2(s)$
   (b) $CF_4(l)$
   (c) $NH_3(g)$
   (d) $H_2S(g)$
   (e) $SO_2(g)$
   (f) $PCl_3(g)$
   (g) $C_2H_2(g)$
   (h) $NF_3(g)$
   (i) $CH_3Cl(g)$
   (j) $C_3H_8(g)$
   (k) $C_2H_2(g)$
   (l) $C_2H_5OH(l)$
   (m) $CsCl(s)$
   (n) $CH_3OCH_3(g)$
3. List the compounds in Question 2 that contain multiple (double or triple) bonds. K/U
4. Draw a Lewis structure and a structural formula for each of the following polyatomic ions: K/U C
   (a) HCO_3^-
   (b) ClO^-
   (c) CO_3^{2-}
   (d) H_3O^+
5. Explain the differences between the formation of ionic bonds and the formation of covalent bonds. K/U
6. You have been given four substances that are all white solids at room temperature. List three specific tests that you could perform to figure out whether they are ionic or molecular compounds. T/I
7. Nitrogen dioxide, NO_2, is a major source of air pollution. At low temperatures, NO_2 levels are decreased as the unpaired electrons on the NO_2 molecules link up to form dinitrogen tetroxide, N_2O_4, a colourless gas. In the winter, the pollution is still there as dinitrogen tetroxide. Draw a Lewis structure for dinitrogen tetroxide. K/U C
8. Thalidomide is a pharmaceutical drug that was withdrawn from use in 1961 after it was discovered that thousands of pregnant women who had used the drug had given birth to babies with significant birth defects. Research the story of thalidomide, its molecular shape, and its effects. T/I

GO TO NELSON SCIENCE

2.3 Chemical Bonding and Electronegativity

In the atomic tug-of-war game, are all the atoms equally strong? Now that you are familiar with ionic and covalent bonds, we will look at them a little more closely. Evidence shows that the atoms of some elements attract electrons more strongly than do the atoms of other elements.

electronegativity the ability of an atom to attract bonding electrons to itself

Electronegativity

In Chapter 1, you learned that several atomic properties follow periodic trends. As you now explore chemical bonding in more detail, it is time to learn about another atomic property: electronegativity. **Electronegativity** is the ability of an individual atom, when bonded, to attract bonding electrons to itself. The American chemist Linus Pauling (1901–1994) developed the concept of electronegativity in 1922 (**Figure 1**). He also proposed a scale of electronegativity values. Simply put, comparing the electronegativity values of two atoms indicates the likelihood of those two atoms being part of an ionic or a molecular compound.

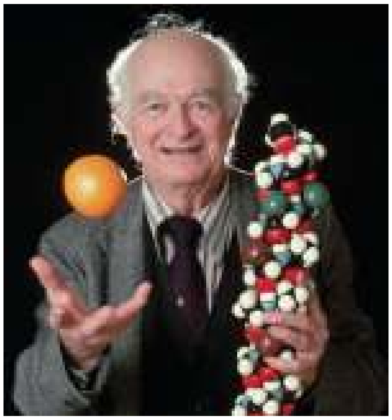

Figure 1 Linus Pauling won the 1954 Nobel Prize in Chemistry for, among other work, developing a better understanding of chemical bonds.

Pauling assigned the highest electronegativity value to fluorine (4.0) and the lowest value to francium (0.7). Fluorine is the element that has the greatest ability to attract a bonding pair of electrons. Francium has the lowest ability to attract bonding electron pairs to itself. In a way, electronegativity indicates how strong an atom is in a tug-of-war over bonding electrons. An element with a high electronegativity is very good at pulling a pair of electrons toward itself, whereas an element with a low electronegativity is not.

In general, the electronegativity of the elements increases from left to right on the periodic table and decreases from top to bottom (**Figure 2**). To explain this trend, it helps to consider atomic radius. Remember that atomic radius decreases as you move from left to right across the periodic table (Section 1.7). As the size of an atom increases, the attraction it has for a shared electron pair weakens. Therefore, electronegativity follows the opposite trend to atomic radius on the periodic table. As atomic radius increases, electronegativity decreases and vice versa.

The electronegativity of an element cannot be measured experimentally. Instead, it is calculated for each element using physical properties such as ionization energy. Accepted values of electronegativity are given in the periodic table at the back of this textbook. Values for some of the more common elements are also given in **Table 1**.

Table 1 Electronegativity Values of Selected Elements

Element	Electronegativity
H	2.2
C	2.6
N	3.0
O	3.4
F	4.0
Na	0.9
Mg	1.3
Cl	3.2
Ca	1.0

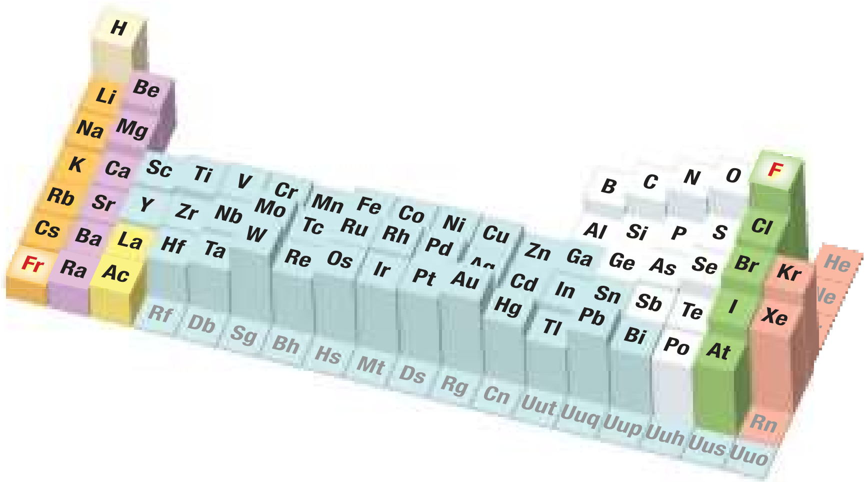

Figure 2 The periodic trends for electronegativity are the opposite of the trends for atomic radius.

Determining Whether a Compound Is Ionic or Molecular

As we started this chapter, we used the general rule that an ionic compound forms when a metal combines with a non-metal and that a molecular compound forms when two or more non-metals combine. Scientists have discovered, however, that the reality is not quite so simple. It turns out that electronegativity values help us to predict the nature of chemical bonds between elements.

To predict whether a bond is ionic or covalent, we must calculate the difference in the elements' electronegativities. This quantity is called the **electronegativity difference**, and has the symbol ΔEN (pronounced "delta E N") (**Table 2**).

electronegativity difference (ΔEN) the difference in electronegativities of two bonded atoms or ions

The electronegativity difference indicates how likely it is that an electron is transferred from one element to another. As you know, when an electron is transferred, the bond between the elements is an ionic bond. The greater the electronegativity difference, the more likely it is that the bond is ionic. For example, we can look at the electronegativity difference between sodium and chlorine. The difference is determined by subtracting smaller electronegativity value (0.9) from the larger one (3.2).

Table 2 Electronegativity Differences of Selected Bonds

Bond	Electronegativity difference (ΔEN)
H–H	0.0
C–H	0.4
N–H	0.8
O–H	1.2
Mg–Cl	1.9
Na–Cl	2.3
Ca–F	3.0

$$\begin{aligned}\text{electronegativity difference for Na}-\text{Cl} &= 3.2 - 0.9 \\ &= 2.3\end{aligned}$$

This relatively large difference indicates that chlorine is strong enough to win the tug-of-war for the electron. As a result, chlorine gains an electron to become a chloride ion, Cl^-, and sodium loses an electron to become a sodium ion, Na^+.

What would happen if two identical atoms were competing for a shared pair of electrons? The electronegativity difference between the two atoms—for example, two hydrogen atoms—would be 0. Electrons are shared equally between the atoms, so the bond that forms is a covalent bond (Section 2.2). The tug-of-war is a perpetual tie!

In fact, the ionic versus molecular nature of these bonds is quite gradual, like a spectrum. When the electronegativity difference is large ($\Delta EN \geq 1.7$), the bond is considered to be ionic. When the electronegativity difference is small ($\Delta EN < 1.7$), the bond is considered to be covalent.

Covalent Bonds: Non-polar and Polar

As you learned in Section 2.1, a covalent bond forms when two atoms share a pair of electrons. This occurs if the atoms have an electronegativity difference of less than 1.7. However, there are several types of covalent bonds. The size of the electronegativity difference affects the type of covalent bond that forms.

non-polar covalent bond a covalent bond formed between atoms with identical (or very similar) electronegativities

Non-polar Covalent Bonds

When the atoms are identical ($\Delta EN = 0$), they share the electrons equally. This type of covalent bond is called a **non-polar covalent bond**. The atoms in a molecule of chlorine gas, $Cl_2(g)$, are held together by a non-polar covalent bond (**Figure 3**). The tug-of-war for the electron pair is a tie because the chlorine atoms have an equal ability to attract the electrons. This is true for all diatomic elements: the atoms in their molecules are held together with non-polar covalent bonds.

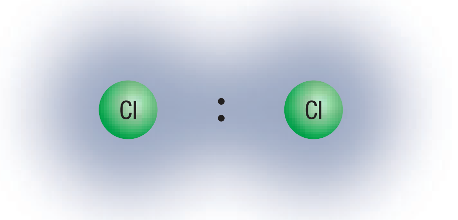

Figure 3 Two identical atoms share electrons equally, so a non-polar covalent bond forms between them.

Polar Covalent Bonds

When two covalently bonded atoms have significant electronegativity differences ($0 < \Delta EN < 1.7$), they do not share electrons equally. Consider the bond that forms between an atom of hydrogen, H, and an atom of chlorine, Cl (**Figure 4**). The atom with the higher electronegativity, Cl, attracts the electrons away from the atom with the lower electronegativity, H. Electrons have a negative charge. The chlorine atom is therefore surrounded by a localized negative charge, or negative "pole." At the same time, the hydrogen atom develops a positive pole. We represent these negative and positive poles with the symbols $\delta-$ and $\delta+$ respectively.

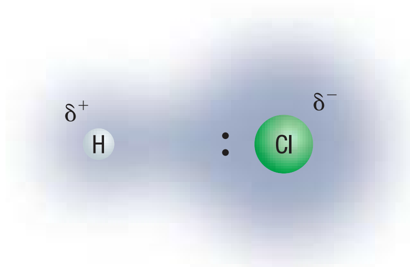

Figure 4 Two different atoms form a covalent bond in which the electron pair is not shared equally.

polar covalent bond a covalent bond formed between atoms with significantly different electronegativities resulting in a bond with localized positive and negative charges or poles

When atoms attract an electron pair unequally, a **polar covalent bond** forms. The most electronegative atom wins the tug-of-war for the electron pair, but not by enough to actually transfer an electron.

To summarize, the nature of a chemical bond depends on how bonding electron pairs are shared between two atoms. **Figure 5** shows this model of bonding as a continuum. A chemical bond between two atoms with an electronegativity difference of 0 is a non-polar covalent bond. A bond between two atoms with an electronegativity difference between 0 and 1.7 is a polar covalent bond. A bond between two atoms with an electronegativity difference greater than or equal to 1.7 is an ionic bond. These are not hard and fast rules but guidelines on a continuum of values and behaviour. We could have just as easily subdivided covalent bonds into "extremely," "moderately," "somewhat," and "hardly" polar!

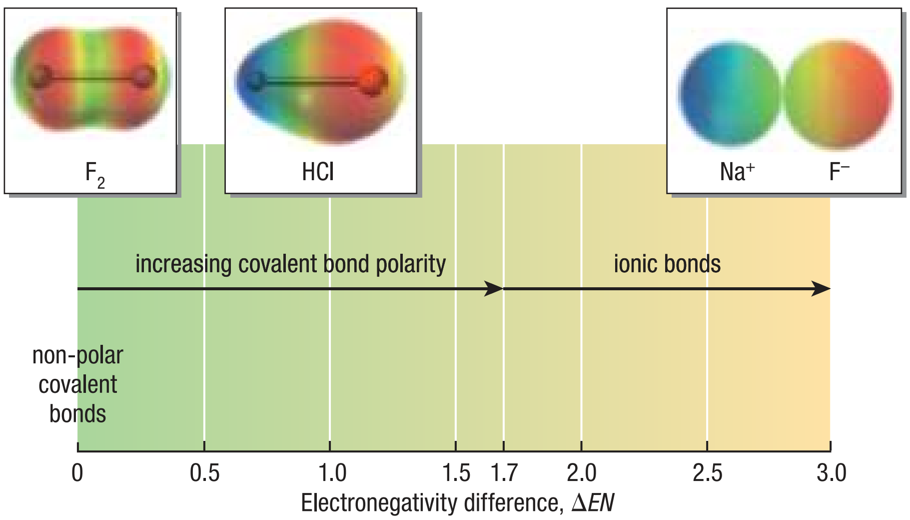

Figure 5 Electronegativity difference indicates the degree to which a chemical bond is ionic or covalent.

Remember that many molecules have more than one bond. These bonds are not always identical, even when they involve the same elements. As we continue to explore the nature of molecular compounds in Chapter 3, you will see that it is important to consider all of the bonds in a molecule.

2.3 Summary

- Electronegativity is a measure of an element's relative ability to attract bonding electron pairs.
- The electronegativity difference between two atoms can be used to classify the nature of the chemical bond.
- A large electronegativity difference ($\Delta EN \geq 1.7$) indicates that the bond between the two atoms is an ionic bond.
- A smaller electronegativity difference ($\Delta EN < 1.7$) indicates that the bond between the two atoms is a covalent bond. Covalent bonds can be further classified as either polar or non-polar.
- When two atoms have an intermediate electronegativity difference ($0 < \Delta EN < 1.7$), they share bonding electrons unequally in a polar covalent bond.
- When two atoms have the same electronegativity ($\Delta EN = 0$), they share bonding electrons equally in a non-polar covalent bond.

2.3 Questions

1. Explain how the trends of electronegativity in the periodic table relate to those of atomic radius. K/U
2. Using only their relative position on the periodic table, arrange the following elements in order of increasing electronegativity: K/U

   K Cs Br Fe Ca F Cl
3. Describe how we can use electronegativity values to predict the types of bonds that will form in a compound. K/U
4. Distinguish between a polar covalent bond and a non-polar covalent bond. K/U
5. For each bond listed, determine the electronegativity difference and predict what type of bond (non-polar covalent, polar covalent, or ionic) would form between the two elements. T/I
   (a) Ca–S
   (b) H–F
   (c) P–H
   (d) C–Cl
   (e) C–O
   (f) Li–Cl
6. Identify the more polar bond in each of the following pairs: T/I
   (a) H–F and H–Cl
   (b) O–H and C–H
   (c) C–N and N–N
7. Predict whether the chemical bonds in each of the following compounds will be non-polar covalent, polar covalent, or ionic: T/I
   (a) $N_2(g)$
   (b) $NH_3(g)$
   (c) $H_2O(l)$
   (d) $FeO(s)$
   (e) $MgCl_2(s)$
8. (a) What is the most ionic bond possible between any two elements in the periodic table? Explain your answer, referring to electronegativities.
   (b) Imagine a compound that contains this bond. What properties would you expect the compound to have? T/I A
9. Linus Pauling was one of the greatest chemists of all time. He had success in areas such as chemical bonding, protein and DNA structure, and the chemistry of vitamin C. He was also an influential peace activist. The United States government withheld his passport from him, considering him a threat to national security. On the announcement of his Nobel Peace Prize in 1962, his passport was quickly returned. Pauling was also devoted to his family. Research the accomplishments and life of Linus Pauling. Summarize your findings in a format of your choice. T/I C A
10. Chemists predicted, based on periodic properties, that it ought to be possible to make compounds of xenon and fluorine. Neil Bartlett, at the University of British Columbia, was the first to produce and identify these compounds. Research Bartlett's discovery. Prepare a brief presentation of your findings. T/I C A

GO TO NELSON SCIENCE

2.4 Chemical Formulas and Nomenclature

Figure 1 The name "saltpetre" may originate from the Latin term meaning "stone salt."

Names help us to identify people and to distinguish between various biological species. Similarly, chemical names and chemical formulas allow us to distinguish among millions of different elements and compounds.

Nomenclature Systems

Chemists use either a chemical formula or a chemical name to identify a specific compound. Every compound has a unique name. Before the 1700s, chemicals were named in a variety of ways. Traditional names such as water, sugar, ammonia, and bluestone became established throughout the ages. Saltpetre is the common name for two compounds: potassium nitrate and sodium nitrate. Saltpetre can be extracted from the sand of dry saltpetre dunes (**Figure 1**).

Table 1 Common Names of Compounds

Chemical formula	Common name
H_2O	water
NH_3	ammonia
KNO_3	saltpetre
K_2CO_3	potash
CaO	lime or quicklime
C_2H_5OH	grain alcohol
$KC_4H_5O_6$	cream of tartar
$MgSO_4$	Epsom salts

Table 1 lists common names for several compounds. It would now be extremely difficult to learn the names of all the compounds if they all had random and unique names. As more and more chemical compounds have been discovered and synthesized, a uniform, comprehensive, and standard naming system has become essential.

Fortunately, the International Union of Physical and Applied Chemistry (IUPAC) has established a logical naming (nomenclature) system. With this system, chemists around the world can communicate the identities of chemicals without confusion.

Ionic Compounds

The IUPAC system has different rules for naming ionic compounds and molecular compounds. We can determine whether a compound is ionic or molecular by examining the difference in the elements' electronegativities (Section 2.3).

A compound that contains only two different ions is called a **binary ionic compound**. Sodium chloride, NaCl, aluminum chloride, $AlCl_3$, and magnesium oxide, MgO, are examples of binary ionic compounds. Binary ionic compounds can be made up of more than two ions, but there must be only two different kinds of ions.

In contrast, a compound that contains ions made of two or more elements is a **polyatomic ionic compound**. A polyatomic ion is a single unit with a charge shared among its atoms. The atoms within the ion are joined by covalent bonds.

Many polyatomic ions belong to an important group known as oxyanions. An **oxyanion** is a negatively charged polyatomic ion that contains oxygen. Examples include chalk (calcium carbonate), $CaCO_3(s)$, and the food preservative sodium nitrite, $NaNO_2(s)$. Many acids, such as sulfuric acid, $H_2SO_4(aq)$, and nitric acid, $HNO_3(aq)$, contain oxyanions. You will learn more about these acids in Chapter 10.

binary ionic compound a compound that consists of ions of only two elements

polyatomic ionic compound a compound that consists of ions of more than two elements

oxyanion a negatively charged polyatomic ion that contains oxygen

Writing Chemical Formulas for Ionic Compounds

Recall that ionic compounds form when billions of non-metal atoms take electrons from vast numbers of metal atoms. There are two or more ions in the formula unit for a compound. For example, $AlCl_3(s)$ is the formula unit of the compound formed when 3 chlorine atoms each remove 1 electron from an aluminum atom (**Figure 2**).

Figure 2 When electrons are transferred, the overall charge on a formula unit is zero.

An ionic compound is electrically neutral. The **zero-sum rule** states that the total positive charge of the cations in an ionic compound must be equal to the total negative charge of the anions. As an example, sodium chloride has a chemical formula of NaCl. In each formula unit the +1 charge of the sodium ion and the −1 charge from the chloride ion add up to zero. The result is a neutral compound.

zero-sum rule the statement that the sum of the positive charges equals the sum of the negative charges in an ionic compound

Tutorial 1 Writing Chemical Formulas for Ionic Compounds

Writing chemical formulas for ionic compounds is fairly straightforward if you follow the steps outlined below.

Sample Problem 1: Writing the Chemical Formula of a Binary Ionic Compound

Write the chemical formula for sodium oxide.

Step 1. Write the symbols of the elements with the metal first and the non-metal second.

Na O

Step 2. Refer to the periodic table to determine the ionic charge of the ions. Write the charges above the symbols.

1+ 2−
Na O

Step 3. Determine how many of each type of ion are required to bring the total charge to zero.

$2(+1) + 1(-2) = 0$

Na O

Step 4. Write the chemical formula with subscripts representing the number of each type of ion.

Na_2O_1

Step 5. Make sure that the subscripts are a lowest whole number ratio. Omit the subscript "1."

The chemical formula for sodium oxide is Na_2O.

Sample Problem 2: Writing the Formula of a Polyatomic Ionic Compound

Write the chemical formula for aluminum sulfate.

Step 1. Write the symbols of the elements with the metal first and the polyatomic ion second.

Al SO_4

Step 2. Determine the ionic charge of the ions. Write the charges above the symbols. Remember that the charge on a polyatomic ion is shared over the entire ion.

3+ 2−
Al SO_4

Step 3. Determine how many of each type of ion are required to bring the total charge to zero.

$2(+3) + 3(-2) = 0$

Al SO_4

Step 4. Write the chemical formula with subscripts representing the number of each type of ion.

$Al_2(SO_4)_3$

Step 5. Make sure that the subscripts are a lowest whole number ratio. Omit the subscript "1."

The chemical formula for aluminum sulfate is $Al_2(SO_4)_3$.

Practice

1. Write the chemical formula for each of the following binary ionic compounds: T/I C
   (a) magnesium oxide (b) aluminum fluoride (c) potassium oxide
2. Write the chemical formula for the following polyatomic ionic compounds: T/I C
   (a) magnesium hydroxide (b) sodium bicarbonate (c) aluminum phosphate

LEARNING TIP

Ionic Charges
The ionic charges are usually written in the top right corner of each element's cell in the periodic table. You could also remember that all Group 1 elements form ions with a charge of 1+, Group 2 ions have a charge of 2+, Group 16 ions have a charge of 2–, and Group 17 ions have a charge of 1–.

WEB LINK

You will need to practise applying these rules. To try some online chemical nomenclature games,

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LEARNING TIP

Polyatomic Ions
Refer to Appendix B4, Table 4, for the formulas of polyatomic ions.

Naming Ionic Compounds

Sometimes you will encounter a chemical formula, rather than a chemical name. How do you translate this formula into a name, according to IUPAC rules?

Tutorial 2 Naming Ionic Compounds

When naming binary ionic compounds, the first part of the name comes from the metal and the second part comes from the non-metal (**Table 2**). Note that the name of the second element is modified to end in *–ide*.

Table 2 Naming Binary Ionic Compounds

Chemical formula	Metallic element	Metallic ion	Non-metallic element	Non-metallic ion	IUPAC name
$CaCl_2$	calcium, Ca	calcium, Ca^{2+}	chlorine, Cl	chloride, Cl^-	calcium chloride
Na_2O	sodium, Na	sodium, Na^+	oxygen, O	oxide, O^{2-}	sodium oxide
AlN	aluminum, Al	aluminum, Al^{3+}	nitrogen, N	nitride, N^{3-}	aluminum nitride

CASE 1: NAMING IONIC COMPOUNDS WITH MULTIVALENT IONS

In Section 1.3 you learned that some metal atoms are multivalent: they form a variety of cations, each with a different charge. Iron can form Fe^{2+} ions or Fe^{3+} ions, reacting with oxygen to form FeO or Fe_2O_3. If you were to name these compounds according to the rules you have learned so far, both substances would be "iron oxide." To clearly specify which ion is present, write the appropriate Roman numeral in curved brackets after the name of the ion. **Table 3** lists some common ionic compounds containing multivalent ions.

Table 3 Naming Multivalent Ionic Compounds

Chemical formula	Metallic element	Metallic ion	Non-metallic element	Non-metallic ion	IUPAC name
CuCl	copper, Cu	copper(I), Cu^+	chlorine, Cl	chloride, Cl^-	copper(I) chloride
$CuCl_2$	copper, Cu	copper(II), Cu^{2+}	chlorine, Cl	chloride, Cl^-	copper(II) chloride
SnO	tin, Sn	tin(II), Sn^{2+}	oxygen, O	oxide, O^{2-}	tin(II) oxide
SnO_2	tin, Sn	tin(IV), Sn^{4+}	oxygen, O	oxide, O^{2-}	tin(IV) oxide

To name an ionic compound containing a multivalent metal, first determine the charge of the non-metal ion in the compound, then use the zero-sum rule to determine the charge of the metal ion. Assume that the charge on common non-metal ions is always the same. For example, the charge on the ions of Group 16 elements is always -2 (e.g., O^{2-}) and the charge on the ions of Group 17 elements is always -1 (e.g., Cl^-).

LEARNING TIP

The Crisscross Method

You may have learned this alternative method for determining the charges on ions, given the formula of an ionic compound:

Fe_2O_3

$Fe^{3+}O^{2-}$

While the crisscross method can be a helpful tool, it is important to understand why it works, rather than simply memorizing how it works.

Sample Problem 1: Naming a Compound Containing a Multivalent Metal

Write the IUPAC name for Fe_2O_3.

Step 1. Identify any ions of multivalent metals or polyatomic ions in the compound.

This compound contains iron, which is a multivalent metal.

Step 2. Write the chemical formula with the charge of the anion (non-metal ion), above its symbol.

$$\overset{-2}{Fe_2O_3}$$

Step 3. Find the possible charges for the multivalent metal from the periodic table or from Table 1 in Section 1.3.

The charge on the Fe ion may be +2 or +3.

Step 4. Find the total negative charge for the chemical formula. There are 3 oxygen ions, each with a −2 charge. Total negative charge = −6

Step 5. Use the zero-sum rule to find the charge on each of the positive ions.

The total positive charge is +6. This charge is shared between two iron ions, so the charge on each Fe ion is +3.

Step 6. Write the name of the compound using the rules for naming ionic compounds. Indicate the charge of the multivalent metal with a Roman numeral in parentheses.

The IUPAC name of Fe_2O_3 is iron(III) oxide.

Practice

1. Write the chemical name for each of the following ionic compounds: K/U
   (a) $CuSO_4$ (b) CuCl (c) $SnCl_4$ (d) SnO

You may encounter chemical names such as cupric nitrate or stannic oxide (Section 1.3, Table 1). This classical naming system is still used in some industries for multivalent metals with just two possible charges. In this system, the Latin name for the element with the suffix *-ous* is used for the metal ion with the smaller charge. The Latin name with suffix *-ic* is used for the metal ion with the larger charge.

CASE 2: NAMING IONIC COMPOUNDS WITH POLYATOMIC IONS

There is a very useful and logical naming system for the oxyanions. Take a closer look at the four oxyanions that contain chlorine and oxygen in **Table 4**. These polyatomic ions all have the same charge but they have different numbers of oxygen atoms. The chlorate ion, ClO_3^-, is known as the parent ion. It is typically the most stable and common combination of the nonmetal and oxygen. The name of the parent oxyanion includes the stem of the non-metal element's name and the suffix *–ate*. In this case, *chlor* + *ate* or chlorate.

Try to become familiar with the parent oxyanions such as chlorate, ClO_3^-, nitrate, NO_3^-, sulfate, SO_4^{2-}, phosphate, PO_4^{3-}, and carbonate, CO_3^{2-}. You will encounter them often. The following additional naming rules work around these names. In Table 4, *x* represents the number of oxygen atoms in the parent ion: ___*ate*. In this case it is the chlorate ion, so *x* is 3.

- When the oxyanion has one more oxygen atom than the *-ate* oxyanion, add the prefix *per* and the suffix *-ate* to the stem of the non-metal name.
- When the oxyanion has one less oxygen atom than the *-ate* ion, add the suffix *-ite*.
- When the oxyanion has two fewer oxygen atoms than the *-ate* ion, add the prefix *hypo* and the suffix *-ite*.

Consider the oxyanions sulfate, SO_4^{2-}, and sulfite, SO_3^{2-}. They have the same electrical charge but different numbers of oxygen atoms, giving them very different properties.

Table 4 Naming Rules for Oxyanions

Number of O atoms	Naming rule	Example
$x + 1$	per____ate	perchlorate, ClO_4^-
x	____ate	chlorate, ClO_3^-
$x - 1$	____ite	chlorite, ClO_2^-
$x - 2$	hypo____ite	hypochlorite, ClO^-

WEB LINK

To learn more about the research into connections between sulfate and autism,

GO TO NELSON SCIENCE

Sample Problem 2: Naming Compounds Containing Polyatomic Ions

Determine the IUPAC name for the compound $Pb(ClO_4)_2$.

Step 1. Identify any ions of multivalent metals or polyatomic ions in the compound.

This compound contains lead, which is a multivalent metal, and a polyatomic ion containing chlorine and oxygen.

Step 2. Write the chemical formula with the charge of the anion above its symbol.

$$\overset{-1}{Pb(ClO_4)_2}$$

Step 3. Find the possible charges for the multivalent metal from the periodic table or from Table 1 in Section 1.3.

The charge on the Pb ion may be +2 or +4.

Step 4. Determine the name and charge of the polyatomic ion.

It is an oxyanion with one more oxygen atom than chlorate, so it is perchlorate. Each perchlorate ion has a charge of −1.

Step 5. Find the total negative charge for the chemical formula.

2 perchlorate oxyanions, each with a −1 charge, have a total negative charge of −2.

Step 6. Use the zero-sum rule to find the charge on each of the positive ions.

The total positive charge is +2. There is one lead ion, so the charge on that Pb ion is +2.

Step 7. Write the name of the compound. Indicate the charge of the multivalent metal with a Roman numeral in parentheses. Use the correct name for any polyatomic ions.

The name of $Pb(ClO_4)_2$ is lead(II) perchlorate.

Practice

2. Determine the chemical name for each of the following ionic compounds: T/I
   (a) $Pb(SO_3)_2$ (c) Cu_3PO_4 (e) NaClO
   (b) $Pb(NO_3)_2$ (d) $Fe(OH)_3$ (f) $(NH_4)_2CO_3$

hydrate an ionic compound that contains water as part of its crystal structure

HYDRATES: NAMING AND WRITING CHEMICAL FORMULAS

A **hydrate** is an ionic compound that includes water molecules within its crystal structure. Bluestone is the common name for a hydrate of copper(II) sulfate. When water is evaporated from a solution of copper(II) sulfate, some water molecules are "trapped." When this water, called the water of hydration, is removed, the anhydrous form of the ionic compound remains. ("Anhydrous" means "without water.") The water of hydration for any given compound is always in a fixed ratio with the formula unit.

The Egyptians used a mixture called natron in the process of embalming bodies (**Figure 3**). One of the compounds in natron is anhydrous sodium carbonate. It extracted the water from the corpse to produce sodium carbonate heptahydrate ($Na_2CO_3•7H_2O$). This drying process was a critical step in preparing the corpse because with no water present, decay was minimized and the body was preserved for years. In a more recent application, anhydrous calcium chloride is used to absorb water in packaging for electronic equipment.

Figure 3 This artist's impression depicts an ancient Egyptian embalming process that used the dessicant natron.

Table 5 Prefixes Used in the Names of Hydrates and Molecular Compounds

Number of atoms or water molecules in the chemical formula	Prefix
1	mono or mon
2	di
3	tri
4	tetra
5	penta
6	hexa
7	hepta
8	octa
9	nona
10	deca

Tutorial 3 Naming and Writing Formulas of Hydrates

CASE 1: NAMING HYDRATES

The traditional naming convention for hydrates specifies the name of the ionic compound, the fact that water is associated with it, and the ratio in which the two substances occur. In bluestone, 5 water molecules are associated with each formula unit of copper(II) sulfate. This hydrate is called copper(II) sulfate pentahydrate. The traditional naming system for hydrates uses the name of the ionic compound, a Greek prefix (**Table 5**), and the word "hydrate."

Although you do not have to remember them, there are other naming systems for hydrates. One involves using the name of the ionic compound, plus the number of water molecules, plus the word "water." For example, magnesium sulfate heptahydrate would be called magnesium sulfate-7-water. Another gives the name of the ionic compound, the word "water," and the ratio of formula units to water molecules: magnesium sulfate—water (1/7).

Sample Problem 1: Writing the Name for a Hydrate

Determine the traditional name for a compound in which nickel(II) chloride is bonded with water in a ratio of 1 to 6.

Step 1. Look up the prefix (in Table 5) for the number of water molecules associated with each formula unit.

There are 6 water molecules, so the prefix is *hexa*-.

Step 2. Add the prefix to "hydrate" at the end of the compound's name.

The name is nickel(II) chloride hexahydrate.

CASE 2: WRITING CHEMICAL FORMULAS FOR HYDRATES

Writing the chemical formulas of hydrates is quite straightforward. The chemical formula for copper(II) sulfate pentahydrate is $CuSO_4 \cdot 5H_2O$ (**Figure 4**). Indicate the ratio of chemical formulas by writing the number of water molecules before the formula for water.

Sodium sulfate decahydrate has the chemical formula $Na_2SO_4 \cdot 10H_2O$. Similarly, we can communicate cobalt(II) chloride hexahydrate as $CoCl_2 \cdot 6H_2O$.

Sample Problem 2: Writing the Chemical Formula for a Hydrate

Write the chemical formula for barium hydroxide octahydrate.

Step 1. Write the formula for the ionic compound without the associated water. (See Tutorials 1 and 2 in this section.)

Barium hydroxide is $Ba(OH)_2$.

Step 2. Look up the number of water molecules for the prefix associated with "hydrate."

The prefix is *octa*-, so there are 8 water molecules with each formula unit.

Step 3. Write the compound's formula followed by the number of water molecules and the formula for water.

The chemical formula is $Ba(OH)_2 \cdot 8H_2O$.

Practice

1. Write the name and chemical formula for each of the following hydrates: K/U C
   (a) a compound of calcium and chlorine in which there are 2 water molecules for each formula unit of the ionic compound
   (b) sodium sulfate with 10 water molecules per formula unit

Figure 4 Hydrated copper(II) sulfate is a distinctive blue colour. When the same copper compound has no water in its structure, it is white.

Molecular Compounds

Even though there are fewer non-metals than metals, there are many more combinations of non-metals resulting in many more molecular compounds. Nitrogen and oxygen alone can form four different compounds: N_2O, NO, NO_2, and N_2O_4. Nitrogen dioxide, NO_2, is a deadly, reddish-brown gas. Dinitrogen monoxide, N_2O, also known as nitrous oxide or laughing gas, is used as an anesthetic. Each compound must have a unique name to distinguish it from the others. Many molecular compounds have traditional or common names, which may or may not give clues about their chemical composition. Fortunately, an established naming system for molecular compounds avoids any confusion.

Investigation 2.4.1

Ionic or Molecular? (p. 85)
In this investigation you will try to identify substances using only information about their properties.

Tutorial 4 Formulas of Molecular Compounds

The names of molecular compounds include prefixes to indicate the numbers of atoms in the molecular formula. These prefixes are the same ones used for hydrates (Table 5).

CASE 1: NAMING MOLECULAR COMPOUNDS

The prefix is always attached to the element to which it refers. For example, the chemical name dinitrogen tetraoxide indicates that there are 2 nitrogen atoms and 4 oxygen atoms in the molecular formula. Therefore the chemical formula for dinitrogen tetraoxide is N_2O_4. Note that the prefix *mono*- is used only for the second element in the compound. For example, CF_4 is named carbon tetrafluoride and not monocarbon tetrafluoride.

Sample Problem 1: Naming Molecular Compounds

Determine the chemical name for $P_2O_3(s)$.

Step 1. Look up the prefix for the number of atoms of each element (Table 5).

Place the prefix *di* in front of "phosphorus" and *tri* in front of "oxygen."

Step 2. If applicable, omit *mono* from the first element.

The name is diphosphorus trioxide.

Sample Problem 2: Naming Molecular Compounds

Determine the chemical name for water, H_2O.

Step 1. Look up the prefix for the number of atoms of each element.

Place the prefix *di* in front of "hydrogen" and *mon* in front of "oxygen." This is the shortened form of *mono* because "oxygen" begins with a vowel.

Step 2. If applicable, omit *mono* from the first element.

The chemical name for water is dihydrogen monoxide.

CASE 1: WRITING FORMULAS FOR MOLECULAR COMPOUNDS

The same rules apply for writing chemical formulas for molecular compounds. Attach the prefix to the appropriate element (Table 4). For example, the chemical formula for sulfur hexafluoride is SF_6, and the formula for dinitrogen pentoxide is N_2O_5.

Practice

1. Name each of the following molecular compounds: T/I
   (a) CCl_4 (b) NO_2 (c) P_2O_5 (d) CF_4
2. Write the chemical formula for each of the following molecular compounds: T/I
   (a) carbon monoxide (b) sulfur dioxide (c) phosphorus pentafluoride

Research This

What's in a Name?

Skills: Researching, Analyzing, Communicating SKILLS HANDBOOK A5.1

It is interesting to find out what ingredients are actually in various prepared foods. In this activity, you will find a list of ingredients for a food item and analyze the contents.

1. Find a list of ingredients for a food item. It could be the list on a package at home, at a grocery store, or online. The list should include at least five named chemical compounds. Bring in the list, or a copy of it.
2. Research information on the chemicals listed.

A. Which of the ingredients on your list are ionic compounds? K/U

B. Which of the ingredients on your list are molecular compounds? K/U

C. How were you able to distinguish between ionic and molecular compounds? K/U

D. Determine the chemical formula of each compound. T/I C

E. Were all of the compounds named according to the rules outlined in this section? Comment on your observations. K/U A

GO TO NELSON SCIENCE

2.4 Summary

To name a chemical compound . . .

Does the formula begin with H?

- **Yes** → Are there two atoms, both the same?
  - **Yes** → **It is the diatomic gas H_2.**
  - **No** → **It is water or an acid.** (more detail in Unit 4)
- **No** → Are the bonds covalent (ΔEN <1.7)?
  - **Yes** → **It is a molecular substance.** Are all the elements the same?
    - **Yes** → **It is a molecular element.** Its name is the same name as the element; e.g., Cl_2 is chlorine.
    - **No** → **It is a molecular compound.** Name the first element using the proper prefix (*di*, *tri*, etc.), but never *mono*. Name the second element with the proper prefix (including mono) and the "ide" ending.
  - **No** → **It is an ionic compound.** Does the chemical formula begin with a multivalent metal ion?
    - **Yes** → Name the first element followed by its charge as a Roman numeral (I, II, III, etc.); e.g., FeO is iron(II) oxide. → Does the formula contain a polyatomic ion?
    - **No** → Does the formula contain a polyatomic ion?
      - **No** → **It is a binary ionic compound.** Name the first element, then name the second element with the *ide* ending, e.g., NaCl is sodium chloride.
      - **Yes** → Does the formula include H_2O?
        - **Yes** → **It is a hydrate.** Name the ionic compound as normal and include the proper prefix (*di*, *tri*, etc.) with the word "hydrate" after; e.g., $CuSO_4{\bullet}5H_2O$ is copper(II) sulphate.
        - **No** → **It is a polyatomic ionic compound.** Name the ions in the order in which they appear in the chemical formula, e.g., NH_4Cl is ammonium chloride, NH_4NO_3 is ammonium nitrate.

2.4 Questions

1. Write the chemical formula for each of the following ionic compounds: K/U C
   (a) lithium chloride
   (b) potassium sulfide
   (c) iron(II) chloride
   (d) aluminum oxide
   (e) sodium sulfate
   (f) tin(IV) oxide
2. Describe the IUPAC rules for naming a compound with a multivalent ion. K/U T/I
3. Write the IUPAC name for each of the following compounds: K/U
   (a) $MgCl_2$
   (b) Cs_2O
   (c) FeS
   (d) Na_3PO_4
   (e) NH_4NO_3
   (f) $Al_2(SO_4)_3$
   (g) $Mg(ClO_3)_2$
   (h) $Pb(BrO_3)_2$
   (i) $ZnHPO_4$
   (j) NaCN
4. Write the IUPAC name for each of the following molecular compounds: K/U
   (a) PCl_5 (b) N_2O_5 (c) CF_4 (d) SO_2
5. **Figure 5** shows the hydrated and anhydrous forms of cobalt(II) chloride. Write the chemical formulas for the hydrated and anhydrous forms of this compound. K/U C

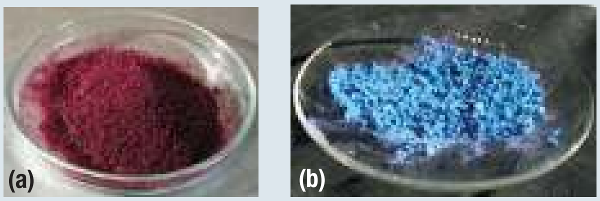


Figure 5 (a) Cobalt(II) chloride hexahydrate and (b) anhydrous cobalt(II) chloride have different physical properties.

6. Write the chemical formula for each of the following molecular compounds: K/U C
   (a) phosphorus trichloride
   (b) carbon tetrachloride
   (c) nitrogen monoxide
   (d) disulfur dichloride
7. Write the names for the following substances: K/U
   (a) KOH(s) (used in making soap)
   (b) $NaNO_2$(s) (a meat preservative)
   (c) CuCl(s) (used to colour firework displays)
   (d) NaOH(s) (drain cleaner)
   (e) $CaCO_3$(s) (the main component of limestone and chalk)
8. Why is it important to have a standardized IUPAC nomenclature system? K/U

Explore an Issue in Chemical Compounds

SKILLS MENU

- Defining the Issue
- Researching
- Identifying Alternatives
- Analyzing
- Defending a Decision
- Communicating
- Evaluating

Sugar versus Artificial Sweeteners

We love sweets! Sugar is one of the most popular additives in food and beverage products. The Canadian Sugar Institute estimates that Canadians consume an average of 63 g of sugar a day, or 23 kg a year! Sugar is present in food and beverages in many forms, such as sucrose, glucose, and fructose. Sucrose (regular table sugar) comes primarily from sugar cane and sugar beets. A sucrose molecule is made up of a glucose molecule and a fructose molecule joined together. Sugars are a source of quick energy. Nutritionists warn us, however, that sugars have no other nutritional value: they are "empty calories."

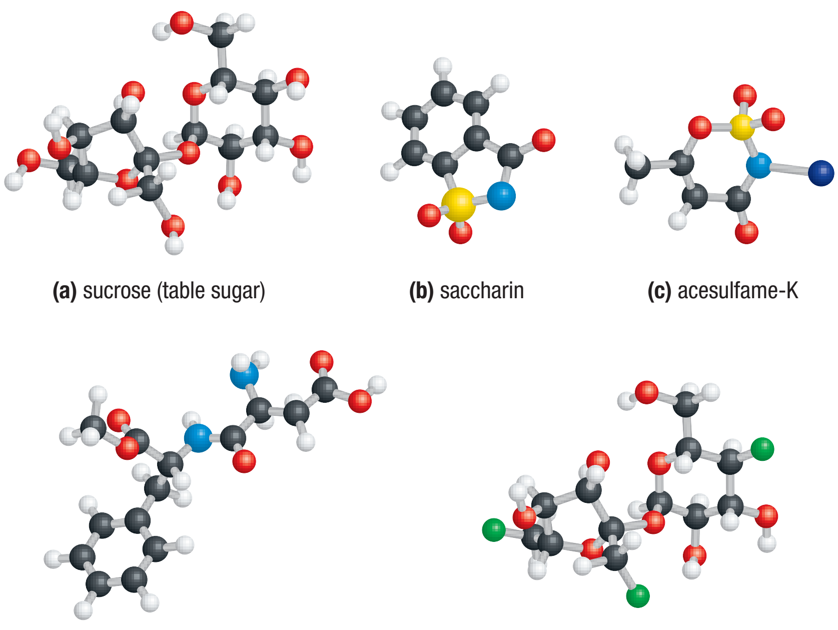
Consuming too much sugar can lead to increased chances of developing type 2 diabetes, becoming obese, and developing tooth decay. As a result, the Heart and Stroke Foundation recommends that we should consume no more than 48 g of added sugar per day. As a can of pop contains up to 39 g of sugar, it is very easy to exceed these limits. People who have diabetes sometimes struggle to reduce their sugar intake. Many would like to have the sweet taste of sugar without the calories.

There are several sweeteners that we can use instead of sugar. Sweeteners are either carbohydrates that are sweeter than sucrose or non-carbohydrate compounds classed as artificial sweeteners (**Table 1**). One of the most common carbohydrate sweeteners is high fructose corn syrup (HFCS). Because fructose is much sweeter than glucose, manufacturers engineer sucrose molecules so that most of the glucose is converted to fructose. The resulting syrup is sweeter than sucrose so a smaller quantity is needed to produce the same sweetness as sugar. The soft-drink industry mainly uses HFCS.

There are several artificial sweeteners in use today including saccharin, aspartame, acesulfame-K, and sucralose. Most artificial sweeteners are synthetic compounds that bear little chemical resemblance to sucrose. Sucralose, however, is identical to sucrose except that three chlorine atoms replace three oxygen and hydrogen groups in each molecule (**Figure 1**).

Table 1 Comparison of Sugars and Artificial Sweeteners

Sweetener	Approximate energy content (cal/g)	Relative sweetness (sucrose = 100)
Sugars		
glucose	3.9	75
lactose	n/a	16
sucrose	4.2	100
fructose	n/a	175
Artificial sweeteners		
aspartame	0.5	18 000
saccharin	0	45 000
sucralose	0.4	60 000



(a) sucrose (table sugar) **(b)** saccharin **(c)** acesulfame-K **(d)** aspartame **(e)** sucralose

Figure 1 Molecular structures of sucrose (table sugar) and artificial sweeteners

While consuming less sugar is generally a healthy step, there may be health problems associated with artificial sweeteners. Saccharin, $C_7H_5NO_3S$, has been linked to bladder cancer in rats. It is now banned in Canada, but is still used in the United States. Aspartame, $C_{14}H_{18}N_2O_5$, is legally approved as a safe alternative to sugar but it breaks down at temperatures above 30 °C and in the digestive tract (**Figure 2**). One of the breakdown products, phenylalanine, is a common compound in many foods but is dangerous to individuals who cannot metabolize it properly. This rare genetic condition is known as phenylketonuria (PKU). A buildup of phenylalanine can lead to brain damage. Acesulfame-K, $C_4H_4KNO_4S$, is similar to aspartame but it is more stable when heated. Finally, sucralose, $C_{12}H_{19}Cl_3O_8$, is widely considered to be one of the safest artificial sweeteners, with animals studies showing adverse affects only at doses much higher than the normal daily intake.

The Issue

A large soft-drink company has contracted your product development team to recommend a sweetener for a new kind of pop.

Goal

To develop options and selection criteria, assess the choices, and ultimately defend a decision regarding the best sweetener to use.

Research

SKILLS HANDBOOK A5.1

Work in a small group to learn more about the risks associated with excessive sugar intake and with the use of artificial sweeteners. Consider criteria such as calories, costs, and health concerns. Find answers to the following questions.

- What sweetening additives are presently used and why?
- What are the health concerns associated with the various options?
- What are the marketing concerns (such as customer demographics, cost, and perception) for your product, the new pop?

Identify Solutions

- Brainstorm a list of possible solutions.
- Set selection criteria (calories, cost, health concerns, etc.) and develop a rating system for these criteria.
- Based on your research, evaluate the different alternatives according to your selection criteria.

Make a Decision

Which sweetener will your team choose? Outline the rationale you used to reach your decision. Be prepared to defend your decision.

Communicate

Prepare a presentation of your findings and recommendation for the company executives. Remember that most of them are not chemists or nutritionists. However, they will make the final decision on whether or not to proceed with your recommendation.

CAREER LINK

Most Canadians have a sweet tooth: we like the taste of sugar. Dieticians are trying to educate people to reduce the amount of sugar in their diet. To find out more about the work of a dietician,

GO TO NELSON SCIENCE

Figure 2 Aspartame and sucralose are among the most commonly available calorie-free sweeteners in Canada.

WEB LINK

To get started with your research into the pros and cons of artificial sweeteners,

GO TO NELSON SCIENCE

UNIT TASK BOOKMARK

Some of the decision-making processes that you use in this section will also be useful as you work on the Unit Task, which is described on page 134.

Plan for Action

You have now found out quite a lot about the benefits and drawbacks of artificially sweetened pop. Is it appropriate to have diet pop available for sale in your school? Who makes the decisions about what drinks are sold? Plan a campaign to influence that decision. Make sure that you support your suggestions with researched facts.

GO TO NELSON SCIENCE

Investigation 2.1.1 CONTROLLED EXPERIMENT

SKILLS MENU

- Questioning
- Researching
- Hypothesizing
- Predicting
- Planning
- Controlling Variables
- Performing
- Observing
- Analyzing
- Evaluating
- Communicating

Comparing Ionic and Molecular Compounds

In this investigation you will formulate a testable question that deals with the physical properties of ionic and molecular compounds. You will design an experiment to answer your question. Pay special attention to identifying and controlling experimental variables.

Testable Question

Write a question about a physical property of ionic and molecular compounds that your experiment will attempt to answer. You might examine electrical conductivity, ability to dissolve in water, texture, hardness, or melting and/or boiling points.

Hypothesis

SKILLS HANDBOOK A2.2

State a hypothesis related to your question. Your hypothesis should include a prediction and a theoretical explanation for that prediction.

Variables

Identify the variable that you will purposely manipulate in this experiment, the variable that you will measure, and the variable(s) that you will keep controlled throughout the experiment.

Experimental Design

Write your own design for your experiment. Consider how you will measure and/or control your experimental variables. Consider what evidence you need to collect in order to answer your question.

Equipment and Materials

SKILLS HANDBOOK A1, A3.1

The following equipment and materials will be available. List those you will require.

- lab apron
- chemical safety goggles
- 50 mL beakers
- test tubes
- test-tube rack
- test-tube holder
- scoopula
- eye dropper
- retort stand with support ring and clamp
- stirring rod
- low-voltage conductivity tester
- well plate
- crucibles
- Bunsen burner clamped to a retort stand
- spark lighter
- clay triangles
- thermometer
- sucrose, $C_{12}H_{22}O_{11}(s)$
- iron(III) oxide, $Fe_2O_3(s)$
- wax, $C_{25}H_{52}(s)$
- sodium hydrogen carbonate, $NaHCO_3(s)$
- potassium chloride, $KCl(s)$
- sodium iodide, $NaI(s)$
- lauric acid, $C_{12}H_{24}O_2(s)$
- distilled water
- ice
- unknown compound #1
- unknown compound #2

There may be open flames in the laboratory. Tie back long hair and secure loose clothing. Do not leave flames unattended.

Procedure

1. Write a procedure, following your experimental design, to answer your testable question. Pay special attention to any safety concerns.
2. Obtain your teacher's approval for your procedure. Perform the experiment by closely following your procedure, including all the safety precautions. Record all observations.
3. Obtain samples of unknown compound #1 and unknown compound #2. Perform tests to determine whether these compounds are ionic or molecular compounds.

Analyze and Evaluate

SKILLS HANDBOOK A2.6

(a) Did your experiment allow you to answer your specific testable question? Explain. T/I
(b) Were you able to control and manipulate the variables as planned? Explain. T/I
(c) Use your evidence to answer your testable question. T/I
(d) According to your evidence, are the unknown compounds ionic or molecular? T/I
(e) Did your evidence support your hypothesis? Explain. T/I

(f) Based on your experience, evaluate your experimental design and suggest further studies and improvements. T/I

(g) As a class compile a table to compare the properties of ionic and molecular compounds. T/I

Apply and Extend

(h) Look up another physical property of your substances to learn more about ionic and molecular compounds. Cite your references. T/I

GO TO NELSON SCIENCE

Investigation 2.4.1 OBSERVATIONAL STUDY

SKILLS MENU

- Questioning
- Researching
- Hypothesizing
- Predicting
- Planning
- Controlling Variables
- Performing
- Observing
- **Analyzing**
- **Evaluating**
- **Communicating**

Ionic or Molecular?

Table 1 provides data on the physical properties of seven compounds. Your task is to classify the compounds as either ionic or molecular.

Purpose

To use provided data to determine whether each of the listed compounds is ionic or molecular

Procedure

SKILLS HANDBOOK A2.4

1. Copy **Table 1** into your notebook. Examine the data.

Analyze and Evaluate

(a) Consider the differences in physical properties for ionic and molecular compounds, and predict which compounds are ionic and which are molecular. Complete the two right-hand columns of your table. Most of the substances have been discussed in this chapter. T/I

(b) Which properties were the most useful in categorizing the compounds? T/I

(c) Which properties were the least useful in categorizing the compounds? T/I

(d) Were there any observations that confused your decisions? Explain. T/I

Apply and Extend

SKILLS HANDBOOK B3

(e) Compare and contrast the physical properties of ionic and molecular compounds. K/U

(f) Extend your table by adding three more rows. Add three more compounds of your choice. Research their properties and complete each column. T/I

(g) Suggest a use for each compound in your table. A

GO TO NELSON SCIENCE

Table 1 Observations of Physical Properties of Seven Compounds, A to G

Compound	Appearance at room temperature	Mixes or dissolves in water?	Conducts electricity in water?	Melting point (°C)	Boiling point (°C)	Other distinguishing features	Ionic or covalent?
A	fine white powder	only very slightly	no	825 (decomposes)			
B	blue crystals	yes	yes	110 (decomposes)		crystal structure	
C	colourless liquid	n/a	n/a		100		
D	colourless liquid	no	n/a		96	pungent odour	
E	white crystals	yes	yes	801			
F	white crystals	yes	no	186			
G	colourless gas	no	n/a		−57		

Summary Questions

1. Create a study guide based on the points listed in the margin on page 54. For each point, create three or four sub-points that provide further information, relevant examples, explanatory diagrams, or general equations.
2. Look back at the Starting Points questions on page 54. Answer these questions using what you have learned in this chapter. Compare your latest answers with those that you wrote at the beginning of the chapter. Note how your answers have changed.
3. Make a concept map to summarize how bonding accounts for the characteristic properties of ionic and molecular compounds.

Vocabulary

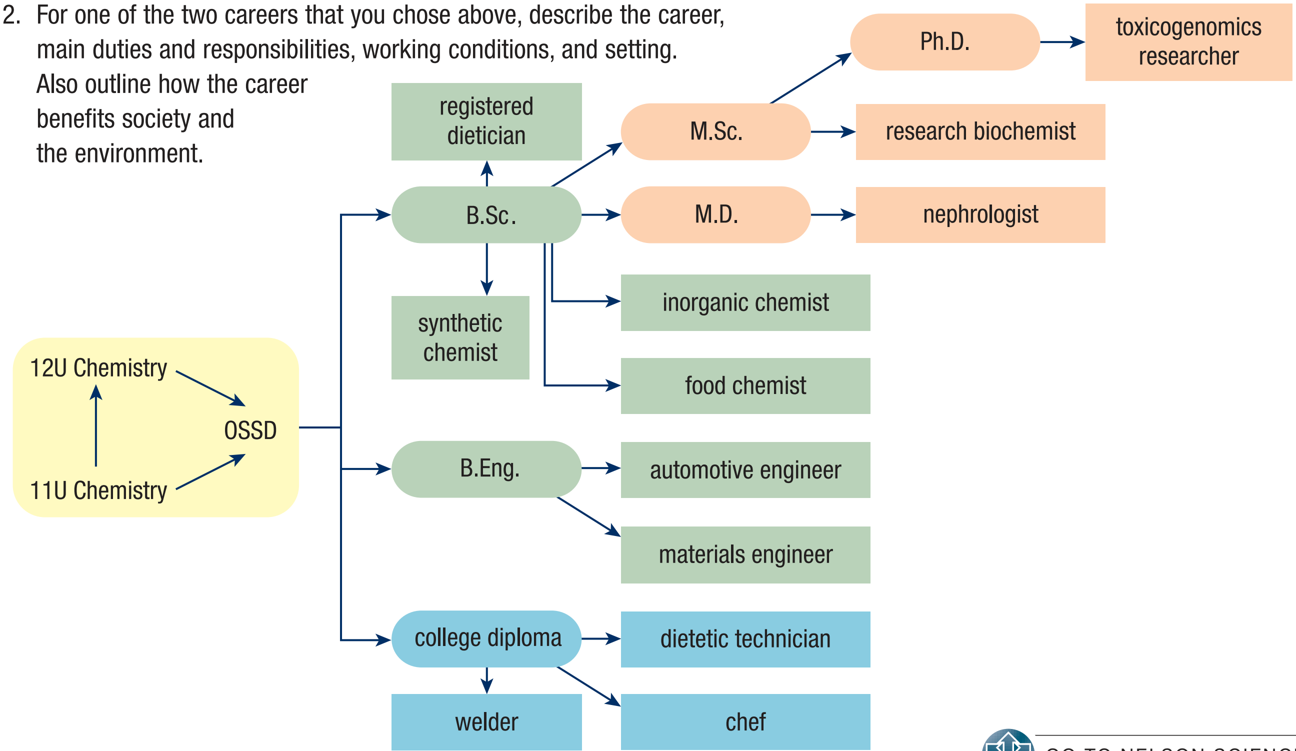
chemical bond (p. 56)	diatomic (p. 61)	lone pair (p. 63)	polar covalent bond (p. 72)
electrolyte (p. 57)	molecular compound (p. 61)	structural formula (p. 63)	binary ionic compound (p. 74)
ionic bond (p. 58)	covalent bond (p. 62)	electronegativity (p. 70)	polyatomic ionic compound (p. 74)
formula unit (p. 59)	bonding electron (p. 62)	electronegativity difference (ΔEN) (p. 71)	oxyanion (p. 74)
ionic compound (p. 59)	bonding capacity (p. 62)	non-polar covalent bond (p. 71)	zero-sum rule (p. 74)
molecular element (p. 61)	Lewis structure (p. 62)		hydrate (p. 78)

CAREER PATHWAYS

SKILLS HANDBOOK A7

Grade 11 Chemistry can lead to a wide range of careers. Some require a college diploma or a B.Sc. degree. Others require specialized or postgraduate degrees. This graphic organizer shows a few pathways to careers mentioned in this chapter.

1. Select two careers, related to chemical compounds and bonding, that you find interesting. Research the educational pathways that you would need to follow to pursue these careers. What is involved in the required educational programs? Prepare a brief report of your findings.
2. For one of the two careers that you chose above, describe the career, main duties and responsibilities, working conditions, and setting. Also outline how the career benefits society and the environment.



GO TO NELSON SCIENCE

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

For each question, select the best answer from the four alternatives.

1. Which is most likely to be true for an ionic compound? (2.1) K/U
   (a) It is soft and waxy in the solid state.
   (b) It has a high melting point.
   (c) In the liquid state, it is a poor conductor of electricity.
   (d) In the solid state, it is a good conductor of electricity.
2. Which substance is most likely to be an electrolyte? (2.1) K/U
   (a) $C_6H_{12}O_6(s)$
   (b) $CH_3OH(l)$
   (c) $Na_2CO_3(s)$
   (d) $CCl_4(l)$
3. Which molecule has the highest number of valence electrons? (2.2) K/U
   (a) $PH_3(g)$
   (b) $SiF_4(g)$
   (c) $CO(g)$
   (d) $Cl_2(g)$
4. Which combination of elements is most likely to form ionic bonds? (2.3) K/U
   (a) potassium and fluorine
   (b) carbon and oxygen
   (c) hydrogen and sulfur
   (d) xenon and fluorine
5. The IUPAC name for the compound $MgCl_2(s)$ is
   (a) magnesium chloride
   (b) magnesium dichloride
   (c) magnesium(II) chloride
   (d) magnesium chlorite (2.4) K/U
6. The IUPAC name for the compound $Cu_3(PO_4)_2(s)$ is
   (a) copper phosphate
   (b) tricopper diphosphate
   (c) copper(II) phosphate
   (d) copper phosphide (2.4) K/U
7. The IUPAC name for the compound $N_2O(g)$ is
   (a) nitrogen oxide
   (b) dinitrogen monoxide
   (c) mononitrogen dioxide
   (d) nitrogen(II) oxide (2.4) K/U
8. The correct chemical formula for the compound tin(II) carbonate is
   (a) Tn_2CO_3
   (b) $SnCO_3$
   (c) Sn_2CO_3
   (d) $Sn_2(CO_3)_2$ (2.4) K/U
9. A correct name for the compound $CaCl_2{\bullet}2H_2O(s)$ is
   (a) calcium chloride
   (b) calcium chloride diwater
   (c) calcium chloride dihydrate
   (d) calcium chloride dihydrogen monoxide (2.4) K/U

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

10. The molecular formula for the compound sodium chloride is NaCl(s). (2.1) K/U
11. The chemical bond that forms between a hydrogen atom and a chlorine atom will be non-polar covalent. (2.2, 2.3) K/U
12. A molecule of the compound nitrogen trihydride, also known as ammonia, would have one lone pair in its Lewis structure. (2.2) K/U
13. A molecule of carbon dioxide would have carbon–oxygen single bonds in its Lewis structure. (2.2) K/U
14. In the periodic table, electronegativity values increase from top to bottom and from left to right. (2.3) K/U
15. In the periodic table, electronegativity values tend to increase as atomic radius decreases. (2.3) K/U
16. The chemical formula for the compound aluminum oxide is Al_2O_3. (2.4) K/U
17. The chemical formula for the meat preservative sodium nitrite is Na_3N. (2.4) K/U
18. The chemical formula for the compound tin(IV) sulfide is TiS_2. (2.4) K/U
19. The charge on the chromate ion in the compound potassium chromate, $K_2CrO_4(s)$, is +2. (2.4) K/U
20. A can of pop can contain up to 39 g of sugar. (2.5) K/U

To do an online self-quiz,
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CHAPTER 2 REVIEW

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

Knowledge

For each question, select the best answer from the four alternatives.

1. A calcium atom loses 2 electrons to become stable. Which shows the Lewis symbol of the particle that would be formed? (2.1) K/U
   (a) $\left[:\ddot{\underset{\cdot\cdot}{Ca}}:\right]$
   (b) $\left[\ Ca\ \right]^{2-}$
   (c) $\left[\ Ca\ \right]^{2+}$
   (d) $\left[\ Ca\ \right]$
2. Which is a molecular element? (2.2) K/U
   (a) helium
   (b) beryllium
   (c) oxygen
   (d) neon
3. Silicon is a very commonly used component for computers, cellphones, and other electronic devices. The bonding capacity of a silicon atom is
   (a) 1
   (b) 2
   (c) 3
   (d) 4 (2.2) K/U
4. Which of the following is a typical physical property of a molecular compound? (2.2) K/U
   (a) has a crystal lattice
   (b) dissolves in water to conduct electricity
   (c) is held together by ionic bonds
   (d) has a relatively low boiling point
5. Which is a molecular compound? (2.2, 2.3) K/U
   (a) NH_3
   (b) $CaCl_2$
   (c) NO_3^-
   (d) Na_2O
6. How does electronegativity vary in the periodic table? (2.3) K/U
   (a) It increases from top to bottom and from left to right.
   (b) It decreases from top to bottom but increases from left to right.
   (c) It increases from top to bottom but decreases from left to right.
   (d) It decreases from top to bottom and from left to right.
7. What type of bond is likely to form between the elements molybdenum, Mo, and chlorine, Cl? (2.3) K/U
   (a) non-polar covalent
   (b) polar covalent
   (c) ionic
   (d) none of the above
8. Using only relative position in the periodic table, predict which combination of elements would have the greatest ΔEN. (2.3) K/U
   (a) As, Br
   (b) Rb, Cl
   (c) Ca, As
   (d) O, Se
9. Carbohydrate compounds contain the elements
   (a) carbon, hydrogen, oxygen, and nitrogen
   (b) carbon, hydrogen, oxygen, and sulfur
   (c) carbon, hydrogen, oxygen, and phosphorus
   (d) carbon, hydrogen, and oxygen (2.4) K/U
10. Which is a polyatomic ionic compound? (2.4) K/U
    (a) $NH_3(g)$
    (b) $Na_2SO_4(s)$
    (c) $HC_2H_3O_2(l)$
    (d) $C_6H_{12}O_6(s)$
11. Which is closest to the Canadian Sugar Institute's estimate of the average yearly consumption of table sugar by Canadians? (2.5) K/U
    (a) 2 kg per person
    (b) 11 kg per person
    (c) 16 kg per person
    (d) 23 kg per person

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

12. The equation below shows the formation of a cation.
    $\cdot\ddot{\underset{\cdot\cdot}{Br}}: + e^- \rightarrow \left[:\ddot{\underset{\cdot\cdot}{Br}}:\right]^-$ (2.1) K/U
13. Sodium chloride is a molecular compound. (2.1, 2.2) K/U
14. A compound that is waxy and soft is most likely molecular. (2.2) K/U
15. The majority of known compounds are molecular. (2.2) K/U
16. When oxygen atoms and hydrogen atoms bond together to form water molecules, electrons are transferred between the atoms. (2.2) K/U
17. In molecules such as O_2 and CO_2, the atoms share three pairs of electrons, thus forming double bonds. (2.2) K/U

18. In comparison to non-metallic elements, metallic elements tend to have lower electronegativities. (2.3) K/U
19. A compound formed between potassium and fluorine would have polar covalent bonding. (2.3) K/U
20. Molecular elements have non-polar covalent bonding. (2.2, 2.3) K/U
21. The correct name for the compound $(NH_4)_2SO_3$ is ammonium sulfate. (2.4) K/U

Match each term on the left with the most appropriate description on the right.

22. (a) no lone pairs — (i) N_2
 (b) one lone pair — (ii) CO_2
 (c) two lone pairs — (iii) BH_3
 (d) three lone pairs — (iv) HCN
 (e) four lone pairs — (v) HOCl
 (f) five lone pairs — (vi) Cl_2
 (g) six lone pairs — (vii) HCl (2.2) K/U

Write a short answer to each question.

23. Give an example of an electrolyte and explain why it is classified as an electrolyte. (2.1) K/U
24. Distinguish between the terms "anion" and "cation." (2.1) K/U
25. How does the shape of a sodium chloride crystal reflect the crystal lattice structure of this compound? (2.1) K/U
26. What is special about the molecular structure of the elements hydrogen, oxygen, nitrogen, fluorine, chlorine, bromine, and iodine? (2.2) K/U
27. If the electronegativity difference between two elements is 1.7, what kind of compound would they form? (2.3) K/U
28. Distinguish between traditional/common names for compounds and IUPAC names. (2.4) K/U
29. Describe the "zero-sum rule" for writing chemical formulas for ionic compounds. (2.4) K/U
30. State three adverse health consequences of consuming too much sugar. (2.5) K/U

Understanding

31. When mixed with water, do all ionic compounds produce solutions that are good conductors of electricity? Explain your answer. (2.1) T/I
32. **Figure 1** shows a model of a crystal structure. (2.1) K/U C
 (a) Sketch a model of the structure's formula unit.
 (b) Write a chemical formula to represent the structure in the form of X_nY_m, in which n and m are appropriate whole numbers.

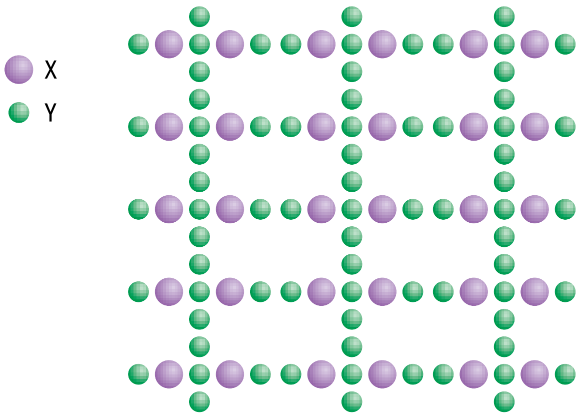


Figure 1

33. Draw Lewis symbols to illustrate the formation of the following ionic compounds: (2.1) K/U C
 (a) magnesium oxide
 (b) aluminum sulfide
 (c) lithium bromide
34. By gaining 2 electrons, a sulfur atom attains an electron arrangement like that of the noble gas argon. (2.1) K/U
 (a) Explain why the sulfur atom does not become an argon atom.
 (b) State what the sulfur atom does become when it gains 2 electrons.
35. Glucose, also known as blood sugar, has the molecular formula $C_6H_{12}O_6(s)$. Why can glucose not be represented with a formula unit of CH_2O? (2.2) K/U
36. Draw a Lewis structure and structural formula for each of the following molecules or ions: (2.2) K/U C
 (a) $Br_2(l)$
 (b) $OF_2(g)$
 (c) $NF_3(g)$
 (d) HCN(g)
 (e) $CS_2(l)$
 (f) $COCl_2(g)$
 (g) NO^+
 (h) PO_3^{3-}
 (i) $H_2S(g)$
 (j) $CCl_4(l)$
37. List the substances or ions in Question 36 that have double or triple bonds. (2.2) K/U
38. Which of the entities in Question 36 contain polar covalent bonds? (2.3) K/U
39. Explain why electronegativity increases as atomic radius decreases. (2.3) K/U

40. For each of the following pairs of elements, predict whether the bond between them will be non-polar covalent, polar covalent, or ionic: (2.3) K/U
    (a) N and H
    (b) Al and F
    (c) N and O
    (d) F and F
    (e) Br and K
41. Write the chemical formula for each of the following ionic compounds: (2.4) K/U C
    (a) copper(II) oxide (used to remediate hazardous waste)
    (b) aluminum nitrate (used to tan leather, found in antiperspirants)
    (c) manganese(II) chloride (used to make dry cell batteries)
    (d) barium fluoride (used to make certain lenses)
    (e) lead(IV) oxide (used to make matches)
    (f) iron(III) sulfate (used in pickling baths for aluminum and steel)
42. Write the chemical formula for each of the following molecular compounds: (2.4) K/U C
    (a) carbon disulfide (used as an insecticide and fumigant)
    (b) diarsenic trioxide (used as a wood preservative and in semiconductors)
    (c) dichlorine monoxide (used to make bleach)
    (d) diantimony pentoxide (used in flame retardants)
43. Write the IUPAC name for each of the following ionic compounds: (2.4) K/U
    (a) $SrS(s)$ (used in fireworks and luminous paints)
    (b) $(NH_4)_2SO_4(s)$ (used as a fertilizer for alkaline soils)
    (c) $SnF_2(s)$ (used in toothpastes)
    (d) $FePO_4(s)$ (used in fertilizers)
    (e) $Ca(OH)_2(s)$ (known as slaked lime)
    (f) $MgCO_3(s)$ (used as an antacid and to make table salt free-flowing)
44. Write the IUPAC name for each of the following molecular compounds: (2.4) K/U
    (a) $NF_3(g)$ (used as a cleaning agent in the manufacture of LCD displays)
    (b) $B_2O_3(s)$ (used to make borosilicate glass)
    (c) $I_2O_5(s)$ (used to determine the concentration of carbon monoxide in gas samples)
    (d) $BrO(g)$ (used as a fungicide and disinfectant)
45. Write the chemical formula for the hydrated compound iron(III) nitrate nonahydrate. (2.4) K/U C
46. Name the hydrated compound $ZnCl_2{\bullet}6H_2O(s)$. (2.4) K/U
47. How does the sugar content of a can of pop compare to the maximum quantity of added sugar that a person should consume in a day, according to the Heart and Stroke Foundation? (2.5) K/U A
48. Why does the soft-drink industry use mainly high fructose corn syrup as a sweetener instead of sucrose? (2.5) K/U

Analysis and Application

49. People routinely consume Aspirin and vitamin A, yet these two substances can be toxic, causing problems such as softening of bones, vomiting, seizures, and even death. Explain how a toxic substance can be consumed safely. (2.1) T/I A
50. What would you predict about the electrical conductivity of a typical sports energy drink? Explain your answer. (2.1) K/U A
51. Draw a diagram that shows how potassium bromide, $KBr(s)$, would dissolve in water. Explain your diagram. (2.1) K/U C
52. Predict which compound would have the higher melting point: sodium chloride, $NaCl(s)$, or aluminum oxide, $Al_2O_3(s)$. Explain your answer. (2.1) T/I
53. Sodium metal is used to obtain other reactive metals from their compounds and as a coolant in some nuclear reactors. Sodium, along with chlorine gas, can be prepared industrially by passing an electric current through sodium chloride placed in a special vessel called a Downs cell. The electric current converts sodium ions into sodium atoms and chloride ions into chlorine atoms. (2.1) K/U T/I C
    (a) In order for the electricity to flow in a Downs cell, in what state of matter must the sodium chloride be? Why?
    (b) Which ions gain electrons and which ions lose electrons in a Downs cell? Use Lewis symbols to illustrate the changes that the sodium and chloride ions undergo.
54. Imagine that you have been given a sample of an unknown solid compound. Explain how observations of melting point, hardness/brittleness, and electrical conductivity could help you to determine if the compound is ionic or molecular. (2.1, 2.2) K/U T/I
55. Element C is metallic and has 1 valence electron. Element D is non-metallic and has 6 valence electrons. (2.1, 2.2) K/U T/I C
    (a) What kind of compound will form between C and D? How do you know?
    (b) Using Lewis symbols, show how C and D will combine.
    (c) Write a chemical formula for the compound that will form.
    (d) What type of chemical formula did you write in (c)?

56. United Agri Products Canada Inc. is a very large supplier of crop protection chemicals and chemical nutrients for plants. One of their products is a herbicide called Mecoprop, which is a mixture of the two molecular forms shown in **Figure 2**. One of these forms actively kills weeds, but the other does not. (2.2) K/U T/I A

(a) What is the difference in the structure of the two molecular forms of Mecoprop?

(b) How might the herbicidal effects of Mecoprop be explained by a "lock and key" model?"

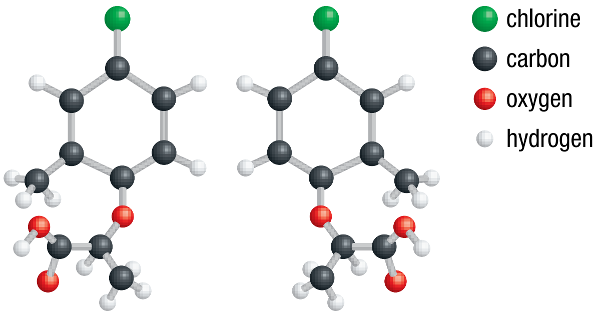


Figure 2

57. Nitrogen monoxide, NO(g), is emitted by internal combustion engines. It contributes to air pollution by participating in the formation of nitrogen dioxide, NO_2(g), and ozone, O_3(g). Nitrogen monoxide is both reactive and toxic. (2.2) K/U C A

(a) Explain why a Lewis structure that obeys the octet rule cannot be drawn for a nitric oxide molecule. Support your argument with two example structures.

(b) How do you think the molecular structure of nitrogen monoxide accounts for its high reactivity and toxicity?

58. One of the exceptions to the octet rule is that boron can form stable compounds with 6 electrons surrounding each boron atom. Boron compounds such as boron trifluoride, BF_3(g), are quite reactive (**Figure 3**).

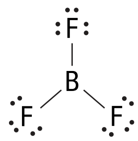


Figure 3

For example, a BF_3 molecule can share an electron pair on a fluoride ion to form the tetrafluoroborate ion, BF_4^-:

$BF_3 + F^- \rightarrow BF_4^-$ (2.2) K/U T/I C

(a) Draw a Lewis structure for the BF_4^- ion.

(b) Explain why the tetrafluoroborate ion is more stable than boron trifluoride.

59. Three different compounds have the molecular formula C_3H_8O (if no rings of atoms are allowed). Draw a structural formula for each compound. (2.2) K/U T/I C

60. Hydrazine, N_2H_4(l), is used as a corrosion inhibitor, as rocket fuel, and in the manufacture of certain foams. It may also be present in tobacco smoke. The Government of Canada is presently assessing the health hazards it poses due to its toxicity and potential to cause cancer. (2.2) K/U T/I C

(a) Are the hydrogen atoms likely to be connected to each other, in a hydrazine molecule?

(b) Draw a Lewis structure and structural formula for hydrazine.

(c) Hydrazine can react to form two different compounds with the molecular formula $C_2H_8N_2$. One of the compounds is a rocket fuel and the other is used to induce tumours in laboratory animals. Draw a Lewis structure and structural formula for each of these compounds. Note that the carbon atoms do not bond to one another.

61. Historically, chemists believed that compounds of the noble gases would never be prepared. How might the octet rule have played a part in this belief? Include a Lewis structure in your answer. (2.2) K/U T/I C

62. How is it possible for a compound to contain both ionic and covalent bonds? Give an example and draw the relevant Lewis structures. (2.1, 2.2) K/U T/I C

63. Oxygen atoms have 6 valence electrons, and beryllium atoms have 2 valence electrons. Yet, the two kinds of atom have the same bonding capacity. How do you account for this observation? Include examples of related Lewis structures in your answer. (2.2) K/U C

64. **Figure 4** shows a structural formula for cholesterol as a biochemist might draw it. Individuals with excessive levels of cholesterol in their blood are at risk of developing clogged arteries, which in turn can lead to heart disease and stroke. According to Statistics Canada, about 41 % of adult Canadians have high cholesterol levels. (2.2) K/U T/I A

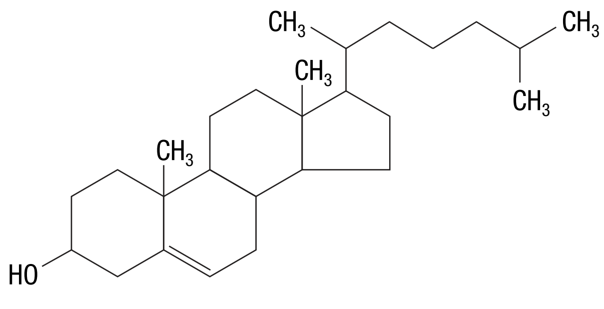


Figure 4

(a) Examine Figure 4 and then write the molecular formula for cholesterol. Note that hydrogen atoms bonded to carbon atoms are not shown in this diagram, and that all of the carbon atoms must form four bonds.

(b) Cholesterol is a waxy solid at room temperature. Is this a surprising observation? Why or why not?

(c) What is the advantage of representing molecules in a form like that in Figure 4?

65. Using Lewis structures as a guide, predict the molecules that each pair of elements below would combine to form. In each case, give your answer in the form of a Lewis structure and a structural formula. Assume that there are no exceptions to the octet rule. (2.2) K/U T/I C
   (a) sulfur and chlorine
   (b) arsenic and fluorine
   (c) germanium and chlorine
   (d) chlorine and fluorine
66. Pauling electronegativity values are not available for helium, neon, and argon. Suggest why. (2.3) K/U T/I
67. Offer reasons for why the trends in electronegativity in the periodic table parallel the trends in first electron affinity (Section 1.7). (2.3) K/U T/I
68. In each case below, predict the type of bond that would be observed. Cite quantitative evidence to support each answer. (2.3) T/I C
   (a) $H_2S(g)$ (gas responsible for rotten egg smell)
   (b) $CaF_2(s)$ (used to make special camera and telescope lenses)
   (c) $SF_6(l)$ (used as an insulator in electric transformers)
   (d) $C_6H_{12}O_6(s)$ (glucose or blood sugar)
69. For each of the pairs of entities listed below, sketch how the bonding electrons would be distributed between the atoms. Use electronegativity difference as a guide, writing the value next to each sketch. Also include the symbols δ^+ and δ^- as appropriate. Look at Figures 3 and 4 in Section 2.3 to see a model for your sketches. Finally, write the type of bond that forms beside each sketch. (2.3) T/I C
   (a) a silicon atom and a chlorine atom
   (b) 2 oxygen atoms
   (c) a potassium ion and a chloride ion
70. Of nitrogen, $N_2(g)$, hydrogen fluoride, HF(g), and hydrogen bromide, HBr(g), which would have bonds with the greatest δ^+ and δ^- charges? Explain. (2.3) T/I
71. Copper(II) sulfate pentahydrate, $CuSO_4 \cdot 5H_2O(s)$, is often used as a fungicide in agriculture and as an electrolyte in copper refining. Explain what will happen to the colour when copper(II) sulfate pentahydrate is strongly heated. The chemical equation below should help you solve this question.

$$CuSO_4 \cdot 5H_2O(s) \rightarrow CuSO_4(s) + 5\ H_2O(g)$$ (2.4) T/I

Evaluation

72. A friend has read that touching tarnished silver to aluminum foil immersed in very hot water can remove the tarnish. However, the water must be able to conduct electricity for the reaction to take place. Your friend is considering dissolving either table sugar, $C_{12}H_{22}O_{11}(s)$, or washing soda, $Na_2CO_3(s)$, in the water to facilitate the tarnish removal method. Which compound would you recommend? Why? (2.1, 2.2) K/U T/I A
73. Amino acids are the building blocks of the proteins in our bodies, linking together in large numbers to create long, chain-like molecules. Our bodies can synthesize some amino acids. Others, however, can only be obtained from our diet. The molecular formula of an amino acid called glycine is $C_2H_5NO_2$. Several proposed Lewis structures for glycine are shown in **Figure 5**. Examine these structures and select the best model for a glycine molecule. Explain your answer, including reasons for rejecting the other structures. (2.2) K/U A T/I A

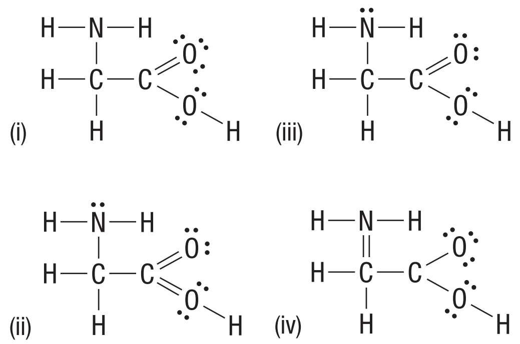
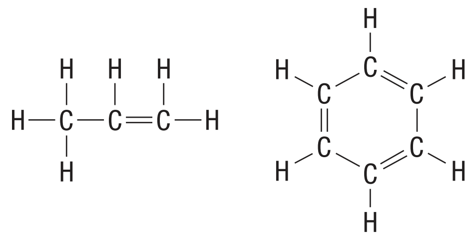


Figure 5

74. **Figure 6** shows possible Lewis structures for propene and benzene. **Table 1** shows experimental data for the lengths of the carbon–carbon bonds in these two compounds as well as the typical lengths of carbon–carbon bonds. (2.2) T/I C A



propene benzene

Figure 6

Table 1 Carbon–Carbon Bond Lengths

Propene (pm)	Benzene (pm)	Typical C–C single bond (pm)	Typical C–C double bond (pm)
134; 151	140	154	134

   (a) Evaluate how well the Lewis structures for propene and benzene agree with the experimental evidence.

(b) Do the carbon atoms in both Lewis structures in Figure 6 satisfy the octet rule? Explain your answer.

(c) Considering your answers to (a) and (b), how useful is the octet rule in predicting the molecular structures of propene and benzene?

75. A chemist is interested in synthesizing molecular compounds composed of noble gas elements combined with other elements. She knows that the most promising approach is to choose reactants that will result in the greatest possible movement of electrons on noble gas atoms toward other atoms as bonds form. What elements would you suggest she use in her research? Explain your choices using specific atomic properties and a description of the type of bond the chemist hopes will form. (2.3) K/U T/I

76. Compounds of a metallic element combined with a non-metallic element are generally ionic. However, some compounds containing a metallic element and a non-metallic element have some characteristics of molecular compounds. A chemist wants to test whether or not these compounds are really more molecular than ionic. The chemist plans to dissolve them in water and measure the electrical conductivity of the resulting solution. (2.1, 2.2, 2.3) K/U T/I

(a) Explain the thinking behind the chemist's plan.

(b) What flaw is there in the chemist's proposed method?

77. A company ships electronic components that must be kept free of moisture during transit. A clerk proposes putting small, porous containers containing calcium chloride dihydrate, $CaCl_2 \cdot 2H_2O(s)$, in the packages to serve as a drying agent. Do you agree with this idea? If so, explain why. If not, suggest an alternative and explain your suggestion. (2.4) K/U A

Reflect on Your Learning

78. Which skill or concept in this chapter was the most difficult for you? Outline a plan for improving your understanding. K/U A

79. It seems that at every level in chemistry, exceptions or oddities remain. Do you believe that, given enough time, science will eventually reach a point where we have an explanation for everything, or is such an outcome impossible to realize? Explain your answer. A

80. Some scientists pursue curiosity-driven research that focuses on "pure" scientific topics. Even so, their discoveries may have very beneficial practical applications. If you were to choose a scientific career, would you rather do pure research or applied research? Explain why. A

81. Considering what you have learned in this chapter about artificial sweeteners and excess sugar consumption, will you alter your diet in any way? Why or why not? A

Research

GO TO NELSON SCIENCE

82. The type of crystal formed by sodium chloride is one of several different crystal forms. Scientists classify crystals into seven systems, each with a characteristic shape. Research the seven crystal systems to find out about their names, symmetries, and ways in which the particles cluster together. Also, find an ionic compound or mineral that illustrates each crystal system. Report your findings in a slide presentation, poster, or website. Include diagrams and images. K/U C A

83. Water molecules assemble into a molecular crystal, ice, when they freeze. Research the crystal lattice of ice. Prepare a diagram or model of the ice structure that exists under ordinary conditions. Use what you have learned to explain why ice displays the unusual property of being less dense than the liquid from which it forms and to explain why water expands as it freezes, breaking bottles and cracking paved streets. K/U C A

84. In many cases, several Lewis structures can be drawn for the same molecule. Chemists can use a concept called formal charge to decide which structure is most likely accurate. Research this concept and report how it is determined and how it helps to weed out unlikely Lewis structures. Give a specific example, showing both the unlikely and likely Lewis structures. T/I C

85. Find out more about how Linus Pauling determined electronegativity values. In addition, compare and contrast his methods and values to those of Robert S. Mulliken and the team of A.L. Allred and E.G. Rochow. Mulliken published an electronegativity scale in 1934, and Allred and Rochow published a scale in 1958. Present your findings in a poster. T/I C

86. A scientist named Gilbert Levin has developed a relatively new low-calorie sweetener, D-tagatose, that is now appearing in selected food products. D-tagatose is sometimes referred to as a "left-handed" sugar. Research Levin's work and left-handed sugars. Report the details of the story of tagatose and the reasoning behind the idea of using "left-handed" sugars as low-calorie sweeteners. Provide examples of structural formulas of relevant sugar molecules. In your report, include the economics of mass-producing such sugars, and the accident that led Gilbert to tagatose. K/U C A

CHAPTER 3

Molecular Compounds and Intermolecular Forces

After completing this chapter you will be able to

- analyze the properties of commonly used chemical substances, evaluate their effects on human health and the environment, and propose ways to lessen their impact
- compare and contrast the physical properties of different molecular compounds
- build molecular models and write structural formulas for molecular compounds
- relate the presence of intermolecular forces to the physical properties of molecular compounds, including water
- compare intermolecular forces to covalent bonds and to ionic bonds
- predict the polar nature of covalent bonds using electronegativity differences
- predict the polar nature of simple molecular compounds
- safely conduct laboratory investigations into the properties of molecular compounds

How Can We Use Our Knowledge of the Physical Properties of Molecular Compounds?

Most compounds in our world today are molecular. Think about what is in your school bag. Your plastic lunch bag is made of long molecules that are very resistant to decomposition. This is good for keeping your lunch separate from your school books, but not so good when you want to dispose of the lunch bag—it will probably remain in the landfill for centuries. The raw materials for making plastic come from non-renewable petroleum. The manufacturing process uses a lot of energy and the by-products can be hazardous. In comparison, the paper in your school books is made from cellulose fibres. Cellulose usually comes from renewable tree pulp. It degrades relatively quickly under the right conditions. Alternatively, we can recycle most types of paper.

Environmental concerns and consumer awareness are increasing. Industries are thinking more about both the composition of their products and the chemical processes that produce them.

"Dry cleaning" is a chemical process that does indeed clean our clothes, but it is neither dry nor clean. Most commercial dry cleaners use a liquid called perchloroethylene, or "perc," C_2Cl_4. Perc is very effective at removing grease and oil, but is not good for our health or the environment. Exposure to perc is linked to many health effects. Improper disposal of perc leads to groundwater contamination, release of greenhouse gases, and other environmental problems. Is there an alternative?

Innovative cleaning companies are using a greener solution: liquid carbon dioxide. As you know, carbon dioxide gas in the atmosphere is a hot topic these days. It is a greenhouse gas. Carbon dioxide changes from a gas to a solid at $-78\ °C$, and then sublimates back to a gas under normal atmospheric conditions. You will probably never see liquid carbon dioxide; it does not exist at the temperatures and pressures of ordinary life. Carbon dioxide gas must be compressed to five times normal pressure to make it a suitable liquid for the dry-cleaning industry to use.

Clothes are placed in large machines, the air is removed, and liquid carbon dioxide is pumped in. The grease and dirt dissolve in the liquid. Spinning and rinsing with more liquid carbon dioxide leaves the clothes fresh and clean.

Just what makes perc and liquid carbon dioxide so good at removing greasy stains from clothes? How does a molecular compound's structure influence its properties? In this chapter, we will explore the answers to these questions.

STARTING **POINTS**

Answer the following questions using your current knowledge. You will have a chance to revisit these questions later, applying concepts and skills from the chapter.

1. Sometimes we discover that products that we have been using for years are dangerous to people or the environment. What might chemists do to respond to such a discovery?
2. Do consumers know everything they should about a product or a chemical process? Justify your response.
3. New green technologies are not always accepted quickly. Suggest reasons why people may not quickly switch to using a new, more environmentally friendly technology.
4. Is there any point in being able to predict the physical properties of a molecular compound? Could this ability be of value to industries and to society? Explain.
5. What do you think is the future of liquid carbon dioxide dry cleaning? Will it become more popular?
6. What measures could help green companies to succeed?

Mini Investigation

Exploring the Properties of Oobleck

Skills: Questioning, Performing, Observing, Analyzing, Communicating

SKILLS HANDBOOK A1.2

In this investigation, you will explore a mixture of cornstarch and water. This mixture is sometimes called "oobleck," from a children's book by Dr. Seuss: *Bartholomew and the Oobleck*. Cornstarch and water are both molecular compounds. When they are mixed together their properties are rather surprising!

Equipment and Materials: newspaper; 250 mL beaker; 600 mL beaker; 50 mL cornstarch; food colouring (optional), water

1. Place a piece of newspaper on your lab bench.
2. Measure about 50 mL of cornstarch in the 250 mL beaker. Transfer the cornstarch to the 600 mL beaker.
3. Measure about 50 mL of water in the 250 mL beaker.
4. Add water to the cornstarch in the larger beaker, while mixing, until the mixture has the consistency of pancake batter.
5. Slowly stick your finger into the mixture and then remove it. Repeat this test, this time moving very quickly. Record your observations.
6. Scoop some of the mixture up into your hands. Squeeze it and then let it go. Record your observations.
7. Follow your teacher's directions for disposal and wash your hands thoroughly.

A. List the properties of this mixture. T/I

B. Propose some explanations for the strange properties of this mixture. T/I

C. Brainstorm three controlled experiments that you could perform on this mixture. For each experiment, list the independent and dependent variables. T/I

D. Suggest some of the challenges of explaining what is going on at a molecular level by observing the properties of a substance. T/I

3.1 Molecular Compounds

Canada is one of the wealthiest nations on Earth. Most Canadians use and consume an incredibly large number of consumer products. This is not necessarily something to be proud of. Fortunately, producers and consumers are taking steps to live in a more sustainable way. As a consumer, you most likely make efforts to minimize your use of materials. For example, you might use a reusable metal container for drinking water instead of using many plastic water bottles (**Figure 1**). You probably also recycle some waste products, such as paper, aluminum cans, and plastic bottles. We know that the things we use affect the world around us. We are trying to change the way we behave and the way we think about consumer products (**Figure 2**). In this section we will consider the production, use, and disposal of molecular compounds to better understand their effects on society and the environment.

Figure 1 Some products are designed to be used many times.

renewable resource a natural substance that replenishes itself as it is used

non-renewable resource a natural resource that cannot be replaced as quickly as it is consumed

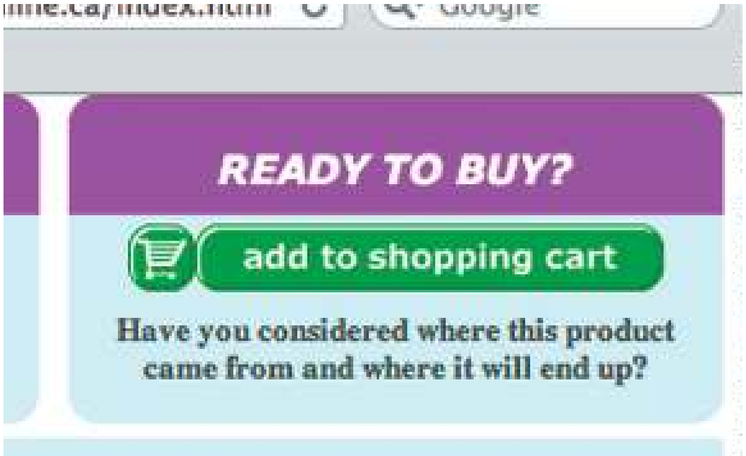


Figure 2 It is important for consumers and manufacturers to look at the production, use, and eventual disposal of a consumer product when considering its purchase.

The Source of Molecular Compounds

The molecular compounds in our consumer products originally came from natural resources. Natural resources are classified as either renewable or non-renewable. A **renewable resource**, such as wood or water, is a resource that replenishes itself as it is used. A **non-renewable resource**, such as crude oil or aluminum ore, is a resource that is not replaced as quickly as we use it. The classification depends on whether or not they continue to be formed as we use them.

Many molecular compounds originate from living things. Organisms produce a huge variety of compounds, many of which are useful to humans. As a tree grows, it continually produces a variety of complex and valuable materials. These include the wood at its core and the various compounds in its bark, fruit, and leaves. Micro-organisms, such as fungi and bacteria, also produce extremely useful compounds. We use these natural substances for food, fuel, and medicines, and as raw materials for other products. If we manage a renewable resource sustainably, it can benefit society indefinitely. If these resources are consumed in a non-sustainable fashion, however, they will be depleted and, ultimately, exhausted.

WEB LINK

To watch videos about the effects of using up non-renewable resources,

GO TO NELSON SCIENCE

Fossil fuels are another large source of molecular compounds. As you probably know, fossil fuels are chemicals that form naturally, over millions of years, from partially decayed remains of ancient plants and animals. Since we are consuming fossil fuels much faster than they are being created, we classify them as non-renewable resources. We can obtain some molecular compounds from renewable resources, such as corn and sugar cane. Most petrochemicals, however, such as plastic, gasoline, and paints, are produced from fossil fuels. A **petrochemical** is a compound that is synthesized by chemically changing certain components of fossil fuels. Petrochemicals are non-renewable resources.

petrochemical a compound manufactured from a fossil fuel

Petrochemicals are everywhere! Natural gas processing plants, oil refineries, and other petrochemical facilities produce a great variety of these molecular compounds. Indeed, it is hard to look anywhere without seeing something that was produced from fossil fuels (**Figure 3**). When we buy cellphones, clothes, or school supplies, we should remember that many molecular compounds modified from fossil fuels are used in these products. It is important that, as consumers, we consider the sources of the raw materials. Synthesizing products from petrochemicals uses a great deal of energy, so we must also consider the fuels used in these processes.

CAREER **LINK**

To find out more about the career of a petrochemical engineer,

GO TO NELSON SCIENCE

Choosing the Right Materials for a Product

When product designers develop a new product, they have to consider what chemical and physical properties the product must have to serve its purpose. They should also consider what will happen to the product at the end of its useful life.

Consider some of the compounds involved in a simple consumer product—a pencil case. A typical pencil case is made of plastic (**Figure 4(a)**). Plastics are molecular compounds composed of very long molecules. They are relatively inexpensive, easy to form into different shapes, waterproof, versatile, and long-lasting. These beneficial properties explain why plastics are used to make so many consumer products. Most plastics are petrochemicals, made from non-renewable resources.

Depending on what type of plastic this pencil case is made of, it may or may not be recyclable. **Recycling** involves reprocessing a product to make other products. Plastics may be recycled into park benches or plastic lumber. However, this plastic pencil case may end up in a landfill for centuries. Many landfills fail to contain the waste, resulting in soil and water contamination.

What are the alternatives to a standard plastic pencil case? The pencil case in **Figure 4(b)** is strong, durable, waterproof, and, in addition, advertised as biodegradable. A **biodegradable** product is one that can be broken down by micro-organisms, using oxygen and water. The common products are carbon dioxide, water, and organic matter. Manufactured products based on natural materials (such as paper and some soaps) are often biodegradable. The biodegradable pencil case is made of a special plastic called polylactide, more commonly known as polylactic acid (PLA). This type of plastic is produced from renewable resources such as corn and other starch-rich foods. The molecules are modified so that they join together to form long molecules. This changes the properties of the compound, making it waterproof and longer-lasting. Substances that are made of PLA are not recyclable but they are biodegradable.

PLA is used for pharmaceutical, industrial, and biomedical products, among other things. You might have seen disposable food containers or cups made from PLA.

Finally, **Figure 4(c)** shows a pencil case made from organically grown cotton, a natural fibre produced by plants raised without synthetic pesticides or fertilizers. Cotton is a renewable resource. This pencil case is not waterproof, and may not be quite as durable as plastic or PLA options.

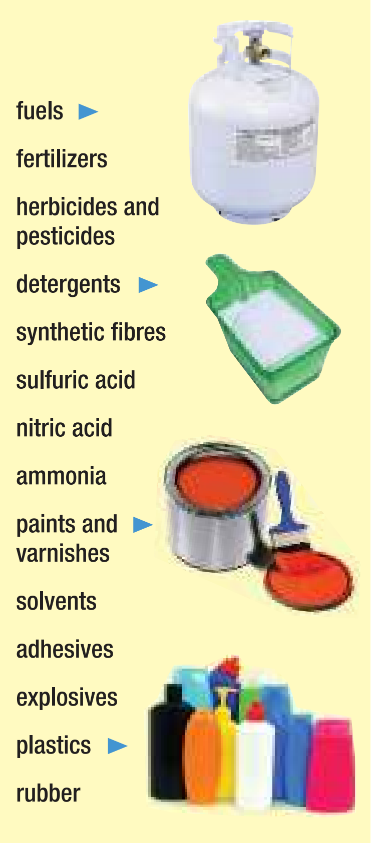


Figure 3 Fossil fuels such as crude oil and natural gas are the raw materials for a vast range of chemical products.

recycling converting a product back into material that can be used to make new goods

biodegradable capable of being broken down (decomposed) rapidly by the action of moisture, heat, and micro-organisms

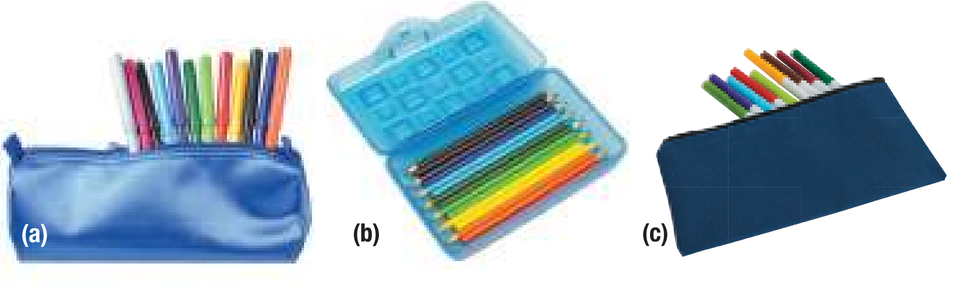


Figure 4 A pencil case could be made from (a) a combination of non-recyclable plastics, (b) biodegradable plastic, or (c) compostable organic cotton.

Like the PLA pencil case, the cotton pencil case is biodegradable. Cotton, like many natural products, can be composted. A material or product is considered **compostable** if it decomposes in a composting site within a set time frame. Food scraps, cotton, wool, and wood are all compostable. Industrial composting sites provide ideal conditions for composting. These conditions include an appropriate mix of waste, exposure to oxygen and water, and the right temperature. The products of composting sustain plant life. All compostable materials are biodegradable.

compostable the ability of a material to decompose naturally, resulting in a product that is able to sustain plant life

The End Game—Reusing, Recycling, and Composting

When we discard a used consumer product, we quickly forget about it. But what happens to it next? It is certainly not gone; it still exists somewhere. Product designers and process engineers must consider the end-of-life or final destination of the materials used in a consumer product.

The chemical bonds that make plastics so durable also make most plastics resistant to any natural decomposition process. These are synthetic compounds and no micro-organisms have evolved to consume them. Therefore, when your pencil case is damaged or unusable, what choice do you have but to throw it away? The environmental effect of our plastic waste is a serious concern. In the past 50 years, over a billion tonnes of plastic has been sent to landfill sites. There, the plastic items will slowly degrade over hundreds or thousands of years.

Plastic garbage is also accumulating at an alarming rate in our waterways, lakes, and oceans. The Great Pacific Garbage Patch in the central North Pacific is vast. Scientists have variously estimated its area as being half the size of Ontario and twice the size of the continental United States. Whatever its actual area, it is an environmental embarrassment.

WEB LINK

The Great Pacific Garbage Patch has formed where ocean currents collect plastic garbage. To see pictures and videos of this astonishing floating garbage dump,

GO TO NELSON SCIENCE

When we consider whether a product is "green," there are many things to bear in mind. One consideration is whether or not the product is reusable, recyclable, or biodegradable/compostable.

Reusable products are generally the most environmentally friendly. It is usually better to use a single item many times than to use many different items only once each. Your reusable water bottle, for example, avoids the need to produce, transport, and dispose of many plastic water bottles.

The process of recycling materials has gained popularity over the past few decades. In many countries it is now common practice to recycle glass, metal, plastic, and paper products. For a material to be recyclable, we must be able to convert it back into material that can be used in manufacturing new goods. For example, a glass bottle must be melted down with thousands of similar glass bottles and made into another glass product. Typically, recyclable materials must remain in their pure form. If too many additional materials are combined with the product (such as adhesives, labels, and paints), it is too difficult or costly for a recycling process to separate and remove these additional materials.

upcycling the process of converting an industrial material or product into something of similar or greater value

Unfortunately, a lot of plastics cannot be recycled. This is because manufacturers use various additives, such as dyes, softeners, and ultraviolet (UV) inhibitors, to give the plastic specific properties. Since they are unsuitable for recycling, most of these plastics end up in the landfill. Sometimes, however, they can be given a totally new life. **Upcycling** involves converting wastes into high-value new products (**Figure 5**).

Figure 5 A stylish bag, produced from used candy pouches, helps to divert garbage from landfill sites.

What about our biodegradable pencil case? Certain plastics incorporate degrading agents so that they, too, will break down eventually. Unfortunately, the definition of "biodegradable" does not specify the time frame or the necessary conditions. "Biodegradable" items may sit in a landfill for 20, 200, or even 2000 years if they do not have maximum exposure to moisture, oxygen, heat, and micro-organisms. Sometimes we cannot be sure how "green" a biodegradable product really is.

Finally, what will happen to our compostable pencil case? It will likely decompose fairly quickly into harmless—even beneficial—products.

3.1 Summary

- Many consumer products are made of molecular compounds.
- Consumer products are produced from raw materials that can be classified as either renewable or non-renewable.
- Renewable resources can be replaced in a sustainable way whereas non-renewable resources cannot.
- Most synthetic molecular compounds are produced from fossil fuels.
- It is important to consider the source of a material, the properties and function of a material, and the destination of a material when designing consumer products.
- Consumer products end up as waste or they are recycled, upcycled, biodegraded, or composted.

UNIT TASK BOOKMARK

Consider the entire lifespan of your product, from manufacture to disposal, as you work on the Unit Task outlined on page 134.

3.1 Questions

1. Give two examples of renewable resources and two examples of non-renewable resources. K/U
2. List three specific physical properties that each of the following products should have: T/I
   (a) the canopy material for an umbrella
   (b) protective packing for electronics equipment
   (c) a bathroom cleaning product
   (d) a DVD case
   (e) a juice container
3. Compare and contrast the terms "biodegradable" and "compostable." K/U
4. Choose five consumer products in your kitchen at home. List the various components of each product, including packaging. In a small group, brainstorm where the materials for each one's production came from and where each product will end up after its useful life. Organize your thoughts in a table or graphic organizer. T/I C
5. Create a table and list six examples of petrochemical consumer products. For each product, identify the end stage of its life. T/I C
6. Research to find five different "upcycled" consumer products. T/I A
7. North Americans discard millions of rubber automobile tires each year. These tires can end up in tire dumps where they take years to decompose, or they can be reused for different purposes. Research new initiatives for the development of new "green" car tires that decompose quickly, and alternative uses for old tires. T/I A
8. Research the Great Pacific Garbage Patch. What is it? What effect does it have on marine wildlife? Can anything be done to make it go away? Assemble your discoveries into a public information campaign. T/I C
9. Industrial composting systems are becoming attractive alternatives to landfill sites. One of the largest composting facilities in the world is the Edmonton Composting Facility, which helps to divert 60 % of household waste from landfill sites (**Figure 6**). The primary building is one of the largest structures in North America—larger than 14 hockey rinks. Research this facility and identify what strategies are used to accelerate the composting timeline. T/I A

Figure 6 A worker at the Edmonton Composting Facility helps to keep the optimal conditions for composting.

GO TO NELSON SCIENCE

Explore an Issue in Molecular Compounds

SKILLS MENU

- Defining the Issue
- Researching
- Identifying Alternatives
- Analyzing
- Defending a Decision
- Communicating
- Evaluating

Figure 1 A mosquito sucking blood from a human arm. Mosquitoes transmit several diseases, including West Nile virus and malaria.

CAREER LINK

Product testing technicians are involved in making sure that consumer products meet a range of guidelines and regulations. To find out more about this career,

GO TO NELSON SCIENCE

DEET in Insect Repellents

Imagine that you are sitting on the dock by a lake enjoying a beautiful Ontario summer sunset. Could life get much better? Except for . . . bzzzzz . . . smack . . . got 'im!

The existence, and persistence, of mosquitoes in summer can be irritating and itchy (**Figure 1**). The issues around mosquito bites go well beyond annoyance, however. The first human cases of West Nile virus in Ontario were confirmed in 2002. Since then, people have become more aware and vigilant about avoiding mosquito bites. There are several ways to do this. You might avoid being outside at dawn and dusk or you might wear protective clothing. The most popular strategy, however, is to use insect repellents to ward off the annoying mosquitoes. (Note that repellents are different from insecticides, which kill insects outright.)

Several brands of insect repellents are available in Ontario (**Figure 2(a)**). These repellents either mask the scent of a human or release a scent that insects avoid. There are many natural insect repellents such as citronella oil, garlic, geranium oil, and peppermint. Research has shown, however, that the synthetic compound known as DEET is particularly effective (**Figure 2(b)**).

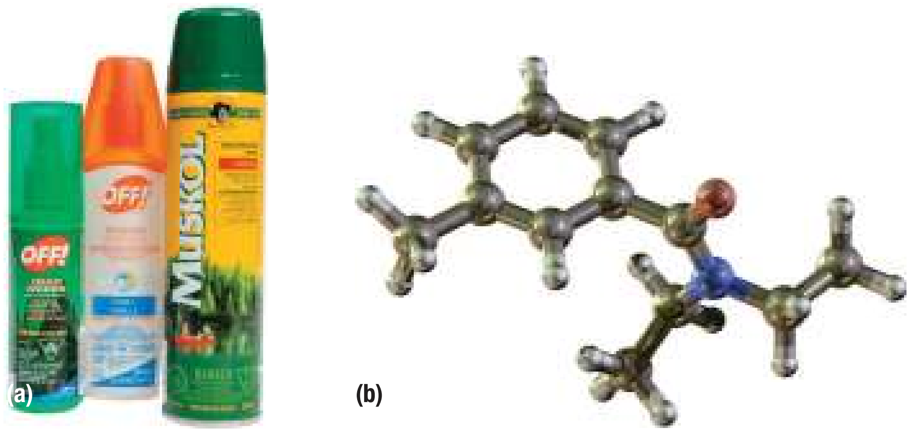


Figure 2 (a) A wide variety of insect repellent products is available. (b) A molecule of DEET

DEET (*N,N*-Diethyl-3-methylbenzamide) was developed by the United States Army to protect soldiers fighting in mosquito-infested areas. It entered commercial use as an insect repellent in 1957. DEET is a slightly yellow, oily substance that can damage plastic materials, various fabrics, and painted surfaces. This damage occurs because the materials partially dissolve in the DEET. As you explore solubility in Unit 4, you will learn more about why this occurs. Because of this property, DEET-based repellents should be kept away from eyeglass frames, synthetic clothing, and other molecular substances.

The concentration of DEET in most products ranges from around 5 % to 30 %. Repellents come in a variety of forms, such as solutions, foams, gels, and sprays.

Millions of people have used DEET over the past 50 years. Nevertheless, there is some controversy surrounding its safety. DEET may be the cause of a small number of seizures in adults.

The Canadian Pediatric Society recommends that repellents for children not exceed DEET concentrations of 10 % and that DEET-based products not be used for infants under six months of age. There is also controversy surrounding the use of DEET by pregnant or breastfeeding women.

The Issue

Your group owns a family resort in northern Ontario. You must decide what types of insect repellents to sell in your general store.

Goal

To decide on the best types of insect repellent to sell in your general store

Research

SKILLS HANDBOOK A5.1

Work in a small group to learn more about the available products and the issues (health concerns, cost, public acceptance, and so on) that surround each one.

- What is the common name of the insect repellent compound in each product?
- What is the concentration of the insect repellent compound?
- How does the active ingredient in each product repel mosquitoes?
- What are the health and environmental concerns?
- What are the marketing concerns (such as customer demographics, cost, and perception)?

As you research, consider your sources of information. Are they trustworthy and unbiased?

WEB LINK

To learn more about the variety of insect repellent products currently available, and their active ingredients,

GO TO NELSON SCIENCE

Identify Solutions

- Set selection criteria for the products you are considering stocking. These criteria could include active ingredient (repellent compound), concentration, cost, health concerns, environmental impact, and customer acceptance. Include other criteria also.
- Develop a rating system for your criteria.
- Based on your research, evaluate the different alternatives according to your selection criteria.

Make a Decision

Which insect repellent products will be sold in your general store? Outline the rationale you used to reach your decision. Be prepared to defend your decision and outline any consequences associated with your choice.

UNIT TASK BOOKMARK

The effectiveness of any product is important. You will want to consider the effectiveness of your green product in the Unit Task (page 134).

Communicate

Develop an information pamphlet, or a web page for your resort's website, that will justify your choices. The information pamphlet will be available to customers in your store. Remember that most customers have little knowledge of chemistry, so you will need to explain your facts and arguments in simple terms.

Plan for Action

Now that you have found out all about insect repellents, which one would you choose for your own personal use? List the criteria that will help you decide. You might consider the active ingredient, researched effectiveness, safety, cost, and so on.

Select three or four products (either commercially available or homemade) and rank them according to your criteria. Which one comes out on top?

GO TO NELSON SCIENCE

3.3 Polar Bonds and Polar Molecules

Table 1 Summary of Chemical Bonds

Ionic bonds	Covalent bonds
form between ions	form between atoms
are created by transfer of valence electrons between atoms to form ions	are created by sharing of pairs of valence electrons between atoms
result in large crystal lattices	result in individual molecules

LEARNING TIP

Reviewing Previous Material
When you see a reference to earlier material, it is a good idea to go back and reread it. In particular, make sure that you understand important definitions. Relevant vocabulary here includes *polar covalent bond; non-polar covalent bond;* and *electronegativity difference,* ΔEN.

In Chapter 2 we focused on the chemical bonds (forces of attraction) between ions or atoms in chemical compounds (**Table 1**). In this section, we will review polar bonds and explore how they influence the properties of a substance.

Bond Polarity

In Section 2.3 you learned about bond polarity: whether a covalent bond is polar or non-polar. Atoms with significantly different electronegativities form polar covalent bonds. In this case, there is unequal sharing of the bonding electron pair. Electrons are in continual motion. In a polar covalent bond the shared electrons tend to be clustered closer to the strongly electronegative atom than to the other atom. The "electron density" is greater around the more electronegative atom. This end of the molecule therefore has a negatively charged end or "pole." The other end of the molecule, where the electron density is less, is the positive pole (**Figure 1**).

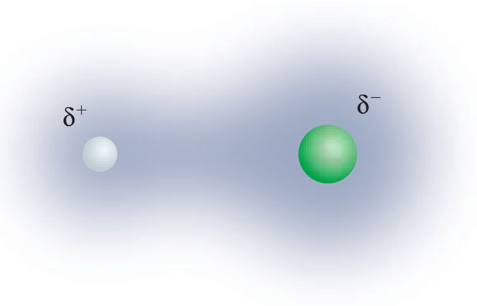
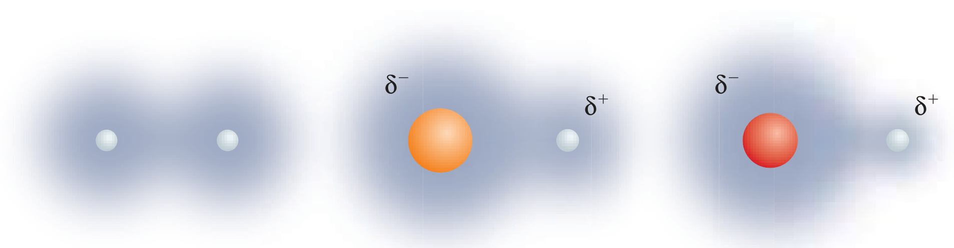


Figure 1 When two atoms are joined by a polar covalent bond, the negative pole, where the electron density is greater, is indicated by "δ^-." The positive pole, where the electron density is less, is indicated by "δ^+." The grey shading represents the electron density.

Look again at Figure 5 in Section 2.3 to see the spectrum of bonds, from non-polar covalent through ionic. You can see that the bond in a molecule of hydrogen chloride, HCl ($\Delta EN = 1.0$), is classified as polar covalent.

The extent to which a bond is polar depends directly on the electronegativity difference. For example, consider the covalent bonds that form between the following: hydrogen and hydrogen, nitrogen and hydrogen, and oxygen and hydrogen (**Figure 2**). See how these bonds would fit on the bonding continuum in Section 2.3.



$\Delta EN = 0.0$
(a) H–H, a non-polar covalent bond

$\Delta EN = 0.9$
(b) N–H, a polar covalent bond

$\Delta EN = 1.4$
(c) O–H, a very polar covalent bond

Figure 2 A covalent bond's polar nature depends on the electronegativity difference between the bonding atoms.

A Phenomenon of Polar Liquids

Scientists have discovered that different liquids behave in rather interesting ways when poured, in a thin stream, near a charged object. The following Mini Investigation lets you observe this phenomenon.

Mini Investigation

Evidence for Polar Molecules

Skills: Performing, Observing, Analyzing, Communicating

SKILLS HANDBOOK A1.2, A2.2, A3

In this investigation you will observe the effects of charged objects on thin streams of various liquids (**Figure 3**). You will look for differences in behaviour between polar and non-polar liquids.

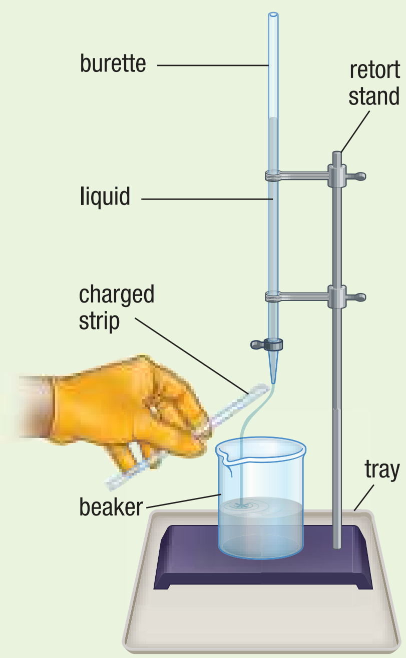


Figure 3 Testing a liquid with a charged strip

Equipment and Materials: chemical safety goggles; lab apron; nitrile gloves; tray; 50 mL glass burette clamped to a stand; small funnel; 400 mL beaker; acetate strip; vinyl strip or balloon; small covered bottle containing water, ethanol, or propanone (acetone); paper towel

Some of the liquids used in this investigation are flammable. They should be used only in a well-ventilated area. There should be no open flames or sparks in the laboratory.

1. Copy **Table 2** into your notebook, adding one more column for your observations.
2. Put on your chemical safety goggles, lab apron, and gloves.
3. Your group will be assigned one of the liquids. Using a funnel, fill the burette with your assigned liquid and set it up securely on the tray.
4. Rub the acetate strip back and forth several times in a piece of paper towel. Paper attracts electrons away from the acetate strip. The acetate will therefore acquire a positive charge.
5. Allow a stream of the liquid to pour into the beaker below.
6. As the liquid is running, hold the charged acetate strip close to the stream of water. Observe the stream closely. Record your observations.
7. Repeat Steps 3 to 6 with a charged vinyl strip or inflated balloon. Vinyl has a greater tendency to gain electrons when rubbed. It therefore acquires a negative charge.
8. Observe the other groups testing the other liquids and examine **Figure 4**.

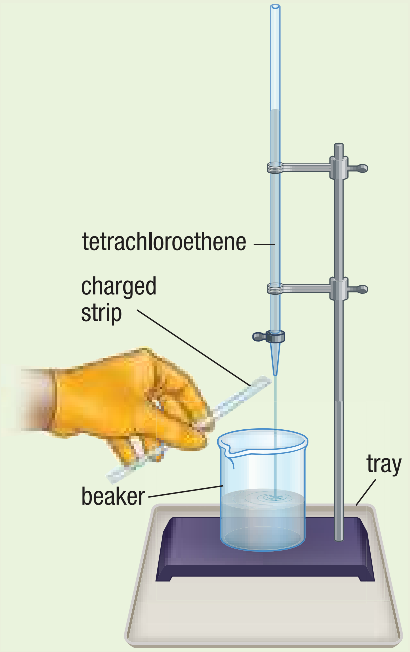


Figure 4 Tetrachloroethene flowing past an acetate strip

9. Empty your burette into a beaker and, using the funnel, return the liquid to its small bottle. Hand the bottle to your teacher.

Table 2 The Effect of Electric Charge on Molecular Liquids

Liquid	Chemical formula	Polar or non-polar
water	$H_2O(l)$	polar
ethanol	$C_2H_5OH(l)$	polar
propanone (acetone)	$CH_3COCH_3(l)$	polar
tetrachloroethene	$C_2Cl_4(l)$	non-polar

A. Is there any connection between the behaviour of the liquids and whether they are polar or non-polar? If so, describe the relationship. T/I
B. What difference occurred when you used the acetate strip compared to the vinyl strip? Did you expect this result? T/I
C. Try to propose an explanation for your observations. T/I
D. Predict what you would observe if you held a vinyl strip beside a stream of hexane (a non-polar liquid). T/I

Molecular Polarity

Look closely at the results from the Mini Investigation: Evidence for Polar Molecules. Is there a connection between the shape of a molecule, whether it contains polar bonds, and how it behaves near a charged object? If we know whether a compound's molecules are polar, can we predict its properties? Indeed we can. As you learned in Section 2.3, the covalent bonds in a molecule may be polar or non-polar. Polar bonds may cause the whole molecule to be polar. A molecule is classified as a **polar molecule** if it has a positively charged end and a negatively charged end (Figure 1). A molecule that has no charged ends—that has its charges equally distributed—is classified as a **non-polar molecule**. We will eventually be able to predict the relative melting point and boiling point of a molecular compound by considering whether it is polar or not. In addition, we will be able to explain why a substance such as water has such special properties and why it is such an important substance.

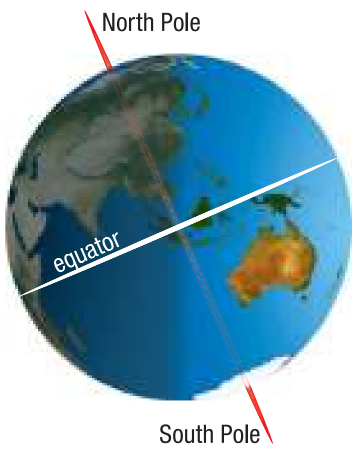
polar molecule a molecule in which the uneven distribution of electrons results in a positive charge at one end and a negative charge at the other end

non-polar molecule a molecule in which the electrons are equally distributed among the atoms, resulting in no localized charges

LEARNING **TIP**

Pole-r

Do you find the term "polar" hard to understand? It might be useful to think about other examples of polar things. For example, Earth has a North Pole and a South Pole. Magnets have north and south poles. Bipolar disorder is a psychiatric condition where one feels excessive moods of elation and then depression. In each of these cases, there are two extremes: two poles.



Explaining the Behaviour of Polar Liquids

As you know, non-polar molecules do not have charged ends. Refer back to Table 2 to see some examples of polar and non-polar compounds.

What happened in the Mini Investigation when a positively charged acetate strip was brought close to a thin stream of a polar liquid, such as water? The stream of liquid bent toward the strip. The polar water molecules arranged themselves so that their negative poles were closer to the positively charged object. When a negatively charged vinyl strip came close to the stream of water, the water was again deflected toward the strip. In this case, the water molecules rearranged themselves so that their positive poles were closer to the negatively charged material (**Figure 5**). Polar molecules are attracted by either kind of charge.

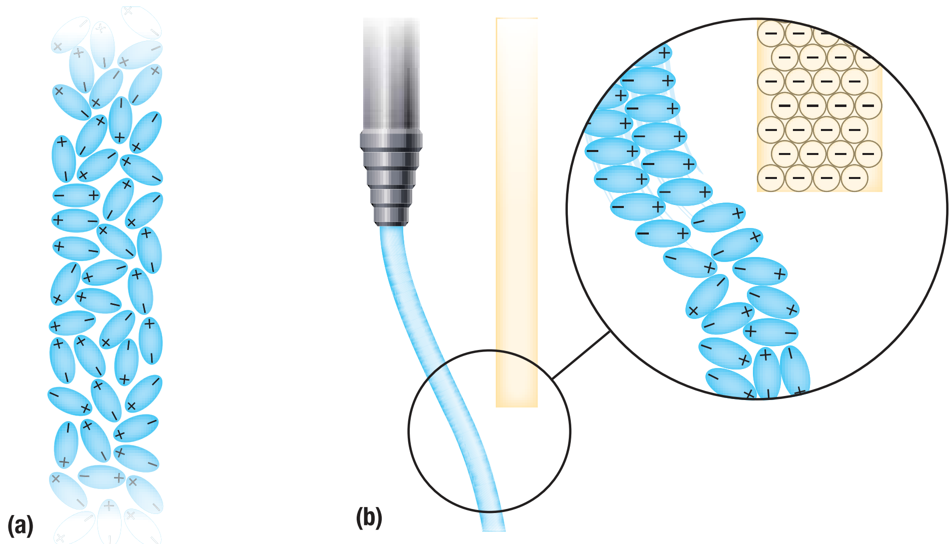


Figure 5 (a) Polar water molecules are randomly oriented in a thin stream of water. (b) Polar molecules in a liquid become oriented so that their positive poles are closer to a negatively charged material.

You have seen some empirical evidence for the existence of polar molecules. If we can determine whether a molecule is polar or non-polar, we will be able to predict and explain why a compound has certain physical properties. We can determine whether or not a molecule is polar by examining two of its characteristics: the polarity of the covalent bonds and the molecule's three-dimensional shape.

Polar Molecules

Consider two molecules: hydrogen chloride, HCl, and water, H_2O. You know that the bond in the hydrogen chloride molecule is a polar covalent bond, so the molecule has a positively charged end and a negatively charged end. The whole molecule is therefore polar (**Figure 6(a)**).

What about water, H_2O? Are water molecules polar? Remember that a water molecule is made up of an oxygen atom and 2 hydrogen atoms. There are two O–H bonds. The difference between the electronegativities of oxygen and hydrogen atoms is 1.4, so each bond is a polar covalent bond. The oxygen atom attracts the bonding electrons most strongly. There is therefore a negative charge around the oxygen atom and a positive charge around each hydrogen atom (**Figure 6(b)**). Since the water molecule is lopsided, or bent, we can consider that it has two ends: a negatively charged oxygen end and a positively charged hydrogen end. Thus, the water molecule is polar.

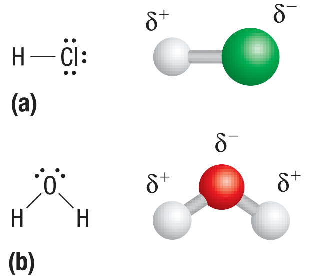


Figure 6 The Lewis structures and molecular shapes of (a) hydrogen chloride, HCl, and (b) water, H_2O. Both are polar molecules.

Non-Polar Molecules

Oxygen, O_2, and carbon dioxide, CO_2, are two examples of non-polar molecules (**Figure 7**). The oxygen molecule, O_2, is diatomic: it contains only 2 atoms. Since both of the covalent bonds in an oxygen molecule are non-polar, the whole molecule is non-polar. To be polar, a molecule must have at least one polar covalent bond. Diatomic molecules made up of identical atoms are always non-polar.

The carbon dioxide molecule, however, is polyatomic: it contains more than two atoms. To be specific, a molecule of carbon dioxide is composed of 3 atoms joined by four covalent bonds: two double bonds, each between an oxygen atom and a carbon atom. The electronegativity difference between oxygen and carbon is 1.0, with oxygen having the higher electronegativity. Consequently, each C=O bond is polar. But look closely at the shape of the CO_2 molecule in Figure 7(b). Notice that the two polar covalent C=O bonds are arranged symmetrically about the central carbon atom. In a symmetrical, linear molecule like carbon dioxide, the positive and negative charges cancel each other out. The molecule is non-polar overall. Note that a molecule can include polar covalent bonds but still be non-polar overall if it is symmetrical.

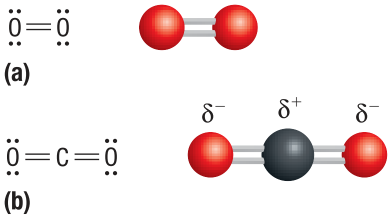


Figure 7 The Lewis structure and molecular shape of (a) oxygen, O_2, and (b) carbon dioxide, CO_2. Both are non-polar molecules.

Research This

Waste Wisdom

Skills: Researching, Analyzing, Communicating, Defining the Issue, Defending a Decision

SKILLS HANDBOOK A5.1

You have just painted your room with a fresh coat of oil-based paint. You carefully wash the paintbrushes in turpentine. Now what are you going to do with the leftover paint and used turpentine? Should you pour them down the sink? Seal them into a container and put them in the garbage? Take them to a hazardous-waste centre?

We are often faced with tough waste-disposal issues. As consumers, how are we to know what to do with hazardous household waste? How can we maintain safe and clean environmental conditions? Are there any guidelines or regulations regarding hazardous-waste disposal? Who develops and enforces them?

1. Choose a consumer product, such as paint, turpentine, motor oil, or batteries, that should not be placed in the regular garbage.
2. Research the correct way to dispose of this product, and the dangers to the environment if it is improperly disposed of.
3. Research any guidelines or regulations regarding disposal of this product in your municipality.
4. Research to find out whether most people comply with the guidelines or regulations.

A. Summarize your research. Do most people follow the rules for proper disposal? Why or why not, in your opinion? T/I A

B. What else could be done to protect the environment from hazardous household waste? Justify your position. A C

GO TO NELSON SCIENCE

How to Determine if a Molecule Is Polar or Non-polar

It is very useful to know whether a molecule is polar or non-polar. This knowledge helps us to predict a compound's physical properties, such as its solubility in other substances. We can predict whether a molecule is polar by considering whether it contains polar covalent bonds and the orientation of these bonds around the central atom in the molecule.

Tutorial 1 Determining the Polarity of Molecules

There is a series of steps you can follow to determine whether a molecule is polar or non-polar. The flow chart in **Figure 8** presents the steps visually.

1. Determine how many atoms of which elements make up a molecule.
2. Draw the Lewis structure for the molecule.
3. Determine how many covalent bonds there are in the molecule.
4. Determine the electronegativity difference for each covalent bond in the molecule. Indicate whether each bond is polar ($0.0 > \Delta EN > 1.7$) or non-polar ($\Delta EN = 0.0$).
5. If there are polar covalent bonds in the molecule, indicate the partial charges. Write "δ^+" by the atom with the lower electronegativity and "δ^-" by the atom with the higher electronegativity.
6. Interpret your diagram. If the molecule only has one polar covalent bond, the molecule is polar. If the molecule has more than one polar covalent bond, the molecule may or may not be polar. Examine the shape of the molecule. Is it symmetrical or asymmetrical? **Table 3** shows which types of compounds tend to be polar and non-polar molecules. Note that there are some exceptions to these general rules.

WEB LINK

There are many powerful computer programs that help chemists draw structures for molecular compounds. To explore some of these programs,

GO TO NELSON SCIENCE

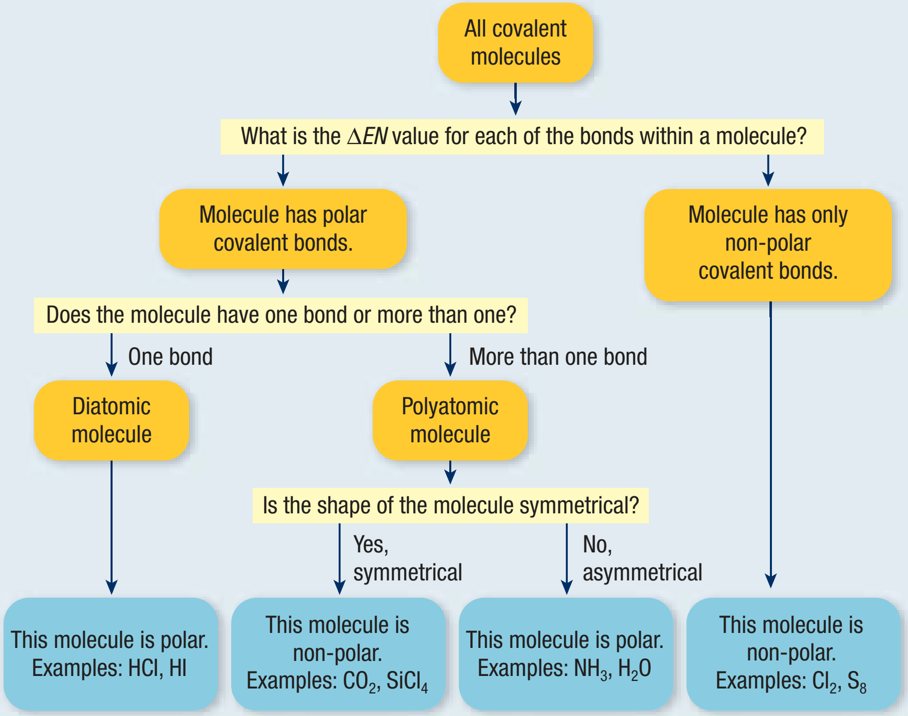


Figure 8 This flow chart can help you to determine whether a molecule is polar or non-polar.

Table 3 Some Rules for Determining Polarity of Polyatomic Molecules

General chemical formula	Polarity	Examples
diatomic; 2 different atoms	polar	HCl, CO
nitrogen and 3 other atoms of the same element	polar	NH_3, NF_3
oxygen and 2 other atoms of the same element	polar	H_2O, OCl_2
carbon and other atoms of two or more elements	polar	$CHCl_3$, C_2H_5OH
diatomic; 2 identical atoms	non-polar	N_2, O_2
carbon and 2 or more atoms of the same element	non-polar	CH_4, CO_2

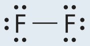
Sample Problem 1: Determining the Polarity of a Diatomic Molecule

Determine whether a molecule of fluorine is polar or non-polar.

Step 1. Determine how many atoms of which elements make up a molecule.

There are two atoms of fluorine in a molecule.

Step 2. Draw the Lewis structure for the molecule.



Step 3. Determine how many covalent bonds there are in the molecule.

There is one covalent bond.

Step 4. Determine the electronegativity difference for each covalent bond in the molecule. Indicate whether the bond is polar or non-polar.

There is no electronegativity difference between the 2 identical atoms. The bond is non-polar. This is therefore a non-polar molecule.

Sample Problem 1 follows our general rule that diatomic molecules made up of 2 of the same atom are always non-polar. (See Table 3.)

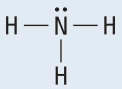
Sample Problem 2: Determining the Polarity of a Polyatomic Molecule

Determine whether a molecule of ammonia is polar or non-polar.

Step 1. Determine how many atoms of which elements make up a molecule.

Ammonia, NH_3, has one nitrogen atom and three hydrogen atoms in each molecule.

Step 2. Draw the Lewis structure for the molecule.



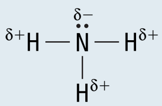
Step 3. Determine how many covalent bonds there are in the molecule.

There are three covalent N–H bonds.

Step 4. Determine the electronegativity difference for each covalent bond in the molecule. Indicate whether the bond is polar or non-polar.

$\Delta EN = 0.9$. Each N–H bond is therefore polar covalent.

Step 5. If there are polar covalent bonds in the molecule, indicate the partial charges.



Step 6. Interpret your diagram. If the molecule only has one polar covalent bond, the molecule is polar. If the molecule has more than one polar covalent bond, the molecule may or may not be polar. Examine the chemical formula and, if possible, the shape of the molecule. Is it symmetrical or asymmetrical? (See Table 3.)

The molecule has three N–H polar covalent bonds. It is asymmetrical with a lone pair on the top. Ammonia is therefore a polar molecule.

Practice

1. Determine whether each of the following molecules is polar or non-polar: T/I
   (a) hydrogen bromide, HBr
   (b) nitrogen gas, N_2
   (c) hydrogen sulfide, H_2S
   (d) ethane, C_2H_6
   (e) tetrachloroethene (perchloroethylene or "perc") C_2Cl_4
   (f) phosphine, PH_3

3.3 Summary

- Whether a bond is polar covalent or non-polar covalent depends on the electronegativity difference between the two bonded atoms.
- Molecular polarity depends on the polarity of the bonds within the molecule and on the shape of the molecule.
- A diatomic molecule will be polar if the covalent bond is polar and non-polar if the bond is non-polar.
- A polyatomic molecule with only non-polar covalent bonds will be non-polar.
- A polyatomic molecule with several polar covalent bonds will be polar if the molecule is asymmetrical and non-polar if the molecule is symmetrical.
- Molecular diagrams and models help us to visualize whether a molecule is symmetrical or asymmetrical.

3.3 Questions

1. Arrange the following sets of bonds in order of decreasing polarity (from most polar to least polar). Refer to the electronegativity values in the periodic table at the back of this textbook. K/U
   (a) K–Br, H–Br, C–H, O–F
   (b) C–O, C–F, O–H, H–H
2. For the most polar bonds in Question 1, draw the bond and indicate which atom would have the partial positive charge, δ^+, and which atom would have the partial negative charge, δ^-. K/U C
3. Draw a diagram showing the orientation of water molecules when a thin stream of water is deflected by a positively charged acetate rod. T/I C
4. Determine whether each of the following molecules is symmetrical or asymmetrical. Explain your reasoning. K/U T/I
   (a) O_2
   (b) H_2O
   (c) CO_2
   (d) CH_4
   (e) NH_3
5. Determine whether each of the following molecules is polar or non-polar: K/U T/I
   (a) CCl_4
   (b) H_2O
   (c) HF
   (d) CF_4
   (e) CH_3Cl
6. **Table 4** shows the results of an investigation similar to the Mini Investigation: Evidence for Polar Molecules. Predict whether liquids (a) to (d) are polar or non-polar. Give your reasons for each prediction. T/I

Table 4 Observations of the Effect of Electric Charge

Liquid	Chemical formula	With acetate strip	With vinyl strip
(a) carbon tetrachloride	$CCl_4(l)$	unaffected	unaffected
(b) methanol	$CH_3OH(l)$	attracted	attracted
(c) hexane	$C_6H_{14}(l)$	unaffected	unaffected
(d) nitrogen trichloride	$NCl_3(l)$	attracted	attracted

7. Research companies that are using carbon dioxide for dry cleaning. Summarize the pros and cons of replacing the conventional cleaning fluid with carbon dioxide. T/I C
8. Two types of computer-generated images (CGI) are commonly used to represent the three-dimensional shapes of molecules: the ball and stick model and the space-filling model. Research these two models, find examples, and list the pros and cons of each model. T/I C

GO TO NELSON SCIENCE

Intermolecular Forces

3.4

You have seen water boil. Bubbles form in the liquid and vapour rises from its surface (**Figure 1**). But what is happening at a molecular level? Similarly, what is happening to carbon dioxide as it sublimates to produce fog in a fog machine (**Figure 2**)? In both cases, individual molecules are gaining kinetic energy. This kinetic energy comes from thermal energy provided by the hot stovetop or the heating element in the fog machine.

Both boiling (evaporating) and sublimating are physical changes: they do not result in any new chemical substances. Water molecules in a pot of water (the liquid state) are the same as water molecules in water vapour (the gaseous state). Similarly, carbon dioxide molecules in a piece of dry ice (the solid state) are the same as carbon dioxide molecules in gaseous carbon dioxide. So what is the difference between a solid or liquid substance, and one in the gaseous state? As a result of having enough energy, the water or carbon dioxide molecules are able to move apart from one another and pass into the gaseous state. So what was holding these molecules together in the first place?

So far, in this chapter, you have learned a lot about covalent bonds: strong forces of attraction that hold the atoms of a molecule together. But covalent bonds only act within molecules, not between them. What is holding one molecule close to an adjacent one? Molecules are attracted to other molecules by much weaker forces, known as intermolecular forces. An **intermolecular force** is an attraction between molecules. When a substance evaporates, melts, or sublimates, the molecules gain kinetic energy from an outside source. This energy allows them to overcome the intermolecular forces holding them close together.

Figure 1 As water reaches its boiling point, some of the water molecules leave the liquid state and join the gaseous state.

intermolecular force the attractive force between molecules

Figure 2 A fog machine makes use of the physical properties of carbon dioxide. Carbon dioxide sublimates at low temperatures because of its weak intermolecular forces.

Comparing Bonds and Intermolecular Forces

There are no intermolecular forces in ionic compounds because ionic compounds do not contain molecules. Ionic compounds are always solids at room temperature. In an ionic crystal, ionic bonds hold all the ions together. There is no difference between the bonds that hold together one formula unit of the compound and those that hold all of the formula units together in a crystal (**Figure 3**).

Ionic bonds are, in general, very strong. This accounts for the high melting point of ionic compounds. Sodium chloride, for example, must be heated to 801 °C before it will become a liquid.

Many molecular substances (carbon dioxide, oxygen, and nitrogen, for example) are gases at room temperature. Others are liquids with fairly low boiling points or solids that melt quite easily (hydrogen chloride and table sugar, respectively). Researchers conclude that the forces between the molecules of molecular solids and liquids are relatively weak (**Figure 4**). The addition of thermal energy can easily overcome these intermolecular forces. It should be noted that chemical bonds, both ionic and covalent, are considerably stronger than intermolecular forces.

LEARNING TIP

International Interactions
The prefix *inter-* is used in many contexts. Some familiar uses include "international," "intermission," and "interstellar." In each case, "inter" means "between." International treaties are agreements that exist between individual nations. Similarly, intermolecular forces are forces that exist between individual molecules.

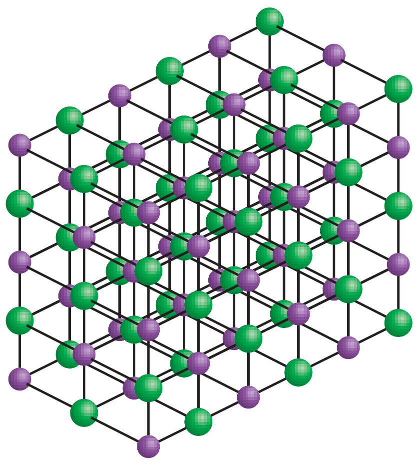


Figure 3 Ionic compounds do not have intermolecular forces. They are held together by ionic bonds, which are much stronger than intermolecular forces.

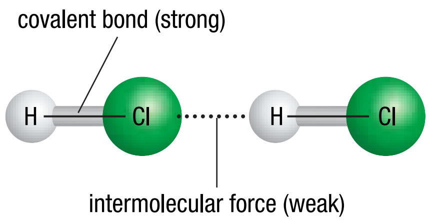


Figure 4 Evidence indicates that, compared with covalent bonds, intermolecular forces are quite weak.

Intermolecular Forces and Physical Properties

Every compound has unique, or characteristic, physical properties. Chemists explain this by suggesting that every compound is made of a unique arrangement of atoms. The strength of the intermolecular forces in every compound is therefore unique.

The strength of intermolecular forces determines the following physical properties of molecular compounds:

- physical state of a compound at a specific temperature and pressure
- melting point
- boiling point
- surface tension
- hardness and texture
- solubility in various solvents

We can make the following generalization about molecular compounds: as the intermolecular forces between molecules increase, the compound's melting point, boiling point, and surface tension also increase.

How can we explain these physical properties? We can hypothesize that a molecular compound that is a gas at room temperature has weak intermolecular forces while one that is a solid has relatively strong intermolecular forces. Even the strongest intermolecular forces, however, are not as strong as covalent bonds.

Figure 5 Johannes van der Waals is remembered for proposing the existence of three different types of intermolecular forces.

Types of Intermolecular Forces

Johannes van der Waals, a Dutch physicist, first developed the theory of intermolecular forces in the late nineteenth century (**Figure 5**). He experimented by cooling a variety of gases into liquids. His work led him to suggest that there are three main types of intermolecular forces. These three forces are dipole–dipole forces, London dispersion forces, and hydrogen bonds.

Dipole–Dipole Forces

As you now know, some molecules are polar and others are non-polar. We can look for patterns in the properties of polar and non-polar compounds. Hydrogen fluoride, for example, is polar. It is a colourless gas at room temperature. **Table 1** compares hydrogen fluoride's melting point and boiling point to those of pure hydrogen and pure fluorine. The evidence indicates that the forces between hydrogen fluoride molecules are considerably stronger than those between hydrogen molecules or between fluorine molecules.

Table 1 Melting Points and Boiling Points of Three Molecular Substances

Substance	Melting point (°C)	Boiling point (°C)
hydrogen fluoride, HF	−83.6	19.5
hydrogen, H_2	−259.1	−252.9
fluorine, F_2	−219.6	−188.1

As a polar molecule, hydrogen fluoride has a negatively charged fluoride end and a positively charged hydrogen end. When two hydrogen fluoride molecules are next to one another, the positive end of one molecule is attracted to the negative end of the next molecule (**Figure 6**). The intermolecular force between oppositely charged ends of polar molecules is called a **dipole–dipole force**. In general, dipole–dipole forces are relatively strong intermolecular forces. They occur between all polar molecules. We can hypothesize that the more polar a substance is, the stronger the dipole–dipole force of attraction that exists.

dipole–dipole force an intermolecular force of attraction that forms between the slightly positive end of one polar molecule and the slightly negative end of an adjacent polar molecule

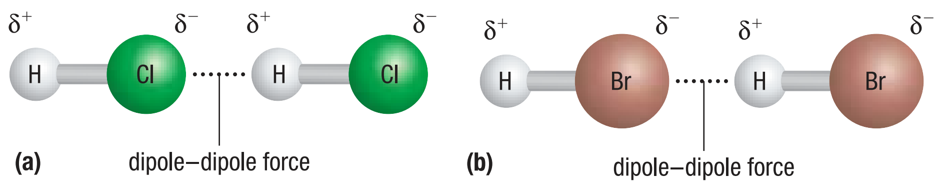


Figure 6 Dipole–dipole forces form between the slightly positive end of one molecule and the slightly negative end of an adjacent molecule. Which is the more polar substance? Which would have a great dipole–dipole attraction?

London Dispersion Forces

Consider some non-polar substances such as hydrogen gas or fluorine gas. Their boiling points are very low (–253 °C and –188 °C, respectively), indicating that the intermolecular forces are very weak. All gases, even the noble gases, condense to form liquids if they are cooled enough. The fact that these substances can form liquids implies that there is some type of attraction between entities when they are cooled. These attractions are caused by temporary shifts in the electron "cloud" in an atom or molecule. This shift produces a temporary dipole for a fraction of a second. The temporary dipole causes a temporary dipole to form in a neighbouring atom or molecule. The attraction of these dipoles is the **London dispersion force**, sometimes shortened to "London force." Since these dipoles are so short-lived, the attraction between them continually forms and breaks. As a result, London forces are very weak. London dispersion forces are named after Fritz London, a German-American physicist.

Table 2 shows some physical properties of several non-polar substances. Notice that helium, hydrogen, argon, and chlorine exist as gases at room temperature and pressure. Bromine, however, has a boiling point of 58.8 °C. It is therefore a liquid at room temperature. Sulfur often exists naturally in a ring-like arrangement of 8 sulfur atoms. In this form, sulfur is a yellow solid at room temperature. There is a relationship between the size of a molecule and its state. We can explain this with London dispersion forces: the larger the molecule, the more electrons and protons there are attracting each other, the stronger the forces, and thus the higher the melting point. Synthetic compounds with large molecules can have enough London dispersion forces to give them very high melting points. These compounds can be as strong as steel.

Investigation 3.4.1

Melting Points of Molecular Compounds (p. 123)
This investigation guides you to collect your own evidence for the relationship between melting temperature and molecular size. You will compare data for several similar compounds of different sizes.

London dispersion force a weak attractive force acting between all entities, including non-polar molecules and unbonded atoms, caused by the temporary imbalance of electrons within entities

Table 2 Comparison of Physical Properties for Several Non-polar Substances

Element	Chemical formula	Boiling point (°C)	Melting point (°C)	Number of electrons per molecule	
helium	He	−268.9	−272.2	2	increasing strength of London forces ↓
hydrogen	H_2	−252.9	−259.1	2	
argon	Ar	−185.8	−189.3	18	
chlorine	Cl_2	−34	−101	34	
bromine	Br_2	58.8	−7.2	70	
sulfur	S_8	444.6	115.2	128	

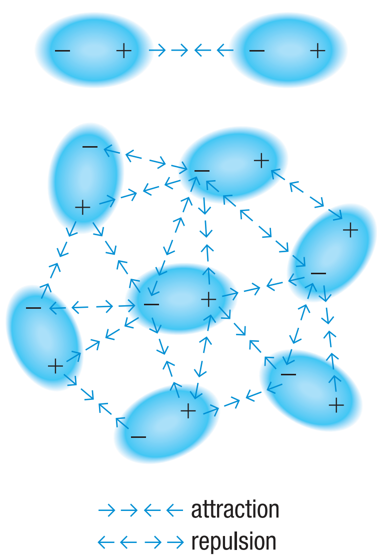


Figure 7 London dispersion forces are caused by a temporary imbalance in the position of electrons. These intermolecular forces can exist between single atoms or molecules.

London forces exist between all molecules, even polar molecules. In a polar compound, however, the London forces are insignificant compared to the much stronger dipole–dipole forces.

Chemists now class dipole–dipole forces and London dispersion forces as **van der Waals forces** in honour of the scientist who first proposed their existence.

van der Waals forces the weak forces of attraction between molecules, including dipole–dipole forces and London dispersion forces

Hydrogen Bonds

Why does water bead into droplets on a smooth surface? Why do raindrops form? How do some spiders walk on the surface of water? All of these questions have the same answer: because of a special bond between water molecules.

We have already seen how dipole–dipole intermolecular forces are created (Figure 6). The boiling point of hydrogen chloride, HCl, is −85 °C. Why would a similar molecule, hydrogen fluoride, HF, have a considerably higher boiling point of 19.5 °C? Chemists theorize there are very strong dipole–dipole intermolecular forces that fall into a separate category known as hydrogen bonds.

hydrogen bond an unusually strong dipole–dipole force between a hydrogen atom attached to a highly electronegative atom (N, O, or F) and a highly electronegative atom in another molecule

A **hydrogen bond** is a particularly strong dipole–dipole force that occurs between two molecules (**Figure 8**). Each molecule consists of a hydrogen atom covalently bonded to a highly electronegative atom of nitrogen, oxygen, or fluorine. (These are the three most electronegative elements of all, partly because they are so small.) Strong attractions form between a positive hydrogen atom of one molecule and the highly electronegative N, O, or F atom of another molecule (**Figure 9**).

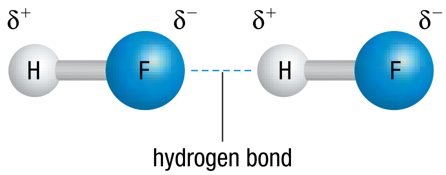


Figure 8 A hydrogen bond is a strong attractive force between a highly electronegative atom in one molecule and a hydrogen atom in a nearby molecule.

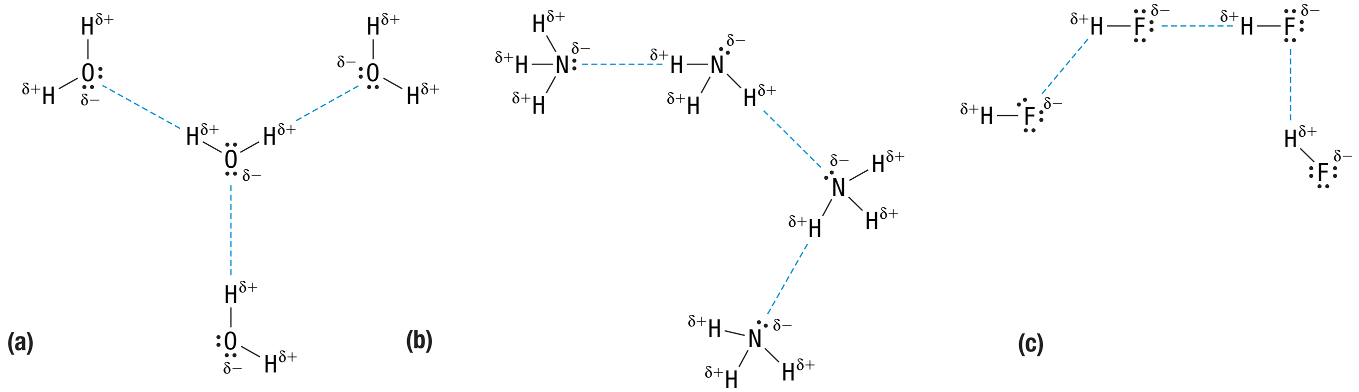


Figure 9 Hydrogen bonds form between molecules that contain hydrogen and (a) oxygen, (b) nitrogen, or (c) fluorine.

There are two main reasons for hydrogen bonding. First is the large difference in electronegativities. Second is the small size of the hydrogen atom, which means that the positive pole is highly concentrated. It strongly attracts the negative pole of a nearby molecule.

HYDROGEN BONDS IN BIOCHEMISTRY

Hydrogen bonds play a significant role in determining the shape and function of large, biologically important molecules. Two examples of such molecules are proteins and deoxyribonucleic acid (DNA). Proteins are essential to living things: they function as enzymes, hormones, antibodies, and structural materials. Proteins are very long molecules, consisting of hundreds or even thousands of atoms. The chain of atoms folds into very specific three-dimensional structures because of attractions between different parts of the chain.

DNA is a very specialized biological molecule that stores the genetic information in a cell. Each DNA molecule is made up of two long chains of structures called nucleotides. This is all assembled in a double helix structure, like a twisted rope ladder. The nucleotides include many nitrogen and oxygen atoms, as well as hydrogen. Hydrogen bonds readily form between the two helixes, holding them together like the rungs of the ladder (**Figure 10**).

The DNA structure has to come apart in order to replicate itself, so the hydrogen bonds between the two chains are broken and re-formed. Very simply put, it is as if the DNA molecule is unzipped and zipped back up again. This is a key part of how cells reproduce, and how genetic information is passed from one generation to another.

WEB LINK

To view three-dimensional structures of DNA and other complex molecular compounds,

GO TO NELSON SCIENCE

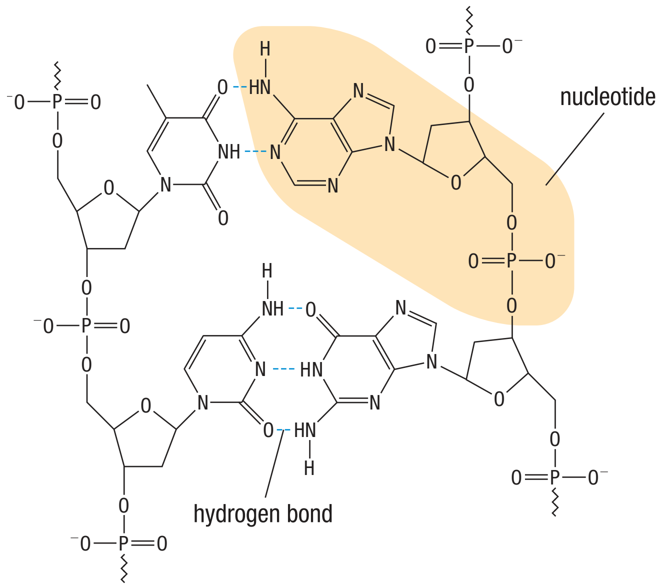


Figure 10 Hydrogen bonds form between hydrogen and oxygen or nitrogen atoms in a DNA molecule. There are many carbon atoms in DNA. In this diagram, the carbon atoms are not shown. Rather, they are implied by a change of direction in the bonds.

Mini Investigation

Relating Boiling Points to Intermolecular Forces

Skills: Hypothesizing, Predicting, Analyzing, Evaluating, Communicating

SKILLS HANDBOOK A2.2, A6.4

In all liquids, intermolecular forces hold the molecules together. These forces, however, become negligible above the boiling point when the substance becomes a gas. During boiling, therefore, intermolecular forces must be overcome by the addition of energy. The temperature at which a liquid boils reflects the strength of its intermolecular forces. A higher boiling point means that more energy has been added to separate the molecules, and thus the intermolecular forces must be strong.

In this investigation you will use provided data to explore the trend in boiling points of the hydrogen compounds of elements in Groups 14 to 17.

A. Write a hypothesis for the trend in boiling points within and between groups. Refer to molecular polarity, dipole–dipole forces, London dispersion forces, and hydrogen bonds in your explanation. T/I

B. Copy **Table 3** into your notebook, adding a column with the title "Number of electrons per molecule." Determine and record the number of electrons for each compound. K/U C

C. Create a graph of boiling point versus number of electrons per molecule. T/I C

D. Use your graph to answer the question, "What is the trend in boiling points of the hydrogen compounds of elements in Groups 14 to 17?" T/I

E. Assuming that the evidence is valid, evaluate your hypothesis and the concept of intermolecular forces used in your explanation. T/I

F. Are there any anomalies (unexpected evidence) in the evidence presented? Suggest an explanation. T/I

Table 3 Boiling Points of the Hydrogen Compounds of Elements in Groups 14 to 17

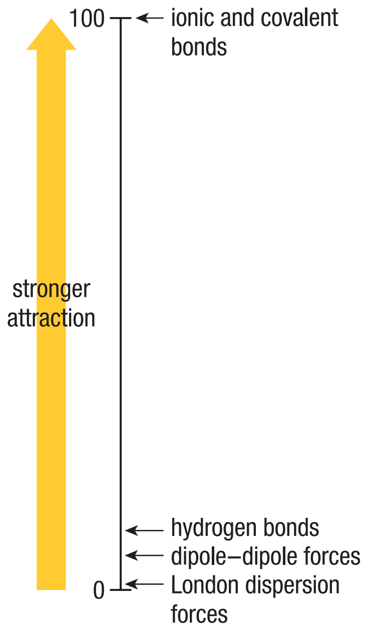
Group	Hydrogen compound	Boiling point (°C)
14	$CH_4(g)$	−162
	$SiH_4(g)$	−112
	$GeH_4(g)$	−89
	$SnH_4(g)$	−52
15	$NH_3(g)$	−33
	$PH_3(g)$	−87
	$AsH_3(g)$	−55
	$SbH_3(g)$	−17
16	$H_2O(l)$	100
	$H_2S(g)$	−61
	$H_2Se(g)$	−42
	$H_2Te(g)$	−2
17	$HF(g)$	20
	$HCl(g)$	−85
	$HBr(g)$	−67
	$HI(g)$	−36

LEARNING TIP

The Strength of Bonds and Forces

Remember that intermolecular forces are very different from covalent bonds. Intermolecular forces act *between* molecules, attracting them toward each other. Covalent bonds hold the atoms *within* a molecule together. Even the strongest intermolecular force—the hydrogen bond—is much weaker than a covalent bond.

If we assign ionic and covalent bonds a strength of 100 on a relative scale, the intermolecular forces would all be below 10 on the scale.



3.4 Summary

- Intermolecular forces are the forces of attraction between molecules.
- There are three types of intermolecular forces: dipole–dipole forces, London dispersion forces, and hydrogen bonds. Dipole–dipole forces and London dispersion forces are known, collectively, as van der Waals forces.
- London dispersion forces exist between all molecules and are relatively weak.
- Dipole–dipole forces exist only between polar molecules. They are intermediate in strength.
- Hydrogen bonds are dipole–dipole forces between an electronegative atom (usually nitrogen, oxygen, or fluorine on one molecule) and a hydrogen atom that is bonded to another electronegative atom on another molecule.
- As intermolecular forces increase, melting point, boiling point, and surface tension increase.
- The polarity of a molecule can be determined by considering its shape and the electronegativity difference of its atoms (**Figure 11**).

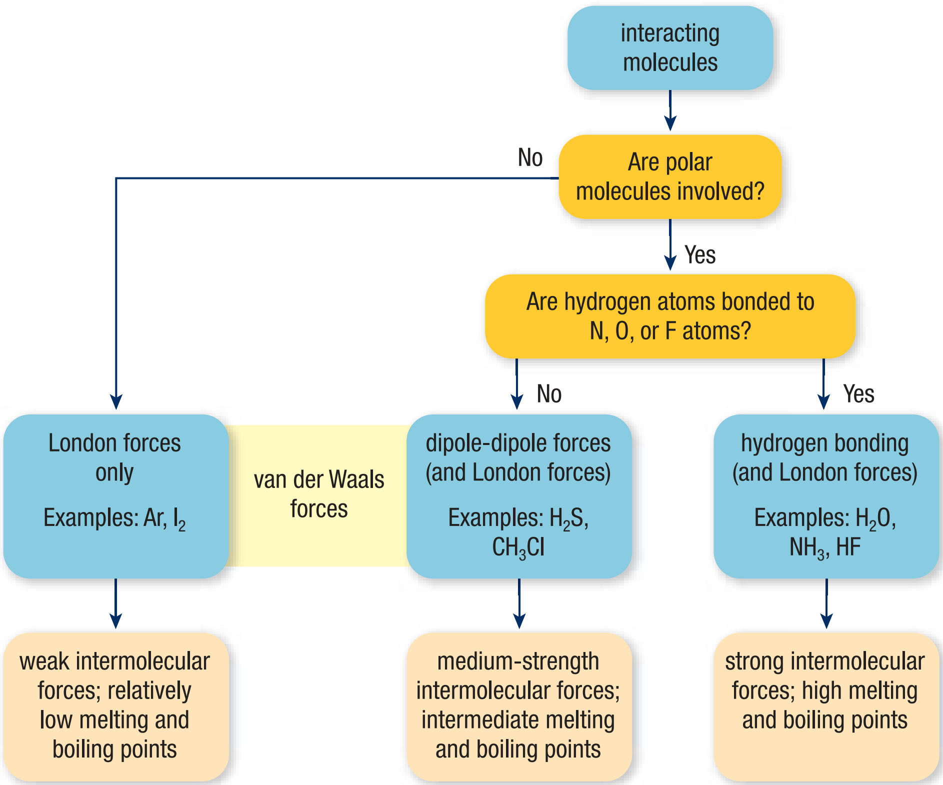


Figure 11 A summary of forces between molecules

3.4 Questions

1. Identify the specific type of intermolecular force illustrated in each diagram in **Figure 12**. K/U

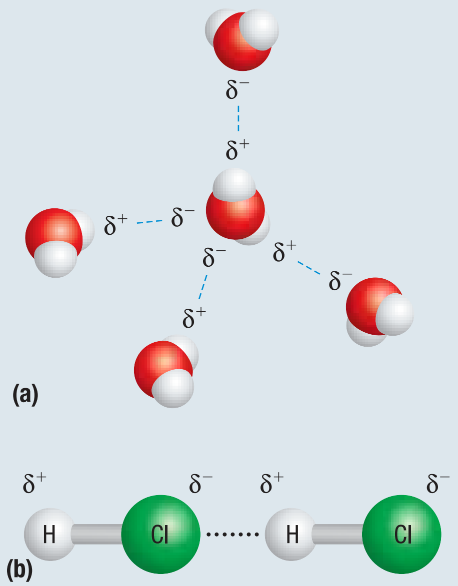


Figure 12 (a) Water molecules and (b) hydrogen chloride molecules

2. List the three types of intermolecular forces in order of increasing strength. K/U
3. Arrange the following types of bonds and forces in a continuum to indicate their relative strengths: covalent bonds, hydrogen bonds, London dispersion forces, dipole–dipole forces. T/I C
4. List all the intermolecular forces that exist between molecules of each of the following substances: T/I
   (a) oxygen gas, $O_2(g)$
   (b) dichlorine monoxide, $Cl_2O(g)$
   (c) water, $H_2O(l)$
   (d) methane, $CH_4(g)$
   (e) chloromethane, $CH_3Cl(g)$
5. Using your knowledge of molecular polarity and intermolecular forces, arrange the following substances in order from lowest melting point to highest: T/I

   H_2O, CH_4, CH_3Cl, NH_3, CH_3OH, H_2, C_3H_8
6. Write a hypothesis predicting whether molecular compounds or ionic compounds tend to have stronger odours. Include a theoretical explanation. K/U T/I
7. Explain each of the following observations about a pair of molecular substances. Refer to intermolecular forces in your explanation. T/I
   (a) Water has a considerably higher boiling point than methane, $CH_4(g)$.
   (b) Bromine is a liquid at ordinary temperatures whereas chlorine is a gas.
   (c) Trichloromethane, $CHCl_3$, has a higher boiling point than tetrachloromethane, CCl_4.
8. Superfluidity is a strange phenomenon that occurs when non-polar substances are subjected to extremely high pressures and low temperatures. This creates a different state known as a Bose–Einstein condensate. Liquid helium in this state will slowly creep up the walls of a container and drip out until the container is empty (**Figure 13**). Research this different state of matter and suggest what might be happening to intermolecular forces at extremely low temperatures. A

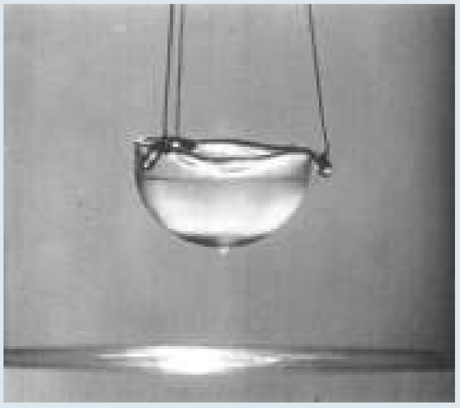


Figure 13 A Bose–Einstein condensate

9. A viscous liquid is one that is partway between a solid and a liquid. It flows, but only slowly. T/I A
   (a) List at least five viscous substances.
   (b) Indicate why the viscosity of each substance is a useful property.
   (c) Identify the intermolecular forces in each substance. You may have to search some references to find this information.
   (d) Referring to intermolecular forces, explain why viscous substances flow so slowly.

GO TO NELSON SCIENCE

3.5 Hydrogen Bonding and Water

Water is a very special molecular compound. It is one of the most abundant compounds on Earth's surface and is essential for life. A large percentage (~22 %) of Earth's fresh water is within Canada's national boundaries. Many countries have limited access to fresh water, which leads to major health problems and food shortages. Our bodies need water to transport nutrients, eliminate waste, and help maintain electrolytic balance in our cells. The water we use, however, must be free of chemical or biological contamination or our health and that of our planet is jeopardized (**Figure 1**).

In nature, water contains dissolved substances such as ions (Ca^{2+}, Mg^{2+}, HCO_3^-, and SO_4^{2-}) and gases (CO_2, O_2). These substances play various roles in sustaining life. The water cycle conveniently recycles and purifies water for us. However, some pollutants accumulate in our aquifers, surface waters, and oceans.

You might think that, because water is so important, we would respect and protect our water sources. Recent analyses, however, show increasing contamination of our drinking water. Contaminants include caffeine, cosmetics, hormones, food additives, and antibiotics. In Chapter 8 you will find out more about substances in water.

We will now explore why water is such an unusual substance. We will apply our knowledge of molecular polarity and hydrogen bonds to explain its properties.

WEB LINK

To find out about other unique properties of water, and the effects that these properties have on the world around us,

GO TO NELSON SCIENCE

Special Properties of Water

Water, as you know, is a very common substance. However, compared with other, similar compounds, it has some very strange physical properties.

Unusually High Melting and Boiling Points

Water is the only substance that commonly exists as a solid, a liquid, and a gas on Earth's surface. Water melts at a relatively high temperature of 0 °C and it boils at 100 °C. These are much higher than the corresponding temperatures for similar molecular compounds, such as ammonia, NH_3, and hydrogen sulfide, H_2S (**Table 1**). This observation implies that the forces holding water molecules to one another are unusually strong.

Table 1 Melting Points and Boiling Points of Three Molecular Compounds

Compound	Melting point (°C)	Boiling point (°C)
water, H_2O	0	100
ammonia, NH_3	−78	−33
hydrogen sulfide, H_2S	−86	−60

The polar covalent O–H bonds and bent shape, along with the two lone pairs of electrons on the central oxygen atom, combine to make water one of the most polar molecules on Earth. As you learned in Section 3.4, this polarity results in hydrogen bonds forming between the water molecules (**Figure 1**). These bonds are responsible for water's relatively high melting and boiling points.

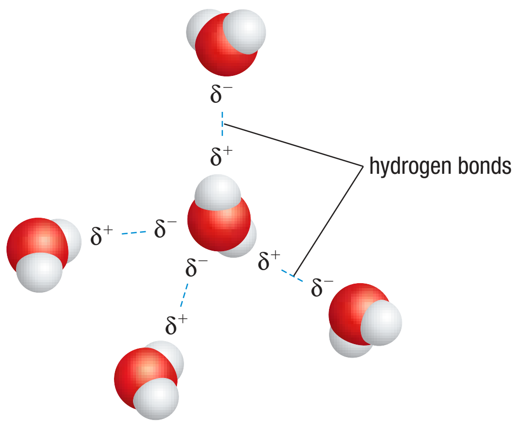


Figure 1 The negative (oxygen) part of one water molecule attracts a positive (hydrogen) part of another.

The Low Density of Ice

In most compounds, such as wax, the solid state is more dense than the liquid state. The solid therefore sinks in the liquid. Solid water (ice), however, floats in liquid water. Remember that, when water is frozen, the molecules cannot move around. Thus, they cannot easily conduct thermal energy from the slightly warmer water below to the air above. Also, water molecules in ice are less likely to evaporate, which would also transfer thermal energy to the air. Ice therefore acts as a blanket, insulating the liquid water from the cold air above. What impact do you think this has on aquatic ecosystems? Imagine what would happen if lakes froze from the bottom up.

Ice is 9 % less dense than liquid water. As liquid water freezes, the molecules lose kinetic energy and slow down. Hydrogen bonds start to arrange the molecules into a crystalline structure (**Figure 2**). This arrangement contains more space between the molecules than does the more random arrangement of liquid water molecules. This expansion results in many effects, from burst water pipes to the freeze–thaw weathering of geological structures (**Figure 3**). In a more complex process, snowflakes form beautiful geometric patterns as airborne water droplets crystallize (**Figure 4**).

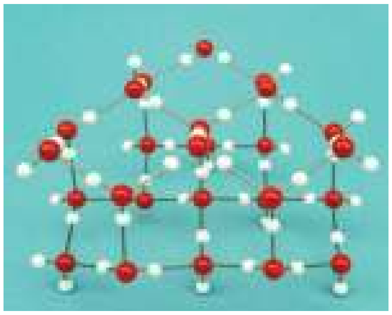


Figure 2 As ice forms, hydrogen bonds (shown as grey connectors) assemble the molecules with more space between them than when the compound is liquid.

Figure 3 The expansion of trapped water as it freezes can crack apart rocks.

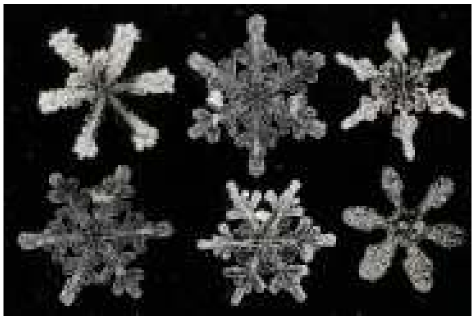


Figure 4 Hydrogen bonds help to make the beautiful arrangements that occur when snowflakes form.

Unusually High Surface Tension

Anyone who has done a belly flop into a lake or swimming pool knows that water packs a punch (**Figure 5**)! Its surface is almost like a skin: quite difficult to break through. This skin forms because of **surface tension**: the attractions between molecules at the surface of a liquid. Surface tension causes beads of water to form on smooth surfaces, and raindrops to form in free fall (**Figure 6**). Surface tension also explains how small objects that are denser than water can remain on the surface of water. Many aquatic insects make use of water's hydrogen bonds to skitter across the surface of the water.

surface tension a phenomenon, caused by forces of attraction between molecules, that leads to the formation of a skin-like film on the surface of a liquid

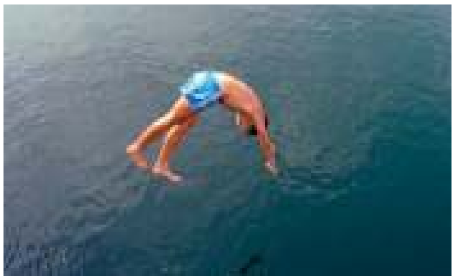


Figure 5 Hydrogen bonds can hurt!

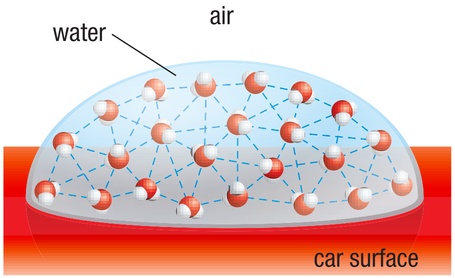


Figure 6 Water beads form on the surface of a car.

Investigation 3.5.1

Surface Tension and Intermolecular Forces (p. 125)
We are so familiar with water that we do not think of it as an unusual compound. This investigation illustrates just how different it is from other liquids.

What causes surface tension? As you know, hydrogen bonds are the main forces of attraction between water molecules. These hydrogen bonds allow the water molecules to hold strongly to each other, pulling together and resisting being broken apart. This causes the water to take a shape that has the smallest possible surface area. Other liquids, such as ethanol and mineral oil, do not form droplets in the same way that water does.

Unusually High Specific Heat Capacity

Water also holds thermal energy extremely well without changing its temperature a great deal. The quantity of the energy that a certain mass of a substance can absorb, and warm up by 1 °C, is called the specific heat capacity of the substance. You will learn more about this property in future chemistry courses. For now, you only need to know that lakes and oceans influence our climate and weather patterns because of the high specific heat capacity of water.

Mini Investigation

Exploring Surface Tension

Skills: Researching, Performing, Observing, Analyzing, Communicating

SKILLS HANDBOOK A1.2

Water has an unusually high surface tension, resulting in a sort of skin on the surface of water where water molecules hold tightly together. If the conditions are right, the water molecules can support an object with a greater density than water, such as a metal paper clip. In this investigation, you will explore this phenomenon and look at ways to reduce or increase the surface tension of water.

Equipment and Materials: chemical safety goggles; lab apron; two 400 mL beakers; glass stirring rod; 5 small metal paper clips; tissue paper; small samples of various substances, such as salt, sugar, dish soap, detergent, baking soda, vegetable oil, baby powder, and powdered sulfur

1. Put on your chemical safety goggles and lab apron.
2. Fill a beaker with water almost to the brim.
3. Drop a metal paper clip into the water from a height of about 10 cm. Record your observations.
4. Now, try to make a paper clip float.
5. If you have no success with Step 4, try the following technique: take a piece of tissue paper about 5 cm square. Gently drop the tissue flat onto the surface of the water. Carefully place a dry paper clip flat onto the tissue. Then use a glass stirring rod to poke the tissue until it sinks, leaving the paper clip floating.
6. Add one of the provided substances to the water. Record your observations.
7. Replace the water with clean water and repeat Step 6 with another substance.

A. Propose theoretical explanations for your observations in Steps 3, 4, 6 and 7. T/I

B. In Step 6, which of the added substances had an observable effect? What was that effect? T/I

C. Did any substances reduce the surface tension of the water? Did any substances increase the surface tension of the water? Propose a theoretical explanation for your observations in Step 6. Include the idea of surface tension in your explanation. T/I

D. How do we make use of substances that decrease the surface tension of water? Research at least two household or commercial applications of this phenomenon. T/I A

GO TO NELSON SCIENCE

3.5 Summary

- Water is extremely important for life on Earth.
- Water is a polar molecular compound that exhibits many unusual physical properties, such as high melting and boiling points, low density when solid, high surface tension, and high specific heat capacity.
- The theory of hydrogen bonding explains water's unusual properties.

3.5 Questions

1. (a) List four unusual physical properties of water.
   (b) Present a theoretical explanation for each of the properties. K/U
2. (a) Name the type of force that is primarily responsible for the unusual properties of water.
   (b) Draw a labelled diagram of the forces that connect water molecules. K/U C
3. Compare the arrangement of molecules in liquid water to that in ice. K/U
4. Consider each of the properties that you listed in Question 1. Predict how they would be different if water were a non-polar molecule. T/I
5. Imagine that water is a non-polar molecule. Brainstorm and list five specific ways that life on Earth would be different. Next, in a drawing or sketch, illustrate how these differences would look. Provide a written explanation for your drawing. T/I C A
6. Liquid water is most dense when it has a temperature of 4 °C. Ice floats on water. How do you think these two properties of water enable aquatic life to survive through a Canadian winter? A
7. Natural water sources and supplies are contaminated in a number of ways. Research the difference between point sources, diffuse sources, and indirect sources of water contamination and give two examples of each. T/I A

GO TO NELSON SCIENCE

Green Chemistry in Action: Choosing the Right Materials

3.6

Green chemistry is fairly new. As you have learned, understanding the bonding and structure of a compound can help us predict its properties and determine its uses. In the past, chemists developed products without considering their effects on the environment. The practice of "green chemistry" involves the invention, design, and use of products and processes that have minimal environmental impact. **Table 1** lists some of the aims of industries that have adopted these green chemistry principles.

Traditionally, companies first produced consumer goods and afterwards discovered any negative effects on the environment. Only then did the companies bring in cleanup technologies. Green chemistry, however, is proactive about designing products and processes that will not harm the environment. This is a better way to both ensure sustainability and minimize health risks. In this section, we will explore three specific cases of green chemistry in action.

Table 1 Some Aims of Green Chemistry

sustainability	starting with renewable materials rather than non-renewable (such as fossil fuels or mined materials)
safety	using and producing less toxic chemicals
process efficiency	using simpler reaction processes with fewer steps
energy efficiency	carrying out processes at lower temperatures or turning waste into usable energy
end-of-life degradation	designing products that degrade into harmless substances after use

Sustainable Materials: Bioplastics

To develop a green alternative, companies begin by considering what raw materials they need to produce their product. Research chemists try to choose renewable raw materials. They also aim to create a product that functions efficiently and presents no risk to human health or to the environment. Every stage of product development and production should be environmentally safe, including the handling of any by-products. The manufacturer should also consider what happens to the product at the end of its useful life.

As you learned in Section 3.1, most plastics are petrochemicals produced from fossil fuels. These plastics are non-biodegradable but many can be recycled at the end of their product life. Bioplastics, on the other hand, are plastics made from chemicals derived from plants (such as corn, potatoes, and peas).

The plastic cup in **Figure 1** is manufactured under the name "Greenware." It is made primarily from polylactate (PLA), which is synthesized from corn. The Greenware cup is marketed as "green" because it is compostable. This means that the cup will break down in a regulated industrial composting facility. Typically, at 55 °C and 90 % humidity, micro-organisms first break the intermolecular bonds and then decompose the molecules themselves. Greenware products break down in approximately 50 days. This technology was awarded the Presidential Green Chemistry Challenge Award in the U.S.A. in 2002.

Bioplastics manufacturers claim that their products significantly reduce greenhouse gas emissions by not using petrochemicals as a raw material. PLA production uses 20 % to 50 % less fossil fuels than the production of traditional plastics. Still, fossil fuels supply the energy for the production processes and for industrial composting at the end of the product's life cycle. Taking this into account, the "greenness" of these products is open to debate.

Other uses of bioplastics include cutlery, food containers, and other types of packaging. While there are many advantages to using bioplastics, there are also arguments against their use. The corn used to provide the raw material is being diverted from the food supply. This raises corn prices and puts pressure on the limited area of agricultural land. It may lead to food shortages in some parts of the world, where farmers can get a higher price for their crops from bioplastics producers than the local population can afford to pay. Some people claim that bioplastics contaminate already established recycling processes. As well, bioplastics are expensive because they are made by emerging technologies. The cost may, however, come down as the technologies become established.

CAREER **LINK**

To find out more about the work of a "green" research chemist,

GO TO NELSON SCIENCE

Figure 1 A bioplastic cup produced primarily from corn

Figure 2 The plastic coating on electrical wires must meet strict requirements for flame resistance.

CAREER **LINK**

If you are interested in the effects of chemicals on organisms in the environment, you might consider becoming an environmental scientist. To find out more about this career,

GO TO NELSON SCIENCE

WEB **LINK**

To find out more about the innovations of bioplastics companies,

GO TO NELSON SCIENCE

Less Toxic Materials: Flame-Resistant Bioplastics

We should all be concerned about the flammability and heat resistance of consumer products. For example, the plastic coating for electrical wires must be flame resistant to reduce the chance of electrical fires (**Figure 2**). As well, materials that are designed for use with hot objects must not melt, deform, or decompose. For example, a laptop computer case must be able to withstand high temperatures. The material must have relatively strong intermolecular forces. The flame-resistant materials that are typically used in these products are toxic. The electronics industry, for example, uses brominated flame retardants (BFRs) in its plastics. BFRs do not break down easily so they are persistent in the environment. They also enter the food chain (for example, by fish eating contaminated sediment) and bioaccumulate in the tissues of carnivores. Environmental scientists are concerned about the effects of these toxic chemicals on the environment and on human health.

Scientists are trying to replace these toxic plastics with bioplastics. PLA alone, however, is quite flammable. What would be a suitable additive, to reduce the flammability? Instead of using BFRs, scientists have developed a "green" alternative. This flame-retardant bioplastic uses a metal hydroxide flame retardant to absorb thermal energy. The energy is then unavailable for breaking intermolecular bonds. Tests show that this significantly improves the bioplastic's ability to withstand heat.

NEC scientists have also developed a bioplastic made from a combination of PLA and fibre from a plant called kenaf. This biodegradable "super plastic" is stronger, and has a greater ability to resist heat, than PLA alone. Such a material promises to be very useful in products where heat resistance is required, such as in electronic devices.

Recycled Materials: Clothing from Pop Bottles

How many pop bottles does it take to make a polar fleece jacket? Strange question? Not at all! For years an outdoor goods and gear company called Patagonia has been using recycled PET bottles to make their fleece products (**Figure 3**). Torontonians alone use about 100 million to 125 million plastic bottles every year. Approximately 50 % are recycled. The remaining bottles end up in landfill sites.

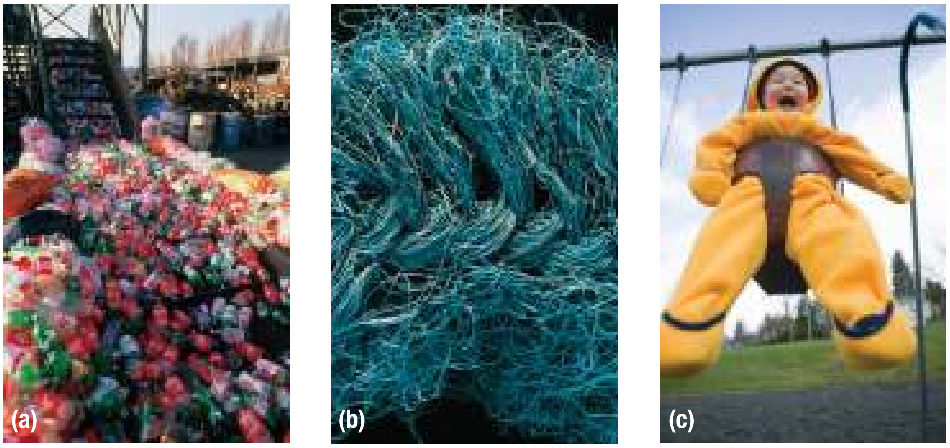


Figure 3 (a) Plastic bottles at a recycling facility (b) Fleece fibres are produced from molten PET—a type of polyester. (c) Fleece products are especially appropriate in Canada!

It takes about 25 pop bottles to make a fleece product. The used bottles are first sorted and cleaned, then heated until they melt. Additional petrochemicals are added to give the material its desired consistency and properties. The liquid is turned into threads by squeezing it through tiny holes in a metal plate. After cooling, these threads go through further steps to turn them into a warm, soft, and durable fabric.

Upcycling: High Fashion from Garbage

The next time you are shopping in a major retail store, watch out for the trashiest products around! A U.S.-based company, TerraCycle, collects non-recyclable materials and turns them into a variety of colourful products (**Figure 4**). Upcycling is the strategy of making something saleable from garbage. The aim of upcycling is to eliminate waste by finding new and useful purposes for garbage.

Figure 4 TerraCycle's motto, "Outsmart Waste," reflects the company's philosophy of making desirable goods from undesirable waste.

Research This

The Dirt on Cleaning

Skills: Researching, Analyzing, Communicating, Defining the Issue, Defending a Decision SKILLS HANDBOOK A1.1, A.5.1

Conventional cleaning products, such as window cleaners, oven cleaners, disinfectants, and drain cleaners, are effective but may be hazardous. This is indicated by a hazardous household product symbol (HHPS) on the label (**Figure 5**). Products with a "poisonous" symbol may contain toxic substances such as benzene, methanol, or formaldehyde. Some of these additives are responsible for long-term and short-term health problems including cancer, asthma, skin irritation, and headaches.

Figure 5 A hazardous household product symbol

There are environmentally friendly alternatives to these conventional cleaners (**Figure 6**). These include green cleaning products and natural cleaners. Natural cleaners include mixtures of substances such as baking soda, vinegar, olive oil, and borax.

Figure 6 How attractive a solution are green cleaners?

1. Research the health concerns related to conventional household cleaning products. Try to find articles and resources that consider the concerns in an objective fashion and that cite supporting evidence.

   Search for answers to the following: A

   - What regulations currently apply to cleaning products in Canada?
   - Identify three or four different types of conventional cleaning products.
   - What are the main ingredients in these products?
   - Research the possible health hazards associated with these cleaning products.
   - Identify green products that could be used in place of the conventional products.
   - What are the advantages and disadvantages of these green cleaning products?
   - What guarantee is there, if any, that a green product is safe?
   - Compare the prices of similar conventional and green cleaning products.

A. Analyze the research: Is it better to use a green cleaner or a conventional one? What criteria did you use to decide? Support your position with evidence. T/I A

B. Will people accept the higher cost of using a green product? Justify your position. A

GO TO NELSON SCIENCE

UNIT TASK BOOKMARK

Many of the ideas addressed in this section will be relevant as you consider your green product for the Unit Task. The task is described on page 134.

Other Areas of Interest in Green Chemistry

The main ideas of green chemistry are summarized in **Figure 7**.

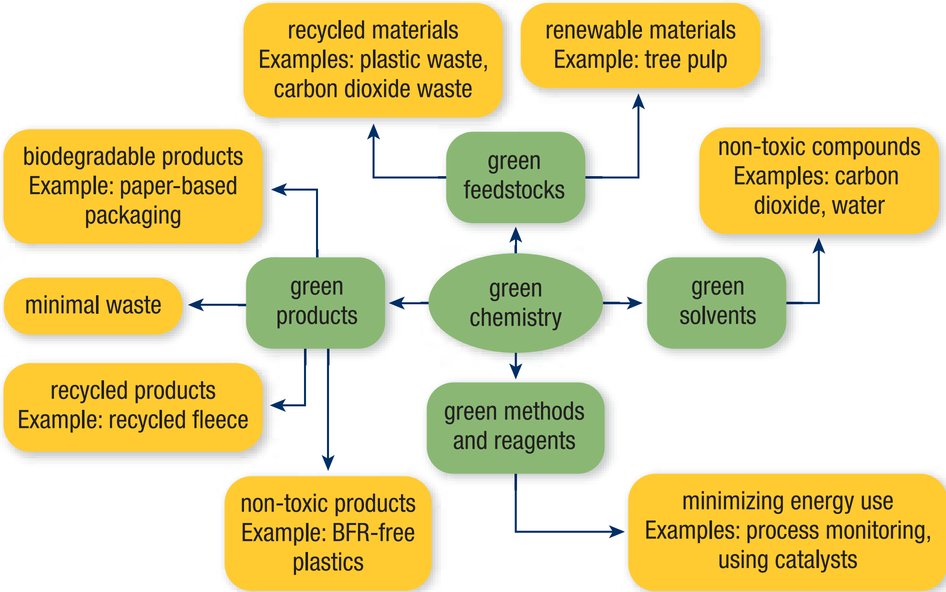


Figure 7

3.6 Summary

- Green chemistry practices include selecting materials that are renewable, non-toxic, or recycled, and considering the end of the product's life.
- The use of food plants to produce raw materials for consumer products is controversial.
- Many products are reusable, biodegradable, or recyclable.
- Making products from recycled materials helps to reduce the quantity of garbage going to landfills.
- Green cleaning products are increasingly available and effective.

3.6 Questions

1. List the four main principles of green chemistry outlined in this section. K/U
2. Consumers are "taxed" in several ways for using non-environmentally friendly products. For example, grocery stores in many municipalities now charge for plastic bags at the checkout. Garages add on a "disposal fee" for used motor oil or car tires. In your opinion, should harmful products be taxed so that they are more expensive than green products? Defend your position. A
3. Bioplastic made with kenaf fibre has an unusual property known as shape memory. Applications of materials with shape memory include self-adjusting orthodontic braces and self-tightening surgical sutures. Research bioplastics with shape memory. Describe some applications in a visual presentation. T/I C A
4. In 2005, Patagonia launched the Common Threads Recycling Program. Research how this program works, and report on the pros and cons of initiatives such as this. T/I A
5. Find out how TerraCycle solves a common problem for recycling industries: obtaining very specific items as its raw materials. T/I A
6. Research and examine the use of "eco-fees" on consumer services and products. Do you think they are a good idea? Give reasons for your decision. T/I A
7. Electrically conductive plastics are made from petrochemicals. These plastics have a much lower density than metal. The reduced mass of wiring in a car would lead to fuel savings. Research these plastics to find out what they are currently used for. List some of the pros and cons associated with their production and use. T/I C A

GO TO NELSON SCIENCE

Investigation 3.4.1 CONTROLLED EXPERIMENT

SKILLS MENU

- Questioning
- Researching
- Hypothesizing
- Predicting
- Planning
- Controlling Variables
- Performing
- Observing
- Analyzing
- Evaluating
- Communicating

Melting Points of Molecular Compounds

We suspect that intermolecular forces affect the melting point of a molecular compound. It seems logical that compounds with stronger intermolecular forces will have higher melting points. But is there a relationship between the size of the molecules and the strength of the forces between them?

In this investigation, you will measure the melting point of one of two molecular compounds. You will then combine your laboratory evidence with that of other research groups to see if a hypothesized relationship between melting point and molecular size is supported.

The molecular compounds used are

- lauric (dodecanoic) acid, $C_{12}H_{24}O_2(s)$ or $CH_3(CH_2)_{10}COOH(s)$ and
- stearic (octadecanoic) acid, $C_{18}H_{36}O_2$ or $CH_3(CH_2)_{16}COOH(s)$.

As you can see from their formulas, a stearic acid molecule is larger than a molecule of lauric acid.

You will research the melting points of four other similar molecular compounds and plot a graph of melting point versus number of electrons per molecule for all six compounds.

Testable Question

What is the relationship between melting point and molecular size?

Hypothesis

The melting points of similar molecular compounds increase as molecular size increase because more intermolecular bonds can form between larger molecules.

Variables

SKILLS HANDBOOK A2.2, A2,4

Read the Experimental Design and Procedure carefully to determine which is the independent (manipulated) variable and which is the dependent (responding) variable. In addition, be sure to identify any controlled variables.

Experimental Design

You will use hot- and cold-water baths to melt a solid compound and then to freeze and remelt it. You will measure the temperature at regular intervals during both the freezing and remelting processes. You will create a temperature-versus-time graph of your observations to determine the freezing point/melting point of the compound. You will compare your evidence with evidence from other groups investigating another similar compound and with information from other references. You will analyze this evidence to test the hypothesis.

Equipment and Materials

SKILLS HANDBOOK A1.1, A1.2

- chemical safety goggles
- lab apron
- hot plate
- wire gauze with ceramic centre
- two 400 mL beakers
- sample of lauric acid or stearic acid in a 15 × 150 mm test tube
- retort stand with test-tube clamp and ring clamp
- two thermometers or temperature probes
- timer that indicates seconds
- paper towel

This investigation involves very hot water. Be careful not to splash the hot water.

Avoid touching the hot-plate surface with your hands. Allow the hot plate to cool completely after turning it off and before moving it. To unplug the hot plate, pull on the plug itself rather than the cord.

The glass of a thermometer bulb is very thin and fragile. Be careful not to bump, pull, or twist the thermometer.

Procedure

1. Put on your chemical safety goggles and lab apron.
2. Add approximately 300 mL of water to a 400 mL beaker. Place the wire gauze and the beaker (secured within the ring clamp) on the hot plate. Gently heat the water to a temperature of about 80 °C. Use a thermometer or temperature probe to measure the temperature. Control the hot plate (turn it down, or switch it off and on) to keep the water hot but not boiling.
3. Place the test tube containing the solid sample in the beaker of hot water on the hot plate. Carefully clamp the test tube to the retort stand so that the test tube does not touch the bottom of the beaker.
4. Warm the test tube in the water bath until the sample is completely melted.

5. Fill the second 400 mL beaker with about 400 mL of cold tap water. Set the beaker on the other side of the retort stand from the hot plate.
6. Lift the test-tube assembly out of the hot water by raising the clamp. Rotate the clamp and test tube over to the beaker of cold water. Lower the test-tube assembly so that the test tube is suspended in the cold water.
7. Hold the thermometer against the side of the test tube so that the bulb is located about halfway between the bottom and the top of the liquid compound in the test tube (**Figure 1**). Keep the thermometer still. It will be held in this position when the liquid compound solidifies. Do *not* attempt to move the thermometer once it is immobilized by the solidified compound.

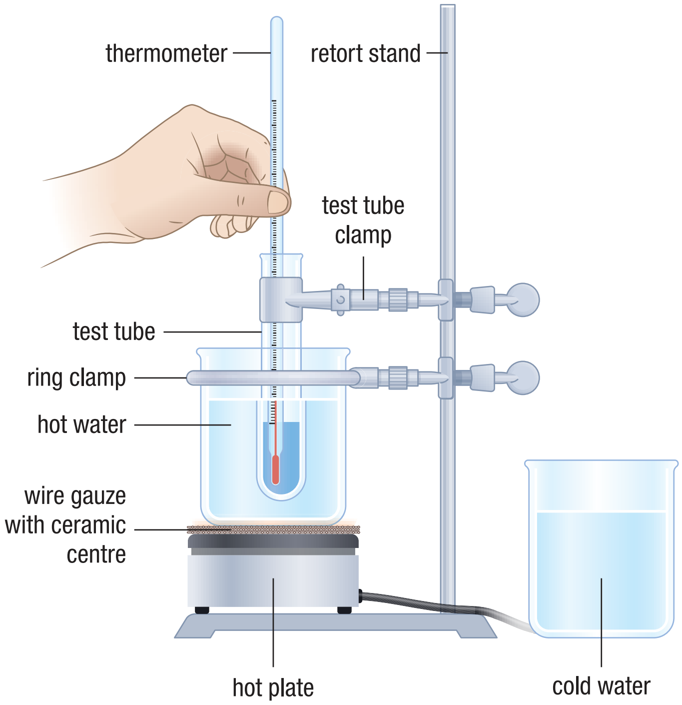


Figure 1 The cold-water bath

8. Measure and record the temperature as precisely as possible every 15 s as the compound cools.
9. When the compound has obviously been completely solidified for a few minutes, stop recording measurements.
10. Carefully lift the test-tube assembly by raising the clamp out of the cold-water beaker. Move the beaker of cold water onto the bench.
11. Rotate the clamp and test tube back to the beaker of hot water on the hot plate. Lower the test-tube assembly into the hot water and begin recording warming measurements by repeating Step 8.
12. When the compound has obviously been completely liquefied for a few minutes, stop recording measurements.
13. Remove the thermometer from the melted sample, and wipe it clean with a paper towel.
14. Repeat Step 6 to solidify the sample before returning it to your teacher.

Analyze and Evaluate

SKILLS HANDBOOK A6.4

(a) Identify the major variables that you measured and/or controlled in this investigation. Which variables were manipulated and which were responding? T/I

(b) Use your observations to plot both melting and freezing temperature–time graphs on the same axes. Interpret the graph shapes to determine the melting-point temperature of your compound as accurately as possible. T/I C

(c) Create a table in which to record the name, molecular formula, number of electrons per molecule, and melting point of six different compounds. Start completing the table with your observations for your own compound. Add the data from other groups for the other compounds. T/I C

(d) Research the melting points for the following molecular compounds: T/I
- capric acid, $CH_3(CH_2)_8COOH(s)$,
- myristic acid, $CH_3(CH_2)_{12}COOH(s)$,
- palmitic acid, $CH_3(CH_2)_{14}COOH(s)$, and
- arachidic acid, $CH_3(CH_2)_{18}COOH(s)$

Add these data to your table. Complete the table.

(e) Create a graph (using graph paper or graphing software) of melting point against number of electrons. T/I C

(f) Interpret your graph to answer the testable question. T/I

(g) Does your answer to the testable question support the hypothesis? Explain. K/U T/I

Apply and Extend

(h) Research the molecular structure of lauric acid and stearic acid. Would you classify these molecules as polar or non-polar? Or do these molecules have non-polar parts and polar parts? T/I

(i) Research the sources and uses of lauric acid and stearic acid. Comment on any differences in how these compounds are used that might be connected to the compounds' unique physical properties. T/I

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Investigation 3.5.1 CONTROLLED EXPERIMENT

Surface Tension and Intermolecular Forces

SKILLS MENU

- Questioning
- Researching
- Hypothesizing
- Predicting
- Planning
- Controlling Variables
- Performing
- Observing
- Analyzing
- Evaluating
- Communicating

In this investigation, you will look for a connection between the surface tension of various molecular liquids and the concepts of molecular polarity and intermolecular forces.

Testable Question

Read the investigation carefully. Write a testable question, related to surface tension and intermolecular forces, that you will attempt to answer during this investigation.

Hypothesis

SKILLS HANDBOOK A2.2, 2.4

Using your knowledge of intermolecular forces, write a hypothesis for your question. Your hypothesis should include a prediction and a theoretical reason for that prediction in the format "If . . . then . . . because . . . ."

Variables

Identify the variable that you will purposely manipulate, the variable that you will be measuring, and the variable(s) that you will keep controlled in this experiment.

Experimental Design

You will measure the relative surface tension of various molecular liquids by placing as many drops as possible of each liquid on the surface of a clean penny.

Equipment and Materials

SKILLS HANDBOOK A1.1, A1.2

- chemical safety goggles
- lab apron
- 15 clean copper pennies
- 5 small beakers
- dropper bottles of:
  - liquid detergent
  - water, $H_2O(l)$
  - ethanol, $C_2H_5OH(l)$
  - mineral oil
- paper towel

Ethanol is flammable. It should be used only in a well-ventilated area. There should be no open flames or other sources of ignition in the laboratory.

Procedure

1. Read the following steps carefully and create a table in which to record your qualitative and quantitative observations. Note that you cannot write a chemical formula for mineral oil because it is a mixture, not a compound.
2. Put on your chemical safety goggles and lab apron.
3. Clean the eye droppers and pennies thoroughly with liquid detergent and water. Rinse and dry them well.
4. Place a clean penny on a dry piece of paper towel.
5. Select one of the liquids. Place drops of the liquid, one by one, on the surface of the penny. Count the number of drops before the liquid spills. Record your observations.
6. Follow your teacher's instructions for disposal.
7. Repeat Steps 4 to 6 two more times, each time with a clean penny, giving three trials for that liquid.
8. Repeat Steps 4 to 7 for each of the other liquids.

Analyze and Evaluate

(a) List the manipulated, responding, and controlled variables in this investigation. C

(b) Rank the liquids according to how many drops of each you could fit on a penny. T/I C

(c) Based on your answer to (b), rank your liquids in order from the one with the greatest surface tension to the one with the least surface tension. T/I C

(d) Based on your observations, predict whether the molecules in the three liquids are polar or non-polar. Research to check your prediction. T/I

(e) Answer your testable question. T/I

(f) Does your evidence support your hypothesis? Explain. T/I

Apply and Extend

(g) List the types of bonds and forces in each of the substances that you tested.

(h) Mineral oil is not a pure substance. Rather, it is a mixture of many different compounds. Research to find out what compounds, or what kinds of compounds, mineral oil contains. A

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CHAPTER 3 SUMMARY

Summary Questions

1. Create a study guide based on the points listed in the margin on page 94. For each point, create three or four sub-points that provide further information, relevant examples, explanatory diagrams, or general equations.
2. Look back at the Starting Points questions on page 94. Answer these questions using what you have learned in this chapter. Compare your latest answers with those that you wrote at the beginning of the chapter. Note how your answers have changed.
3. Think about the possibility of carbon dioxide replacing perc as the world-standard dry-cleaning solvent. What physical properties make carbon dioxide a good cleaning agent? What are the benefits and risks of using carbon dioxide as a dry-cleaning solvent?
4. We have learned how we can predict the physical properties of molecular compounds. Give three specific examples where we have determined the use of a molecular compound because of its physical properties.

Vocabulary

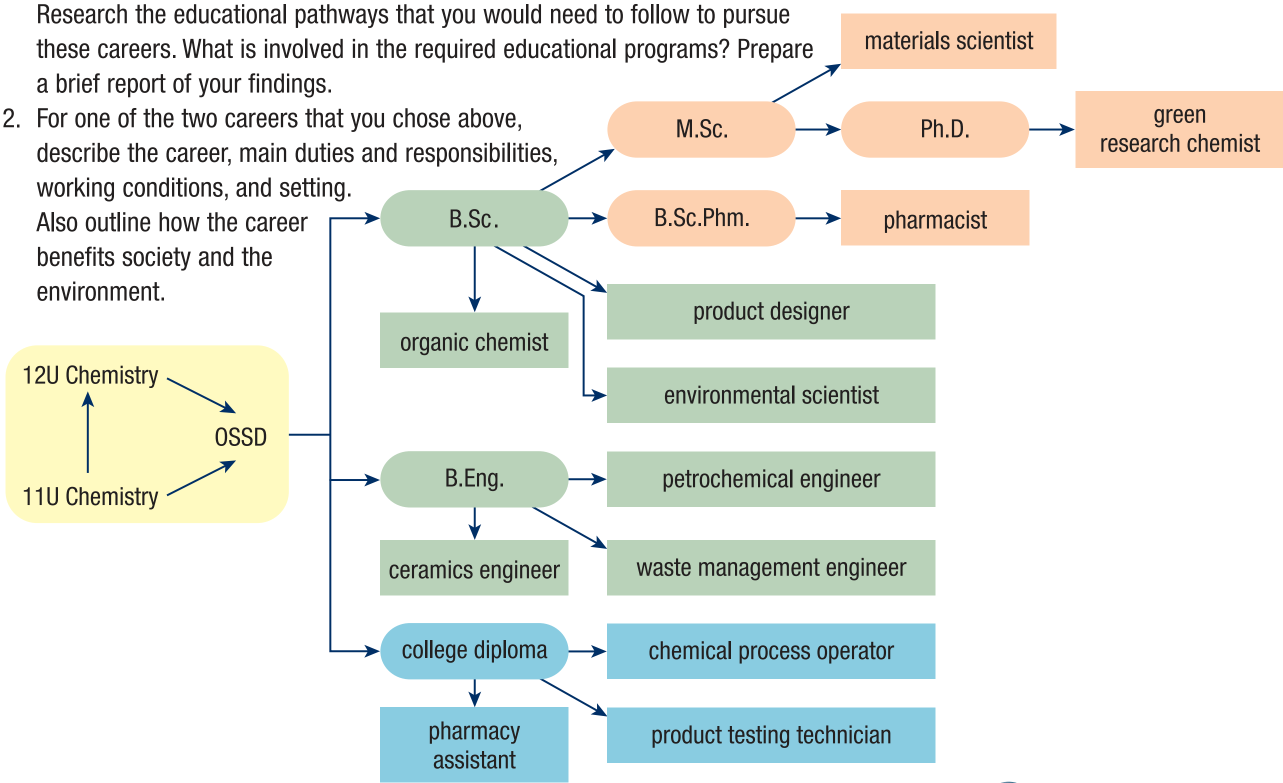
renewable resource (p. 96)	biodegradable (p. 97)	non-polar molecule (p. 104)	van der Waals forces (p. 111)
non-renewable resource (p. 96)	compostable (p. 98)	intermolecular force (p. 109)	hydrogen bond (p. 112)
petrochemical (p. 96)	upcycling (p. 98)	dipole–dipole force (p. 110)	surface tension (p. 117)
recycling (p. 97)	polar molecule (p. 104)	London dispersion force (p. 111)	

CAREER PATHWAYS

SKILLS HANDBOOK A7

Grade 11 Chemistry can lead to a wide range of careers. Some require a college diploma or a B.Sc. degree. Others require specialized or postgraduate degrees. This graphic organizer shows a few pathways to careers mentioned in this chapter.

1. Select two careers, related to molecular compounds, that you find interesting. Research the educational pathways that you would need to follow to pursue these careers. What is involved in the required educational programs? Prepare a brief report of your findings.
2. For one of the two careers that you chose above, describe the career, main duties and responsibilities, working conditions, and setting. Also outline how the career benefits society and the environment.



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CHAPTER 3 SELF-QUIZ

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

For each question, select the best answer from the four alternatives.

1. Which is an example of an upcycled product? (3.1) K/U
   (a) cups made of biodegradable plastic
   (b) sweaters made from llama wool
   (c) insulated lunch bags made from used candy wrappers
   (d) detergents made from petrochemicals
2. Which of the following characteristics of a material would make it most suitable for composting? (3.1) K/U
   (a) petroleum-based
   (b) biodegradable
   (c) recyclable
   (d) reusable
3. Which of the following products is most likely to contain DEET? (3.2) K/U
   (a) insecticides
   (b) insect repellents
   (c) herbicides
   (d) waterproofing agents
4. When poured in a thin stream, which of the liquids listed below would most likely be attracted to a charged vinyl strip that has been rubbed with a paper towel? (3.3) K/U T/I
   (a) methanol, CH_3OH
   (b) carbon tetrachloride, CCl_4
   (c) pentane, C_5H_{12}
   (d) carbon disulfide, CS_2
5. Which of the following substances has covalent bonds that are the most polar? (3.3) K/U
   (a) H_2O
   (b) H_2S
   (c) NH_3
   (d) CH_4
6. Which of the following substances will likely have the highest boiling point? (3.4) K/U T/I
   (a) butane, C_4H_{10}
   (b) pentane, C_5H_{12}
   (c) hexane, C_6H_{14}
   (d) heptane, C_7H_{16}
7. Dipole–dipole forces would be observed in
   (a) $F_2(g)$
   (b) $HF(g)$
   (c) $O_3(g)$
   (d) $CH_4(g)$ (3.4) K/U T/I
8. Which of the following is most important in helping the intertwining strands of DNA link properly? (3.4) K/U
   (a) polar covalent bonds
   (b) dipole–dipole forces
   (c) London dispersion forces
   (d) hydrogen bonds
9. Which of the following substances can be synthesized from corn? (3.6) K/U
   (a) PLA
   (b) PET
   (c) perc
   (d) CFCs

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

10. Fossil fuels and ore deposits are examples of renewable resources. (3.1) K/U
11. The bonding in nitrogen trifluoride, $NF_3(g)$, would be non-polar covalent. (3.3) K/U
12. In a molecule of oxygen difluoride, OF_2, the oxygen atom would have a partial positive charge. (3.3) K/U T/I
13. Of the molecules CO_2, SiH_4, and PF_3, the molecule with asymmetrically distributed bonds is CO_2. (3.3) K/U T/I
14. Of the three intermolecular forces, hydrogen bonding is the strongest. (3.4) K/U
15. Hydrogen bonds exist between all molecules. (3.4) K/U
16. Covalent and ionic bonds are considerably stronger than intermolecular forces. (3.4) K/U
17. Ice floats on water because water expands as it freezes. (3.5) K/U
18. Surface tension in water primarily results from hydrogen bonds. (3.5) K/U
19. Water's relatively high specific heat capacity results in a relatively large temperature change for a given amount of thermal energy absorbed. (3.5) K/U
20. Scientists increased the strength of PLA plastic by adding BFRs to it. (3.6) K/U

To do an online self-quiz,
GO TO NELSON SCIENCE

Knowledge

For each question, select the best answer from the four alternatives.

1. Which of the following is a non-renewable resource? (3.1) K/U
   (a) corn crops
   (b) solar energy
   (c) fossil fuels
   (d) rainwater
2. Which of the following substances could be a petrochemical? (3.1) K/U
   (a) plastic
   (b) adhesive
   (c) detergent
   (d) all of the above
3. Which of the following is true of DEET? (3.2) K/U
   (a) It is synthetically prepared.
   (b) It is a white and powdery solid.
   (c) It will not harm painted surfaces.
   (d) all of the above
4. The Canadian Pediatric Society recommends that DEET concentrations in insect repellents used for children should not exceed
   (a) 1 %
   (b) 5 %
   (c) 10 %
   (d) 30 % (3.2) K/U
5. If strips of plastic, one with a positive charge and the other with a negative charge, are held near a thin stream of water at different times, the stream of water will be
   (a) bent away from the positive strip and toward the negative strip
   (b) bent toward the positive strip and away from the negative strip
   (c) bent away from both charged strips
   (d) bent toward both charged strips (3.3) K/U
6. Which of the following molecules has only non-polar covalent bonding? (3.3) K/U
   (a) F_2 (c) NH_3
   (b) HCl (d) CF_4
7. Of the following attractions, which is generally the weakest attractive force between particles? (3.4) K/U
   (a) ionic bond
   (b) covalent bond
   (c) dipole–dipole force
   (d) London dispersion force
8. Hexane, C_6H_{14}(l), boils at 69 °C and heptane, C_7H_{16}(l), boils at 98 °C. Which would best account for the difference in the boiling points of these substances? (3.3, 3.4) K/U T/I
   (a) London dispersion forces
   (b) dipole–dipole forces
   (c) hydrogen bonding
   (d) all of the above
9. Which of the following substances is most likely to have hydrogen bonding? (3.4) K/U
   (a) ammonia, NH_3(g)
   (b) methane, CH_4(g)
   (c) phosphine, PH_3(g)
   (d) all of the above
10. Which of the following attractions is most important in holding the molecules together in the crystal structure of ice? (3.5) K/U
    (a) London dispersion forces
    (b) dipole–dipole forces
    (c) hydrogen bonds
    (d) ionic bonds
11. Which of the following is an environmentally unfriendly fire retardant used in plastics? (3.6) K/U
    (a) DNA (c) PET
    (b) BFRs (d) perc

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

12. Most of the compounds we encounter in today's world are molecular. (3.1) K/U
13. When individual drink pouches are used to make a backpack, this is an example of upcycling. (3.1) K/U
14. Citronella oil, garlic, geranium oil, and peppermint are examples of synthetic insect repellents. (3.2) K/U
15. DEET is suspected to have caused a small number of seizures in adults. (3.2) K/U
16. Of the polar covalent bonds H–F, H–Cl, and H–Br, the most polar is H–Br. (3.3) K/U
17. For the polar covalent bond between a silicon atom and a chlorine atom, the partial charges would be shown as δ^-Si–Clδ^+. (3.3) K/U
18. Carbon dioxide sublimates at low temperatures, so it must have weak intermolecular forces. (3.4) K/U
19. Boiling and sublimating are chemical changes. (3.4) K/U
20. The strongest intermolecular attractions in nitrogen trifluoride are London dispersion forces. (3.4) K/U

21. Butane, C_4H_{10}, is the only substance on Earth that commonly exists as a solid, a liquid, and a gas. (3.5) K/U
22. Lakes and oceans influence climate because of water's high specific heat capacity. (3.5) K/U

Match each term on the left with the most appropriate description on the right.

23. (a) polar covalent bond
    (b) ionic bond
    (c) non-polar covalent bond

    (i) $0.0 < \Delta EN < 1.7$
    (ii) $\Delta EN = 0.0$
    (iii) $\Delta EN \geq 1.7$ (3.3) K/U

Write a short answer to each question.

24. We frequently make use of compounds that are produced by micro-organisms. List three categories of these substances that benefit us. (3.1) K/U
25. List four materials that are commonly recycled in many countries. (3.1) K/U
26. Distinguish between an insect repellent and an insecticide. (3.2) K/U
27. What organization originally developed DEET and for what reason? (3.2) K/U
28. In your notebook, write "ionic bonds" or "covalent bonds" as appropriate in front of each of the phrases below. (3.3) K/U
    (a) form between atoms
    (b) form between ions
    (c) are created by a transfer of valence electrons between entities
    (d) are created by the sharing of pairs of valence electrons between entities
    (e) hold together the atoms within molecules
29. Do ionic compounds have intermolecular forces? Explain your answer. (3.4) K/U
30. What are the two main reasons for the formation of hydrogen bonds? (3.4) K/U
31. What relationship could there be between outdoor fleece clothing and plastic bottles? (3.6) K/U

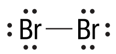
Understanding

32. Could the wood in a forest become a non-renewable resource? Explain your answer. (3.1) K/U
33. How can the strong, very stable bonds in most plastics be both good and bad? (3.1) K/U
34. Why would you want to avoid getting a DEET-based insect repellent spray on your clothes? (3.2) K/U
35. Distinguish between a polar covalent bond and a polar molecule. (3.3) K/U
36. "Electron density is greater near the more electronegative atom in a polar covalent bond." Explain this statement. (3.3) K/U
37. Which structure in **Figure 1** is the better representation of the three-dimensional shape of a water molecule? Explain your choice. (3.3) K/U

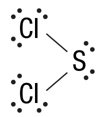
(a) (b)

Figure 1

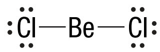
38. Arrange the following sets of bonds in order of decreasing polarity. Refer to the electronegativity values in the periodic table. (3.3) K/U
    (a) Cl–Cl, H–Br, N–F
    (b) C–N, C–S, As–O
39. Draw the "most polar" bond in each set of bonds from Question 38 and use the symbols δ^+ and δ^- to indicate the partial positive charges and partial negative charges. (3.3) K/U C
40. Determine whether the molecules of each substance listed below are polar or non-polar. Explain your reasoning in each case. For each molecule with one or more polar bonds, copy the Lewis structure into your notebook and add the symbols δ^+ and δ^- to indicate the predicted partial electric charge on each atom. (3.3) K/U T/I C
    (a) bromine, $Br_2(l)$



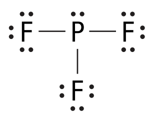
    (b) sulfur dichloride, $SCl_2(l)$



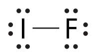
    (c) beryllium dichloride, $BeCl_2(s)$



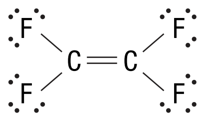
    (d) phosphorus trifluoride, $PF_3(g)$



    (e) iodine fluoride, IF(s)



    (f) tetrafluoroethene, $C_2F_4(g)$



41. Solid carbon dioxide, dry ice, is made of exactly the same molecules as gaseous carbon dioxide. What is the difference between these two forms of carbon dioxide? Use the terms "intermolecular forces" and "energy" in your answer. (3.4) K/U

42. **Figure 2** shows two attractive forces that exist in gaseous hydrogen bromide, HBr(g). One force exists between neighbouring molecules and the other exists within each molecule. (3.4) K/U T/I
    (a) Name each force in Figure 2.
    (b) Which of the two forces is stronger? What experimental evidence supports your answer?

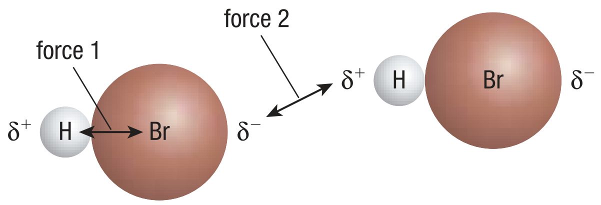


Figure 2

43. Of hydrogen bonding, dipole–dipole forces, and London dispersion forces, which exists between all molecules and atoms? Explain your answer. (3.4) K/U

44. Compare and contrast London dispersion forces, and dipole–dipole forces between molecules. (3.4) K/U

45. At room temperature, water, H_2O, is a liquid and hydrogen sulfide, H_2S, is a gas. (3.4, 3.5) K/U C
    (a) Compare the boiling points of the two compounds.
    (b) Draw a Lewis structure for each substance.
    (c) How do the molecular shapes of H_2O and H_2S molecules compare?
    (d) Account for the difference in boiling points of water and hydrogen sulfide using the concept of intermolecular forces.

46. During winter, why is it important to drain water pipes in unoccupied, unheated homes? (3.5) K/U A

47. Why do lakes freeze from the top down? (3.5) K/U

48. Thermal energy flows more rapidly from one material to another when their temperature difference is greater. A good coolant, therefore, can absorb large quantities of thermal energy without undergoing a drastic temperature increase. Considering these facts, explain why water is often chosen as a coolant. (3.5) K/U

49. List the qualities that a product should have if it is to be environmentally harmless from its manufacture to the end of its useful life. (3.1, 3.6) K/U

50. Why are products made from recycled materials generally "greener" than those made from new materials? (3.6) K/U

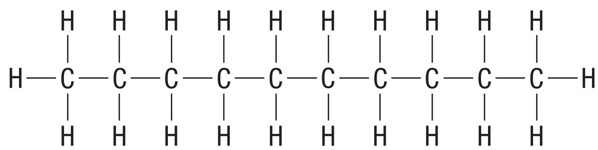
Analysis and Application

51. A newspaper columnist writes an article in which she urges Canadians to use less gasoline by driving smaller, more fuel-efficient cars and by walking or biking whenever possible. The columnist writes, "When we drive big cars and SUVs, we might just as well throw our clothes and many of our household products into the gas tank and learn to live without them." Explain this statement. (3.1) T/I A

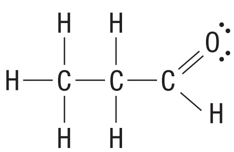
52. List three necessary chemical or physical properties of each of the following products: (3.1) T/I A
    (a) the hose for a gasoline pump at a gas station
    (b) adhesive for patching the holes in inner tubes and inflatable pool toys
    (c) the fabric of a camping tent

53. Suppose you are a researcher working for a chemical manufacturer and you have been asked to develop new scented insect repellents. You need to find substances that effectively repel insects and impart a pleasant odour to the product. Would you start your research by looking at ionic compounds or at molecular compounds? Explain your thinking. (3.2) K/U T/I

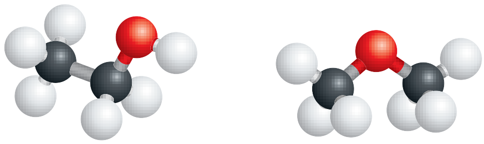
54. Predict what will happen when a thin stream of each liquid below is passed near a vinyl strip that has been rubbed vigorously with a paper towel to give it a negative charge. Explain each prediction. Include diagrams in your explanations. (3.3) K/U C
    (a) decane, $C_{10}H_{22}$



    (b) propanal, C_3H_6O



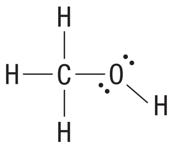
55. Ethanol, $C_2H_5OH(l)$, is used as a fuel in gasohol blends. Methoxymethane, $C_2H_6O(g)$, can also be used as a fuel. Ethanol boils at 78 °C and methoxymethane boils at −24 °C. Explain the difference in boiling points between these two substances by referring to intermolecular forces (**Figure 3**). (3.3, 3.4) T/I



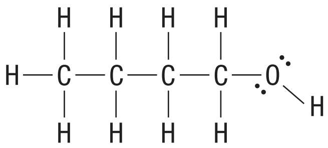
ethanol methoxymethane

Figure 3

56. Methanol, $CH_3OH(l)$, is used for the industrial production of many other chemicals. Butan-1-ol, $C_4H_9OH(l)$, is used in perfumes and to extract essential oils from plant matter. Methanol boils at 65 °C and butan-1-ol boils at 118 °C. Use the concepts of intermolecular forces to explain the higher boiling point of butan-1-ol. **Figure 4** shows the Lewis structure for each substance. (3.3, 3.4) K/U T/I C



methanol



butan-1-ol

Figure 4

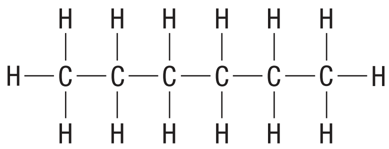
57. Innovative cleaning companies are using special washing machines that use liquid carbon dioxide as the washing liquid. At the end of a wash cycle, almost all of the carbon dioxide is recaptured for repeated use. Why would cleaning clothes with liquid carbon dioxide result in energy savings when the clothes are "dried" in the washing chamber? (3.4) K/U A

58. **Table 1** shows the boiling points of the hydrogen halides: the hydrogen compounds of the Group 17 elements. Use the concept of intermolecular forces in your answers to the following questions: (3.4) T/I C
(a) Why is the boiling point of hydrogen fluoride higher than the other boiling points?
(b) Why do the boiling points increase as one moves down the family?

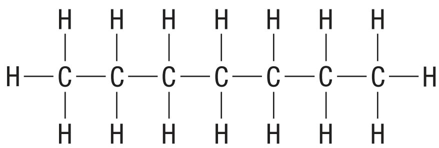
Table 1 Boiling Points of Hydrogen Halides

Compound	Boiling point (°C)
HF	20
HCl	−85
HBr	−67
HI	−36

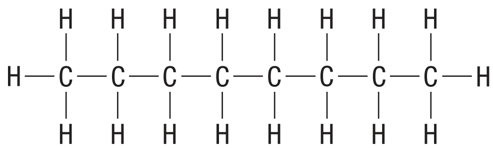
59. Petroleum is a mixture of many different hydrocarbon compounds. The industrial separation of the components of petroleum relies on differences in boiling points. Three typical hydrocarbons obtained from petroleum are hexane, C_6H_{14}, heptane, C_7H_{16}, and octane, C_8H_{18}. **Figure 4** shows the Lewis structure for each hydrocarbon. Use the concept of intermolecular forces to predict the trend in the boiling points of these substances. (3.4) K/U T/I



hexane



heptane



octane

Figure 4

60. Imagine gently floating a dry plastic clip from a bread wrapper on the surface of a container of water (**Figure 5**). If you gently place a single drop of detergent onto the water in the hole at the back of the bread clip, the clip surges forward like a motorboat. Explain this observation using the concept of surface tension. (3.4, 3.5) K/U T/I

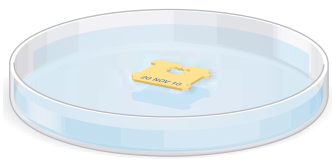


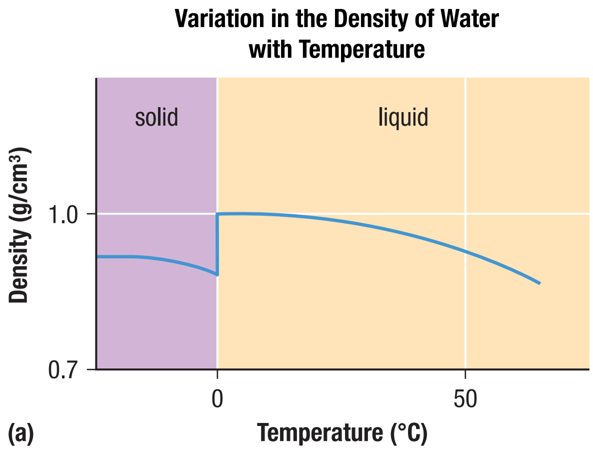
Figure 5

61. When soap is added to water, the soap molecules reduce the intermolecular attractions between the water molecules. How would this affect the ability of water to form beads on a horizontal surface? Why? (3.4, 3.5) T/I A

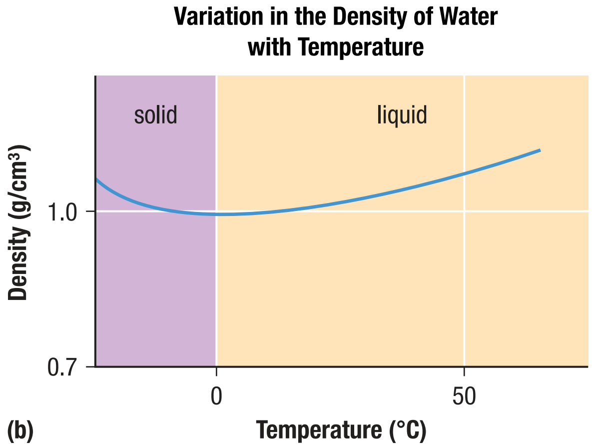
62. The Nanticoke petroleum refinery in southwestern Ontario produces asphalt that is used to pave streets and parking lots. Pavers must seal freshly paved streets to prevent moisture from seeping into the asphalt. Why would water penetrating the asphalt be a problem when the temperatures fall below freezing? (3.5) K/U A

63. A 100 cm^3 sample of liquid water freezes. (3.5) T/I C
    (a) Calculate the volume of the resulting ice. Assume that the density of liquid water at its freezing point is 1.00 g/cm^3 and that the density of the ice is 9 % less than that of liquid water. Show and explain your calculations.
    (b) Explain how your answer to (a) might affect water pipes during the winter.

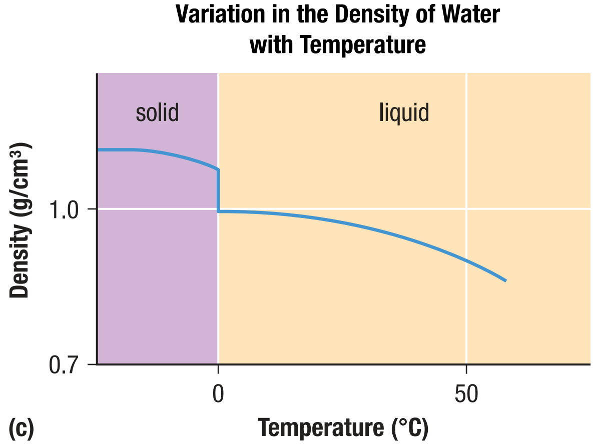
64. Three graphs are shown in **Figure 6** below. Which graph best represents how the density of water varies with temperature? Explain your selection. (3.5) K/U T/I



(a)



(b)



(c)

Figure 6

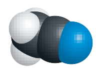
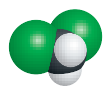
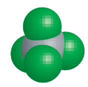
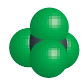
65. Bioplastics offer environmental advantages over traditional petroleum-based plastics. At the same time, bioplastics are an emerging technology and are generally more expensive to produce than petroleum-based plastics. In your opinion, should governments provide tax breaks or subsidies to manufacturers of bioplastics to help offset production costs? Explain your thinking. (3.6) K/U T/I A

Evaluation

66. You are a researcher for a laboratory that tests consumer products. Describe four important tests you would carry out to evaluate which of several reusable water bottles is best. (3.1) T/I A

67. An organic chemist wants to carry out a synthesis reaction in a liquid. The reaction requires a non-polar solvent and a reaction temperature of 60 °C. Evaluate the suitability of the possible solvents listed in **Table 2**. Which solvent should the chemist choose? Defend your choice. (3.3) T/I A

Table 2 Boiling Points of Various Solvents

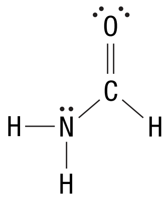
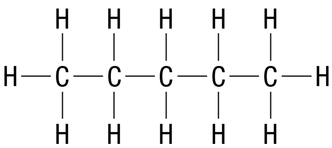
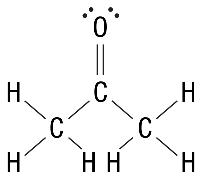
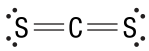
Solvent	Molecular structure	Boiling point (°C)
acetonitrile, C_2H_3N		81
dichloromethane CH_2Cl_2		40
silicon tetrachloride, $SiCl_4$		58
dimethyl sulfoxide, C_2H_6OS		189
carbon tetrachloride, CCl_4		77

68. A capacitor is a component in electronic devices such as flight simulators. The capacitor stores electric charge. It has two metal plates separated by an electrically insulating material. The insulator works best when its molecules have charged ends. A researcher is investigating liquid insulators for capacitors and is considering the liquids shown in **Table 3**. (3.3) K/U T/I A

(a) Evaluate the liquids in Table 3 with regard to the degree to which their molecules have charged ends.

(b) Recommend which liquids the researcher should test as possible insulators for capacitors. Explain your recommendations.

Table 3 Possible Insulator Liquids for Capacitors

Liquid	Molecular formula	Lewis structure
methanamide	CH_3NO	
pentane	C_5H_{12}	
propanone	C_3H_6O	
carbon disulfide	CS_2	

69. A materials engineer is considering building computer cases out of pure PLA plastic. (3.6) K/U

(a) Would pure PLA be a good choice for this product?

(b) What kind of modified PLA plastic might be more suited for the computer cases? Explain why.

70. If the world were to substitute bioplastics for petroleum-based plastics on a large scale, many crops that are now grown for food might be diverted to the manufacturing of the bioplastics. Outline at least one potential problem with this use of food crops, and briefly describe further research topics within this problem. (3.6) K/U T/I A

Reflect on Your Learning

71. After reading this chapter, are you now more likely to buy upcycled products? Why or why not? K/U A

72. Considering what you have learned in this chapter about insect repellents, are you more likely or less likely to use an insect repellent when outdoors? Use specific examples to explain your answer. A

73. Electronic products like cellphones and portable music or media players go "out of date" very quickly, as manufacturers promote new models and features. How has this chapter changed your opinion on the marketing strategy of electronic retailers, if at all? A

Research

GO TO NELSON SCIENCE

74. The Peel Integrated Waste Management Facility in Ontario houses a "single-stream" waste processing facility. Research how single-stream recycling works, focusing on the Peel facility. Report on the advantages and disadvantages of single-stream recycling. Include diagrams in your report. K/U C A

75. You have learned that polylactic acid, PLA, is an environmentally friendly plastic because it is made from renewable resources and is biodegradable. Prepare a pamphlet or audiovisual presentation of your research on the following topics: K/U C A

(a) Research the molecular structure of PLA and the chemical reactions used to produce it.

(b) Research the pros and cons of PLA plastic products.

76. Although ice exists in nature almost entirely in one crystalline structure, a number of different forms of ice form at high pressures. These forms of ice are called phases. Each phase has its own characteristic arrangement of water molecules. Research the various phases of ice and report your findings. Include a diagram of the conditions under which the various phases form and also provide diagrams of the relevant lattice structures. K/U C

77. Bisphenol A is a plastic additive that is a perfect illustration of how industry develops and markets products only to find later that the products may be unsafe to people and the environment. Research the history of the use of BPA and the current regulations regarding its use in Ontario. Based on your findings, write a statement of opinion about the banning of BPA from plastic products. Support your argument. K/U A

It's Not Easy Being Green

As you wander through the grocery store, peruse the aisles of a hardware store, search for back-to-school supplies, or browse products online, you may notice something. More and more products are being advertised as "green" (**Figure 1**). Increasingly, buyers are deciding between green and conventional products. For example, you might come across green cleaning products. Or you may need to decide whether to use green bioplastic grocery bags or reusable green carrying bags.

Figure 1 There are many choices when it comes to carrying the groceries home.

What does it mean when a product is declared "green"? What does the expression "benign by design" really mean? Throughout Unit 1 you have learned a lot about the principles of green chemistry. Specifically, you have learned how producers can make safer and more environmentally friendly products by following the principles of green chemistry:

- choosing the least-toxic chemicals for the production process
- adopting sustainable processes by using renewable or recycled raw materials and by minimizing waste products
- using renewable energy sources
- saving energy by carrying out processes at lower temperatures wherever possible
- addressing disposal issues by designing products that degrade into harmless substances after use

New green products and processes are developed daily in a wide variety of fields. (See Figure 6 in Section 3.6.) Are there regulations and standards for green products? Is there a greener alternative to every consumer product?

You are a member of the Millennial generation, born at the end of the twentieth century. Your generation is more environmentally educated than any before you (**Figure 2**). How are companies targeting you as a potential consumer?

Figure 2 What makes you choose a green product rather than a conventional one?

In this task, you will plan, develop, design, and market a green product. You will be purposeful in your design. The green-ness of your product will illustrate at least one of the principles of green chemistry. You will conduct extensive research on the properties of your product and on the production processes. You will contrast your product with the conventional alternative. You will put the product to the ultimate test: convincing investors to support its production. They are only likely to do so if they believe that consumers will buy it.

The Issue

How can you convince an investor to support the production of a "green" consumer product?

Goal

Your aim is to develop a green product that investors are willing to support. Your product could be

- clothing
- cleaning products
- paints
- printer paper
- grocery bags
- electronic products
- laundry detergent
- toilet paper
- water bottles
- diapers
- books
- pesticides
- office supplies
- disposable plates, cups, and cutlery
- batteries
- dry-cleaning solvents
- fuels
- any other product you might be interested in

Research

In your group, conduct research to learn more about

- the principles of green chemistry
- any regulations or restrictions on companies marketing products as "green"
- what makes green products appealing to investors and consumers
- a selection of products that you are interested in working on

The Conventional Product

With the information you have gathered, narrow down your choices to one or two products that you are interested in pursuing.

Conduct more focused research on a conventional version of the product you are considering, including

- its chemical and physical properties
- manufacturing processes and energy sources
- waste materials or by-products
- packaging
- its likely fate at the end of its useful life
- safety and ease of use, price, packaging, and marketing strategies
- the relative "green-ness" and desirability of the conventional product

Identify Solutions and Make Decisions

Now you will design a green alternative to the conventional product. Similar green alternatives can exist in the marketplace at the same time, so you do not need to come up with something totally new—unless you really want to challenge yourself!

The Green Product

For your green alternative, research the same seven points that you investigated for the conventional product. As you compare the two products, develop responses to the following questions:

- What makes your product greener?
- What makes your production processes greener?
- What makes your green alternative more attractive to a consumer?

Communicate

You may have seen the CBC television program *Dragons' Den*, in which inventors, entrepreneurs, and product designers try to sell their ideas to a panel of potential investors. You will take your turn in the dragons' den! In your group, plan and present your "pitch" to a panel of investors (selected classmates). Your pitch must include

- an attention-grabbing introduction
- the information about your green product, contrasted with the conventional options
- a clear and concise evaluation of your product's green-ness
- a wrap-up to convince the investors to back your product

After listening to your pitch, the investors will confer and give their answers.

ASSESSMENT **CHECKLIST**

Your completed Performance Task will be assessed according to the following criteria:

Knowledge/Understanding

- ✓ Understand some of the principles of green chemistry and the criteria for green consumer products.
- ✓ Research and classify the chemical and physical properties of the green and traditional consumer products.
- ✓ Research and classify the chemical and physical properties used to manufacture the green and traditional consumer products.

Thinking/Investigation

- ✓ Compare the green-ness of the green product against its traditional counterpart.

Communication

- ✓ Prepare an informative, creative, and engaging report that includes all the specified points.

Application

- ✓ Demonstrate an understanding of the negative effects posed by consumer products to our health and to the environment.

For each question, select the best answer from the four alternatives.

1. Wind is matter because it has which two properties? (1.1) K/U
   (a) mass and motion
   (b) mass and volume
   (c) volume and space
   (d) mass and sound
2. Which statement is part of Dalton's atomic theory? (1.2) K/U
   (a) Atoms can be created or destroyed.
   (b) All atoms are identical.
   (c) All matter is made up of atoms.
   (d) Atoms can be divided into smaller parts.
3. What does an atom become when it loses an electron? (1.3) K/U
   (a) a positively charged anion
   (b) a negatively charged anion
   (c) a positively charged cation
   (d) a negatively charged cation
4. Which of the following pairs of atoms are isotopes? (1.4) K/U
   (a) K-39 and K-40
   (b) F-19 and Ne-19
   (c) C-12 and N-14
   (d) H-3 and He-3
5. Which is shown in a Bohr–Rutherford diagram but not shown in a Lewis diagram? (1.5) K/U
   (a) all the electrons in the atom
   (b) the symbol of the element
   (c) the valence electrons
   (d) the location of the element on the periodic table
6. Which scientist is given credit for developing the periodic table upon which the current periodic table is based? (1.6) K/U
   (a) Newlands
   (b) Meyer
   (c) Moseley
   (d) Mendeleev
7. Which property of an element decreases as you proceed from left to right across a row on the periodic table? (1.7) K/U
   (a) atomic radius
   (b) ionic radius
   (c) ionization energy
   (d) electron affinity
8. What makes up an ionic compound? (2.1) K/U
   (a) one anion and one cation
   (b) huge numbers of anions and cations
   (c) molecules
   (d) uncharged atoms
9. What type of bond holds the atoms in a molecule together? (2.2) K/U
   (a) ionic
   (b) metallic
   (c) covalent
   (d) Lewis
10. What is the ability of an atom, when bonded, to attract electrons to itself? (2.3) K/U
    (a) bond character
    (b) ionization energy
    (c) electron affinity
    (d) electronegativity
11. Which is the correct chemical formula for the ionic compound magnesium chloride? (2.4) K/U
    (a) $MgCl$
    (b) Mg_2Cl
    (c) Mg_2Cl_2
    (d) $MgCl_2$
12. What is the correct name of the compound whose chemical formula is SnO? (2.4) K/U
    (a) tin oxide
    (b) tin(I) oxide
    (c) tin(II) oxide
    (d) tin(IV) oxide
13. Which of the following compounds is an example of an artificial sweetener? (2.5) K/U
    (a) sucralose
    (b) glucose
    (c) sucrose
    (d) fructose
14. Which of the following is an example of a non-renewable resource? (3.1) K/U
    (a) fossil fuels
    (b) wood
    (c) water in a river
    (d) bacteria
15. Which of the following insect repellents is a synthetic compound? (3.2) K/U
    (a) garlic
    (b) citronella oil
    (c) geranium oil
    (d) DEET

16. What determines how polar a polar bond is? (3.3) K/U
    (a) whether the bonded atoms have gained electrons or lost electrons
    (b) the electronegativity difference of the bonded atoms
    (c) the difference in ionization energy of the bonded atoms
    (d) the size of the bonded atoms
17. Which of the following properties is characteristic of a compound that contains weak intermolecular forces? (3.4) K/U
    (a) gas at room temperature
    (b) extremely high melting point
    (c) soft and flexible
    (d) liquid at room temperature
18. Which of the following is the strongest type of intermolecular force? (3.4) K/U
    (a) polar force
    (b) dipole–dipole force
    (c) London dispersion force
    (d) hydrogen bond
19. Which of the following is an important physical property of water? (3.5) K/U
    (a) low melting point
    (b) high boiling point
    (c) low surface tension
    (d) low density compared to its solid form
20. For which of these products would a company most likely try to develop a green alternative? (3.6) K/U
    (a) a product made from renewable raw materials
    (b) a product that functions efficiently
    (c) a product that presents a potential problem to human health
    (d) a product that decomposes into environmentally safe substances
21. Which of the following is an example of a green solvent? (3.6) K/U
    (a) alcohol
    (b) biodegradable plastic
    (c) recycled fleece
    (d) water

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

22. Laws change over time as new knowledge is discovered. (1.1) K/U
23. Thomson discovered the electron by using a cathode ray tube. (1.2) K/U
24. According to the octet rule, atoms are stable when they contain full valence shells. (1.3) K/U
25. A group is a horizontal row on the periodic table. (1.5) K/U
26. Mendeleev arranged his periodic table according to increasing atomic number. (1.6) K/U
27. Ionization energy is the amount of energy needed to remove a valence electron from an atom in the gaseous state. (1.7) K/U
28. Solid ionic compounds conduct an electric current. (2.1) K/U
29. A molecular element is made of at least two different kinds of elements. (2.2) K/U
30. Elements with similar electronegativities form a covalent bond. (2.3) K/U
31. The name of the molecular compound NO is nitrogen oxide. (2.4) K/U
32. Research involving synthetic molecules has provided many options for alternative sweeteners. (2.5) K/U
33. High fructose corn syrup is an example of a carbohydrate sweetener. (2.5) K/U
34. Both recycling and upcycling involve converting wastes into new products. (3.1) K/U
35. Repellents contain compounds that kill insects outright. (3.2) K/U
36. You can predict the approximate melting point of a molecular compound if you know how polar it is. (3.3) K/U
37. A non-polar molecule cannot contain polar bonds. (3.3) K/U
38. Intermolecular forces are stronger than chemical bonds. (3.4) K/U
39. Hydrogen bonds are important in determining the shape and function of large molecules, such as proteins. (3.4) K/U
40. Because of hydrogen bonds, water has unusually high melting and boiling points. (3.5) K/U
41. Recycling is the strategy of making a high-value retail product from something that otherwise would have been garbage. (3.6) K/U

Knowledge

For each question, select the best answer from the four alternatives.

1. Which of the following is an example of a theory? (1.1) K/U
   (a) The gravity of Earth pulls objects toward its surface.
   (b) An atom is made up of smaller particles called protons, neutrons, and electrons.
   (c) For every action, there is an equal and opposite reaction.
   (d) Mass is not created or destroyed in normal changes.
2. Which scientist discovered that an atom contains a dense core, which became known as a nucleus? (1.2) K/U
   (a) Bohr
   (b) Chadwick
   (c) Rutherford
   (d) Thomson
3. Which is the correct sequence for analyzing a sample using a mass spectrometer? (1.4) K/U
   (a) accelerate and deflect the ions → ions hit a detector plate → a computer analyzes the results → ionize the sample
   (b) ionize the sample → accelerate and deflect the ions → ions hit a detector plate → a computer analyzes the results
   (c) ionize the sample → ions hit a detector plate → accelerate and deflect the ions → a computer analyzes the results
   (d) accelerate and deflect the ions → ions hit a detector plate → ionize the sample → a computer analyzes the results
4. An element is a soft, silver-coloured solid. It reacts violently and forms a basic solution when it is dropped into water. The element is most likely a member of which of the following groups? (1.5) K/U
   (a) halogens
   (b) alkali metals
   (c) alkaline earth metals
   (d) noble gases
5. Which scientist published an early chemistry book that listed several chemical elements? (1.6) K/U
   (a) Antoine Lavoisier
   (b) Hennig Brand
   (c) Johann Dobereiner
   (d) John Newlands
6. Which of these properties describes an ionic compound? (2.1) K/U
   (a) solid, liquid, or gas at ambient temperatures
   (b) low melting point
   (c) conducts a current in any state
   (d) high boiling point
7. Which types of substances can exist as molecules? (2.2) K/U
   (a) certain elements and compounds
   (b) all compounds
   (c) all elements
   (d) all elements and all compounds
8. The difference in electronegativity between two bonded atoms determines what type of bond forms. Which electronegativity difference indicates the presence of an ionic bond? (2.3) K/U
   (a) 0
   (b) <1.7
   (c) 1.7
   (d) ≥ 1.7
9. Aluminum chloride is an example of what type of compound? (2.4) K/U
   (a) binary ionic
   (b) polyatomic ionic
   (c) hydrate
   (d) molecular
10. Suppose a hydrogen atom forms a polar covalent bond with another atom. Which of the following elements will form the most polar bond with hydrogen? Use the electronegativity values in the periodic table in the back of this textbook to help you. (3.3) K/U
    (a) carbon, C
    (b) oxygen, O
    (c) chlorine, Cl
    (d) bromine, Br
11. What is the name for the attraction that exists between molecules? (3.4) K/U
    (a) covalent bond
    (b) ionic bond
    (c) intermolecular force
    (d) electrical attraction
12. What is the only substance that regularly exists in all three states of matter on Earth's surface? (3.5) K/U
    (a) air
    (b) water
    (c) oxygen
    (d) hydrogen

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

13. IUPAC is an organization that was established to standardize knowledge and procedures in chemistry. (1.1) K/U
14. Bohr–Rutherford diagrams show only the valence electrons in an atom. (1.2) K/U
15. When an atom gains an electron, it becomes a cation. (1.3) K/U
16. Isotopes are atoms of the same element that have different numbers of electrons. (1.4) K/U
17. Meyer's proposed periodic table left spaces for yet undiscovered elements. (1.6) K/U
18. Atomic radius is a measurement of the size of an atom. (1.7) K/U
19. Ionic compounds form from the attraction between anions and cations. (2.1) K/U
20. A polar covalent bond forms when the electronegativity difference between the atoms is 0. (2.3) K/U
21. Upcycling involves taking a used product and making new, high-value items from it. (3.1) K/U
22. Polar molecules have a negatively charged end and a positively charged end. (3.3) K/U
23. The ability of one substance to dissolve into another substance depends on the polarity of the two substances. (3.4) K/U
24. Bioplastics are frequently recycled into other useful products. (3.6) K/U

Match each term on the left with the appropriate chemical formula on the right.

25.	(a) sodium chlorate	(i)	$NaClO$
	(b) sodium chlorite	(ii)	$NaClO_2$
	(c) sodium hypochlorite	(iii)	$NaClO_3$
	(d) sodium perchlorate	(iv)	$NaClO_4$ (1.3, 2.5) K/U

Write a short answer to each question.

26. What is chemistry? (1.1) K/U
27. How did the results of Rutherford's experiment suggest that an atom contains a dense, positively charged core? (1.2) K/U
28. Write the formula for the ion that forms from each of the following atoms. If necessary, draw Lewis symbols or Bohr–Rutherford models of each atom. (1.3) K/U
    (a) Br
    (b) N
    (c) K
    (d) Mg
    (e) S
29. Name each of the following polyatomic ions: (1.3) K/U
    (a) SO_4^{2-}
    (b) NO_3^-
    (c) CO_3^{2-}
    (d) BrO_3^-
    (e) ClO_4^-
    (f) CN^-
    (g) SO_3^{2-}
30. Write the formulas for each of the following polyatomic ions: (1.3) K/U
    (a) chlorate
    (b) ammonium
    (c) hydrogen carbonate
    (d) hypochlorite
    (e) dichromate
    (f) permanganate
    (g) hydrogen sulfate
31. Write a sentence that correctly relates the following terms: radioisotopes, nuclear radiation, radioactive decay. (1.4) K/U
32. The term "periodic law" relates observations made about different elements and their location on the periodic table. (1.5) K/U
    (a) In your own words, state the periodic law.
    (b) Periodic law is often defined in terms of periodic trends. What is a periodic trend?
33. What contribution to the periodic table was made by Johann Dobereiner? (1.6) K/U
34. If an atom has high ionization energy, what is true about its electron affinity? (1.7) K/U
35. (a) What is an electrolyte?
    (b) Why are ionic compounds electrolytes? (2.1) K/U
36. In which state do both ionic and molecular compounds exist under ambient temperatures? (2.2) K/U
37. (a) What is a covalent bond?
    (b) How does a covalent bond differ from an ionic bond? (2.2) K/U
38. What is the difference between electronegativity and electron affinity? (2.3) K/U
39. Name each of the following ionic compounds: (2.4) K/U
    (a) $NaBr$
    (b) MgS
    (c) $CuOH$
    (d) $SnCl_2$
    (e) K_2SO_3

40. Name each of the following molecular compounds: (2.4) K/U
    (a) SO_3
    (b) $AsCl_3$
    (c) NO
    (d) CCl_4
    (e) P_2O_5
41. Write the formula for each of these ionic compounds: (2.4) K/U
    (a) sodium fluoride
    (b) calcium chloride
    (c) mercury(I) oxide
    (d) potassium cyanide
    (e) ammonium sulfate
    (f) magnesium nitrate
    (g) barium phosphate
    (h) magnesium sulfite
    (i) nickel(II) perchlorate
    (j) copper(I) sulfide
42. Write the formula for each of these molecular compounds: (2.4) K/U
    (a) sulfur dioxide
    (b) carbon tetraiodide
    (c) silicon dioxide
    (d) phosphorus trichloride
    (e) dinitrogen tetroxide
43. What are the two sources of most molecular compounds? (3.1) K/U
44. Contrast reusing and recycling. (3.1) K/U
45. Why is there a need for scientists to develop products such as insect repellents? (3.2) K/U
46. In a molecule of hydrogen bromide, would you expect the electron density to be greater around the hydrogen atom or around the bromine atom? (2.3) K/U
47. Under what circumstances is a molecule that contains polar bonds a non-polar molecule? (3.3) K/U
48. (a) List the three different types of intermolecular forces.
    (b) Which of these types are also referred to as van der Waals forces? (3.4) K/U
49. The following three compounds all have similar molecular masses: water, H_2O; ammonia, NH_3; and methane, CH_4. Why is water a liquid at normal temperatures, yet methane and ammonia are gases? (3.5) K/U
50. What property of water causes it to form beads and droplets? (3.5) K/U
51. What does it mean to say that one aim of "green chemistry" is to increase energy efficiency? (3.6) K/U

Understanding

52. You cannot see thoughts or air. (1.1) K/U T/I
    (a) Are thoughts matter? Explain your answer.
    (b) Is air matter? Explain your answer.
53. Suggest a reason why neutrons were discovered later than electrons and protons. (1.2) T/I
54. Copy and complete **Table 1** in your notebook. (1.2) K/U T/I

Table 1 Atomic Data for Three Elements

	Cobalt	Iron	Silver
atomic number	27		
mass number	59	56	
number of protons		26	47
number of neutrons			61

55. (a) What is an octet?
    (b) State the octet rule.
    (c) How does the octet rule relate to the formation of ions? (1.3) K/U
56. When mercury forms a compound, the mercury atom forms an ion with a valence of either +1 or +2. Using the IUPAC system of naming, name the following compounds: (1.3) K/U T/I
    (a) HgCl
    (b) HgO
    (c) $HgNO_3$
57. (a) What are the three most common types of nuclear radiation?
    (b) Describe each type of nuclear radiation. (1.4) K/U
58. Naturally occurring uranium is composed of three major isotopes. Most uranium is uranium-238, which makes up 99.28 % of all natural uranium. Because it undergoes a known pattern of decay, U-238 is important in dating objects. Uranium also exists as uranium-235, which makes up 0.71 % of all natural uranium. U-235 is the form of uranium used in nuclear reactors. Small amounts of uranium-234 (0.0054 %) are also present. From this information, calculate the atomic mass of uranium. (1.4) T/I
59. Examine Figure 3 in Section 1.5. Imagine that the square representing boron, B, is empty and does not contain a diagram. (1.5) T/I
    (a) How might you use the squares to the right and the left of the empty square to draw the Bohr–Rutherford diagram for boron?
    (b) How might you use the square for aluminum, Al, to check that your Bohr–Rutherford diagram is correct?
60. Contrast the modern periodic table and the periodic table proposed by Dmitri Mendeleev. (1.6) K/U T/I
61. Are Dobereiner's triads more similar to a group on the periodic table or a period? Explain your answer. (1.7) K/U T/I

62. (a) What is a chemical bond?
    (b) What is an ionic bond?
    (c) Why is an ionic bond a type of chemical bond? (2.1) K/U T/I
63. Refer to the structure of ionic compounds to explain the following properties of ionic compounds: (2.1) K/U
    (a) high melting point
    (b) hardness
    (c) breaks when struck with a hammer
    (d) electrical conductivity when dissolved
64. The type of covalent bond formed depends on the number of electron pairs shared. For each type of covalent bond below, explain the bond and how it forms. Give an example of each type of bond. (2.2) K/U
    (a) single covalent bond
    (b) double covalent bond
    (c) triple covalent bond
65. Using Figure 5 in Section 2.3 and the electronegativity values on the periodic table in the back of this textbook, identify what type of bond forms when each of these pairs of atoms reacts: (2.3) K/U
    (a) H and C
    (b) N and O
    (c) K and Br
    (d) O and O
    (e) H and O
66. Compare and contrast non-polar covalent bonds and polar covalent bonds. (2.3) K/U T/I
67. Molecular compounds do not follow the same method of naming used by ionic compounds. Why is a different naming convention necessary? (2.4) T/I
68. (a) How does the energy content of an artificial sweetener compare to that of sugar?
    (b) How does the sweetness of the two types of compounds compare? (2.5) T/I
69. Why are compostable or biodegradable products generally considered to be more environmentally friendly than non-biodegradable products? (3.1) K/U T/I
70. (a) Why do you think insect repellents are often applied directly to the skin?
    (b) Why are insecticides usually applied to the environment instead of directly on a person? (3.2) T/I A
71. What two factors determine whether or not a molecule is polar? (3.3) K/U
72. (a) Explain what dipole–dipole forces are.
    (b) What types of molecules contain the strongest dipole–dipole forces? (3.3, 3.4) K/U T/I
73. (a) Which type of intermolecular force exists between all molecules, whether they are polar or not?
    (b) Are these forces more noticeable between polar molecules or between non-polar molecules? Explain your answer. (3.4) K/U T/I
74. Create a graphic organizer of your choice to demonstrate the relationships among the following terms and any other relevant terms:
    - covalent bond
    - ionic bond
    - chemical bond
    - van der Waals forces
    - dipole–dipole forces
    - London dispersion forces
    - hydrogen bond
    - intermolecular bond
    - single covalent bond
    - double covalent bond
    - electronegativity
    - polar bond
    - polar molecule (3.4) K/U C
75. "Green chemistry" involves materials that are less toxic than formerly used materials. In what two ways must toxicity be considered in developing and using green materials? (3.6) T/I A

Analysis and Application

76. Explain how the electron energy levels proposed by Niels Bohr are similar to the rungs on a ladder. (1.2) T/I
77. A small quantity of the element selenium is essential to good health in humans. Among other purposes it is used for proper functioning of the thyroid gland. However, larger quantities of this element are toxic. Most selenium atoms contain 34 protons and 45 neutrons. (1.2) K/U T/I
    (a) What is the atomic number of selenium? Explain your answer.
    (b) Explain how to calculate the mass number of a selenium atom. Then, calculate it.
78. Radioisotopes can be hazardous. However, they also have many uses. (1.4) K/U A
    (a) What do all isotopes used in medical diagnosis and treatment have in common?
    (b) List three useful radioisotopes. Explain how they are used.
79. Examine Figure 1 in Section 1.5. Note that hydrogen is shown in the same column as the alkali metals. (1.5) K/U T/I
    (a) Why is hydrogen shown in the same column of the periodic table as the alkali metal group?
    (b) Why is hydrogen not considered to be an alkali metal?

80. The periodic trend of atomic radius is different from the periodic trend followed by ionization energy, electron affinity, and electronegativity. (1.7) T/I
    (a) Describe how these trends differ.
    (b) Give a theoretical explanation for the difference in the trends.
81. The size of an atom is not the same as the size of an ion formed from the atom. For both of the following elements, explain what happens when the ion forms, and compare the size of an atom and an ion: (1.7) K/U T/I A
    (a) copper
    (b) sulfur
82. Suppose you want to know how easily an atom forms an ion. For each of the following elements, would you be more concerned with its electron affinity or its ionization energy? Explain each answer. (1.7) T/I A
    (a) bromine
    (b) magnesium
83. Potassium chloride, KCl, is used to make fertilizer, in processing food, and as a salt substitute. (2.1) K/U A
    (a) Name and describe the structure of solid potassium chloride.
    (b) List four properties of potassium chloride that are common to other ionic compounds.
84. Propanol, or propyl alcohol, has the chemical formula C_3H_7OH. Propanol is used in disinfectants such as rubbing alcohol. Draw a Lewis structure and a structural formula for a molecule of propanol. (2.2) T/I C A
85. Carbonates are compounds that contain the carbonate polyatomic ion, CO_3^{2-}. They are the main components of limestone and mollusk shells. Draw a Lewis structure and a structural formula for the carbonate ion. (2.2) T/I
86. Think about the different types of ions you learned about in Section 1.3. How is it possible that an ionic compound might contain both ionic and covalent bonds? Provide an example of such a compound in your explanation. (2.3) K/U T/I A
87. For each of the following compounds, write the chemical formulas for both the anhydrous compound and the hydrate: (2.4) K/U A
    (a) Cobalt(II) chloride is blue. When it becomes hydrated, the resulting cobalt(II) chloride hexahydrate is deep rose in colour. As a result of this dramatic colour change, this compound is frequently used to show the presence of water.
    (b) Magnesium sulfate, when hydrated, forms magnesium sulfate heptahydrate. This hydrate is commonly known as Epsom salt, a product of British Columbia. This hydrate is dissolved in warm water and used to reduce inflammation in human tissue.
88. Sucrose is table sugar, glucose is blood sugar, fructose is fruit sugar, and lactose is milk sugar. If you are checking a food label to see if the food contains sugar, what would you look for to show the presence of a sugar? (2.6) T/I A
89. Do you agree with the statement that reusable products are generally more environmentally friendly than any comparable product? Explain your answer. (3.1) T/I C A
90. Compare and contrast renewable resources and non-renewable resources. In your answer, provide two examples of each type of resource. (3.1) K/U T/I
91. Methane, CH_4, is the main component of natural gas. (3.3) K/U T/I
    (a) Is a methane molecule polar or non-polar?
    (b) Explain your answer.
92. Environmentalists often mention the "three Rs": reduce, reuse, recycle. (3.6) T/I A
    (a) Explain what each "R" refers to in daily living, giving relevant examples.
    (b) Rank the "Rs" in order of decreasing importance with regard to protecting the environment. Give reasons for your rankings.
93. Methane, ethanol, and paraffin are all molecular compounds. At ambient temperatures, methane is a gas, ethanol is a liquid, and paraffin is a soft solid. Use this evidence to predict the strengths of the intermolecular forces in these compounds at ambient temperatures. (3.4) K/U T/I A
94. Some of Earth's best sport fishing is found in remote lakes in Canada's Yukon. The sub-arctic climate in this location means that most of these lakes have a thick layer of ice on them for much of the year. What property of water prevents the water in most of these lakes from freezing all the way down to the bottom during the long winter? (3.5) K/U A
95. On a sunny summer day a beach heats up faster than the adjacent water, but at night the beach cools off faster than the water. Explain these observations, referring to at least one property of water. (3.5) K/U T/I A
96. Which uses energy more efficiently, recycling or upcycling? Explain your answer. (3.6) T/I A

Evaluation

97. (a) Explain the difference between theoretical knowledge and empirical knowledge.
    (b) Describe an example of theoretical knowledge.
    (c) Use an example to explain how empirical knowledge can solve a problem. (1.1) K/U T/I A
98. The currently used periodic table classifies elements in many ways. Explain how each of the following terms is used to classify elements on the periodic table: (1.5) K/U A
    (a) metals
    (b) non-metals
    (c) metalloids
    (d) group
    (e) period
99. Imagine a Bohr–Rutherford diagram of a fluorine atom and a Lewis symbol of the same atom. (1.5) T/I
    (a) What do the diagrams have in common?
    (b) When might it be most helpful to use a Bohr–Rutherford diagram?
    (c) When might you choose to use a Lewis symbol instead of a Bohr–Rutherford diagram?
100. (a) What is effective nuclear charge?
    (b) How does effective nuclear charge affect atomic radius?
    (c) How does it affect electron affinity? (1.7) K/U T/I
101. Why are Lewis symbols helpful when representing the formation of an ionic bond? (2.1) T/I C
102. Use the electronegativity values from the periodic table in the back of this textbook to answer the following questions: (2.3) T/I A
    (a) Magnesium and iodine react to form magnesium iodide, MgI_2. Compare the electronegativities of the two elements.
    (b) What type of bond forms between a magnesium atom and an iodine atom?
103. Aluminum sulfide, Al_2S_3, is a binary ionic compound that releases toxic fumes when it is heated or mixed with water. Why is it called a binary compound? (2.4) K/U
104. Several types of intermolecular forces attract molecules to each other. (3.4) K/U
    (a) What type of intermolecular force acts between all molecules?
    (b) What type of intermolecular force acts only between polar molecules?
    (c) What type of intermolecular force acts between polar molecules that contain hydrogen and a highly electronegative element?
105. (a) In general, what force can be used to explain the unique properties of water?
    (b) Several properties of water are unusual, compared to the properties of other compounds of similar mass. List five of these properties. (3.5) K/U
106. Plain water penetrates dry, clay-laden soil poorly. When sprinkled on such soil, the water forms beads and will not seep into cracks. Imagine that you are a researcher at a company that makes lawn and garden supplies. You want to create a product that can be applied to lawns with dry clay soils. Your goal is for water that is sprinkled on the soil to mix with the product and, as a result, create a solution that more readily penetrates and dampens the soil. (3.5) K/U T/I A
    (a) What property of water must your product affect? Explain your thinking.
    (b) What substance might you consider as the main ingredients for your product? Explain your choice.
    (c) What would you have to find through research or experimentation to be sure that your product will be effective, safe, and environmentally friendly?
107. A relative wants to compost not only kitchen waste and yard trimmings, but also biodegradable plastics such as PLA cups. He plans to buy a standard composting bin that he will put in his backyard. He will let the composting mixture stand in ordinary environmental conditions. His hope is that he will have fully digested compost within a month. Evaluate this plan. Are his expectations reasonable? Explain your answer. If you believe that your relative's plan is not sound, offer an alternative. (3.6) K/U T/I A
108. It requires time, effort, and expense to change from materials and processes that have been used for many years to green materials and processes. What are three advantages of converting to green materials and processes? (3.6) T/I A

Reflect on Your Learning

109. (a) What is meant by the statement, "Chemistry is called the central science"?
    (b) Give an example of how chemistry connects physics and biology.
    (c) Give an example of how chemistry connects biology and Earth science. K/U T/I A
110. Think about what you know about a polyatomic ion. T/I A
    (a) What other terms do you know that contain the prefix *poly*-?
    (b) How does the meaning of *poly*- in these terms relate to the meaning of this prefix in "polyatomic ion"?

111. Examine Table 3 in Section 1.3. This table shows several ions, their sources, and their importance to human health. K/U T/I A

   (a) From the sources listed, how many of the ions in the table have you ingested today?

   (b) How do you think what you have eaten today affects your health? Explain, using at least two examples.

112. Radon is a serious health problem when it is present in a home or other building. Examine Figure 8 in Research This in Section 1.4. From the diagram, list four different ways the homeowner could prevent radon gas in the soil from entering the house. T/I A

113. Most noble gases do not react because their atoms already have complete valence shells. However, large noble gas atoms can be made to form compounds with active non-metals, such as oxygen and fluorine. T/I A

   (a) Xenon, Xe, reacts with fluorine to form XeF_6. How many valence electrons are around Xe in this compound? Explain your answer.

   (b) Does this compound follow the octet rule? Explain your answer.

   (c) Why do you think large noble gas atoms form such compounds but small noble gas atoms, such as those of He or Ne, do not?

114. Section 2.3 compares differences in electronegativity to a tug-of-war. Describe another analogy that might be used to explain the effects of differences in electronegativity between bonded atoms. T/I A

115. Choose two compounds that are usually known by their common names, not their IUPAC names. For each compound, provide both its common name and its IUPAC name. K/U

116. Think about consumer products that were made by upcycling. Describe one such product. Include in your description at least three reasons why you think the product was made by upcycling instead of by some other method. T/I A

117. There are both risks and benefits associated with many products and scientific processes. Discuss with a partner the risks and the benefits of using insect repellents that contain DEET. T/I C A

118. Suppose you comb your hair on a cold, dry day. You might notice that some of your hair now stands away from your head and is attracted to the comb. Electrons are transferred from the comb to your hair. The comb becomes positively charged and your hair becomes negatively charged. What do you predict will happen if you run a thin stream of water from a tap and hold the comb alongside? Explain. K/U T/I A

119. In this unit you encountered several ways to represent various entities and how they form compounds. List at least three different ways, and explain a benefit of each one. T/I

120. Impurities often affect the properties of substances. For example, salt added to water lowers its freezing point, raises its boiling point, and affects its surface tension. Why do you think impurities affect these properties of water? K/U T/I A

Research

GO TO NELSON SCIENCE

121. Benzene is a chemical in petroleum. It is also manufactured in chemical plants in Ontario, Alberta, and Québec. T/I A C

   (a) What are some uses of benzene?

   (b) Describe some risks associated with the production of benzene.

   (c) What are some of the responsibilities that go along with the production, transportation, and use of benzene?

122. Some transition metals have only one valence, but many others have multiple valences. The properties and uses of the compounds containing the element vary depending on the valence of the element in that compound. Research the following transition metals and answer the following questions: A

   (a) Compare and contrast the properties and uses of copper(I) chloride, CuCl, and copper(II) chloride, $CuCl_2$.

   (b) Although chromium has valences that include every number from –2 through +6, its most common valences are +3 and +6. What are the names given to chromium when it has each of these valences? What ions contain chromium in each of these valences? How are the different forms of chromium related to human health?

   (c) Historically, mercury and its compounds had numerous medicinal uses. The toxicity of these substances now limits their uses, especially for children. What are some former medical uses of mercury, mercury(I) chloride, and mercury(II) chloride?

123. The British Columbia Cancer Agency (BCCA) is now producing its own radioisotopes. Research this development and answer the following questions: T/I A
   (a) What isotope is BCCA now producing?
   (b) Why did it need this isotope?
   (c) How is it now able to produce the isotope?
   (d) By having the ability to produce radioisotopes, what future advances in cancer diagnosis and treatment might be possible?
124. When IUPAC assigns permanent names to elements, they sometimes decide that the common names should be used as the permanent names. Other names have been chosen to honour different things, such as people or locations. Research each of the following elements. Find out what each name honours. Explain why you think each person or location was important enough to earn this honour. K/U
   (a) yttrium
   (b) uranium
   (c) lawrencium
   (d) ruthenium
   (e) curium
125. Canadian chemist Raymond Lemieux was a pioneer in the study of sugars. T/I A
   (a) Research Lemieux's work and list at least three of his achievements.
   (b) Lemieux's work demonstrates the worldwide effect of scientific research. Research the awards Lemieux received. Which countries honoured this scientist?
   (c) Lemieux investigated oligosaccharides. Research the roots of the different parts of the word. How do the root meanings relate to the structures of the molecules?
126. The charged ends of a polar molecule might be compared to the ends, or poles, of a magnet. T/I A
   (a) Research the poles of a magnet. Find out what they are called and why these names were given to them. How do similar poles of magnets act when they are brought near each other? How do opposite poles of magnets act when they are brought near each other?
   (b) How do things that have the same electrical charge, such as the negative ends of two polar molecules, act when they are near each other? How do two objects that have opposite charges act when they are near each other?
   (c) Use your research results from (a) and (b) of this question to explain why the terms "pole" and "polar" apply both to magnets and to electrically charged items.
   (d) Water molecules are polar. From your research, how do you think water molecules are arranged in a sample of liquid water?
127. Research bioplastics to find answers to the following questions: K/U T/I A
   (a) What is the source for approximately half of the bioplastics produced?
   (b) Technically, all plastics are biodegradable by microbes under the right conditions. Why are traditional plastics considered to be non-biodegradable?
   (c) Microbes are not the only cause of biodegrading of plastics. What other factors can cause biodegrading of plastics?
   (d) How might bioplastics have an advantage over other plastics in the field of medicine?
   (e) How might bioplastics be used effectively in agriculture?

UNIT 2 Chemical Reactions

OVERALL EXPECTATIONS

- analyze chemical reactions used in a variety of applications, and assess their impact on society and the environment
- investigate different types of chemical reactions
- demonstrate an understanding of the different types of chemical reactions

BIG IDEAS

- Chemicals react in predictable ways.
- Chemical reactions and their applications have significant implications for society and the environment.

UNIT TASK **PREVIEW**

The challenge in the Unit Task is to reclaim toxic copper ions from a dump site that contains mine tailings. How will you recover the copper and treat the remaining solution so that it is safe to return to the environment?

The Unit Task is described in detail on page 242. As you work through the unit, look for Unit Task Bookmarks to see how information in the section relates to the Unit Task.

GOING GREEN WITH CHEMISTRY

Green chemistry is good for business. This is the conclusion reached by many manufacturers that rely on chemical processes to transform raw materials into finished goods. Green chemistry, if you recall, is an industry trend to make industrial processes that involve chemicals more environmentally friendly and sustainable. Going green might involve making a process more energy efficient. For example, the chemical reactions involved in making steel need a high-temperature furnace. To maintain these temperatures the furnaces burn huge quantities of fossil fuels. This makes the steel industry a major emitter of greenhouse gases and other air pollutants. Updating to newer, more energy-efficient furnaces allows companies to produce more steel with less energy. More product and reduced energy consumption means the company is more profitable. And we all benefit from a cleaner environment.

Green chemistry can also involve the total redesign of a manufacturing process. One example is the use of corn to make clear plastic-like film used for food wrappers. Traditionally, clear plastic wrap is made from petroleum: a non-renewable resource. Cargill Dow, a chemical company based in the United States, developed a process to make clear wrap that uses sugar—a renewable resource—instead. Sugar is first extracted from corn. Bacteria then use the sugar to produce lactic acid, which is purified and further processed into the plastic-like film. An additional advantage of this product is that it is biodegradable.

In this unit you will examine other creative ways in which green chemistry can make chemical processes greener and more sustainable.

Questions

1. How does reducing energy costs make a chemical process more profitable?
2. What are the advantages of using renewable resources as the starting materials of manufacturing processes? What possible disadvantages may there be?
3. Do you think is it more profitable for a company to prevent pollution from occurring, or to deal with pollution after it is produced? Defend your position.
4. What role do you think governments should play in encouraging green chemistry in industry? Why?

CONCEPTS

- chemical and physical properties of a substance
- the law of conservation of mass
- balanced chemical equations
- environmental effects of chemical reactions
- properties of acids and bases

SKILLS

- name or write the chemical formula of an ionic or molecular compound
- write a balanced chemical equation for a given reaction
- research and collect information
- plan and conduct investigations
- communicate scientific information clearly and accurately

Concepts Review

1. Several clues may alert us that a chemical change is taking place. K/U C
   (a) What evidence of chemical change can you observe in **Figure 1(a)**?
   (b) Write a chemical equation for the reaction occurring in Figure 1(a).
   (c) Hydrogen peroxide, H_2O_2, decomposes rapidly when in contact with some vegetables. What is the evidence of a chemical change in **Figure 1(b)**?
   (d) The gas produced in Figure 1(b) causes a glowing splint to relight. Identify the gas.
   (e) Write a balanced chemical equation for the decomposition of hydrogen peroxide, H_2O_2. Assume that the products are the gas from (d) and water.
   (f) What other clues might be seen, in other reactions, indicating that a chemical change is taking place?

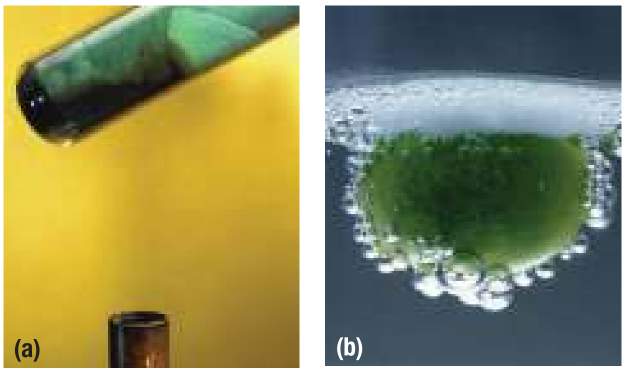


Figure 1 (a) Green copper(II) carbonate decomposes into copper(II) oxide and carbon dioxide when heated. (b) Hydrogen peroxide solution is clear and colourless.

2. **Figure 2** shows models of the combustion of methane, CH_4. K/U T/I C

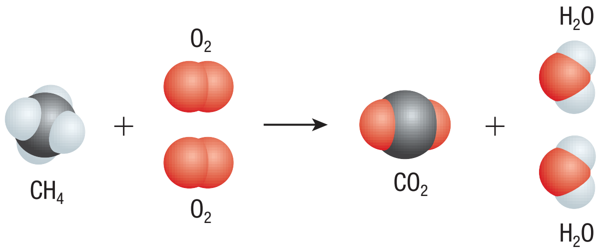


Figure 2 Oxygen atoms are represented by red spheres, carbon by black, and hydrogen by white.

   (a) Write a balanced chemical equation for this reaction.
   (b) Distinguish between coefficients and subscripts in the chemical equation.
   (c) Define the law of conservation of mass.
   (d) Explain why the balanced equation for the combustion of methane obeys the law of conservation of mass.
   (e) When 16 g of methane reacts with 64 g of oxygen, 44 g of carbon dioxide is produced. Predict what mass of water is produced. Explain your prediction.
3. A shiny, silver-coloured element X is a good conductor of electricity. Reacting X with chlorine produces a white solid. K/U
   (a) Classify X as a metal, non-metal, or metalloid.
   (b) If X acquires an ionic charge of +2 when it reacts, what is the chemical formula of the chloride compound that it forms?
   (c) If X has an atomic mass of less than 40 u, to which group of the periodic table does it belong?

Skills Review

4. Write the chemical formula for each of the following compounds. Use the periodic table on the inside back cover of the textbook and the table of ion charges in Section 2.5. K/U C
   (a) potassium sulfide
   (b) ammonium chlorate
   (c) sodium sulfite
   (d) iron(III) nitrite
   (e) calcium phosphate
   (f) carbon disulfide
   (g) potassium dihydrogen phosphate
   (h) barium hypochlorite
   (i) manganese sulfate
   (j) ammonium perchlorate
5. Name the following compounds: K/U
   (a) $AlCl_3$
   (b) $Zn(ClO_3)_2$
   (c) PbO_2
   (d) $(NH_4)_2CO_3$
   (e) Na_3PO_3
   (f) $Ca(HCO_3)_2$
   (g) $Zn(ClO_2)_2$
   (h) SO_2
   (i) N_2O
   (j) $Fe(OH)_3$
6. Balance these chemical equations: K/U C
   (a) $K + O_2 \rightarrow K_2O$
   (b) $P_4 + Cl_2 \rightarrow PCl_3$
   (c) $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$
   (d) $Fe + H_2SO_4 \rightarrow Fe_2(SO_4)_3 + H_2$
   (e) $HBr + Mg(OH)_2 \rightarrow MgBr_2 + H_2O$
   (f) $Na_3PO_4 + CaF_2 \rightarrow NaF + Ca_3(PO_4)_2$
   (g) $(NH_4)_3PO_4 + Sn(NO_3)_4 \rightarrow Sn_3(PO_4)_4 + NH_4NO_3$
7. What safety precautions should be taken in each of the following situations? K/U T/I
   (a) Chemicals or glassware are used in an investigation.
   (b) Open flames are used.
   (c) A flask containing a solution is heated on a hot plate.
   (d) A flammable solvent is brought into the laboratory.
   (e) A substance in a test tube is heated over a Bunsen burner.
   (f) You are asked to smell a sample of a chemical.
   (g) A solution is provided with a label bearing the symbol shown in **Figure 3**.

Figure 3

8. **Figure 4** shows two samples of rainwater. Universal indicator has been added to both samples. T/I A
   (a) Which sample is more acidic? How do you know?
   (b) Which of the following chemicals could be added to make these samples less acidic: HCl(aq), $NaHCO_3$, or NaCl?
   (c) What effect would this addition have on the pH?

Figure 4 Two rainwater samples mixed with an acid–base indicator.

9. The cleanup of an acid spill often requires two steps: containment of the spill followed by treatment. K/U
   (a) What properties should the substances have that are used to contain an acid spill?
   (b) Which substance would you recommend to treat an acid spill at home: sodium hydroxide, NaOH(aq) (in drain-opening products), or sodium hydrogen carbonate, $NaHCO_3$(aq) (baking soda)? Why?
10. Compounds can be broken down or decomposed into simpler substances; elements cannot. How does this difference affect the disposal of hazardous substances? K/U

CAREER PATHWAYS **PREVIEW**

Throughout this unit you will see Career Links in the margins. These links mention careers that are relevant to Chemical Reactions. On the Chapter Summary page at the end of each chapter you will find a Career Pathways feature that shows you the educational requirements of the careers. There are also some career-related questions for you to research.

CHAPTER 4

The Effects of Chemical Reactions

KEY CONCEPTS

After completing this chapter you will be able to

- discover that chemical processes can have unexpected effects on humans and the environment, sometimes far from their source
- classify a variety of chemical reactions as synthesis, decomposition, single displacement, or double displacement
- predict whether certain reactions will occur
- write balanced chemical equations for predicted reactions
- explore synthesis, decomposition, single displacement, and double displacement reactions in the laboratory
- collect experimental evidence to test predictions and identify patterns related to whether reactions will occur
- research a potential solution to an environmental problem and make a decision on whether this solution is appropriate

How Do Chemical Reactions Affect Us and the Environment?

Only a few years ago, an environmentally friendly car or "eco car" was typically slow and not much to look at. Not anymore! The latest eco cars are sleek, "green," and attractive. Also, going green no longer means sacrificing performance. The performance of many eco cars is now comparable to that of their conventional cousins. For example, the Tesla Roadster—an electric car—can accelerate from 0 to 100 km/h in less than 4 s and has a range of about 320 km between recharges, according to the manufacturer.

Most vehicles rely on chemical reactions as their energy source. For example, a conventional car burns gasoline in an internal combustion engine. The problem with this technology is that it pollutes the environment and is very inefficient. The electrical energy used to power an electric car comes from reversible chemical reactions inside its batteries. Electric cars are far more efficient at converting stored energy into motion. Hybrid vehicles use both gasoline and electrical motors. Other eco cars generate energy by burning an alternative fuel such as hydrogen or ethanol.

Advances in our understanding of chemical reactions have also produced the high-tech materials used in eco cars. Strong, yet lightweight materials such as carbon fibre are replacing steel as the materials of choice for automotive bodies. This not only improves their performance but also makes them virtually corrosion free. Chemical reactions also provide the means to recycle and reuse these materials once the car is scrapped.

Before deciding that an eco car is for you, consider all the pros and cons. For example, how "clean" is its energy source? The sales pitch of "zero emissions" is not a guarantee that electric cars run without producing any pollution. These vehicles are only as "clean" as their source of electricity. Another consideration is the impact on society of producing the energy source. For example, the alternative fuel ethanol is primarily made from corn. Diverting a large percentage of the annual corn crop to ethanol production affects the cost of food. And, finally, new technology is always expensive. Be prepared to pay more for the eco car. Is the extra cost worth it?

Almost every product and service that we buy has an impact on society and the environment. This chapter will help you to be more aware of the chemical reactions that are involved in these impacts so that you can make better-informed choices.

STARTING POINTS

Answer the following questions using your current knowledge. You will have a chance to revisit these questions later, applying concepts and skills from the chapter.

1. How can the products of chemical reactions be predicted?
2. Use a specific example to show how a chemical reaction can be beneficial in one application and harmful in another.
3. What role does chemistry play in solving environmental problems?
4. What role do chemical reactions play in a sustainable lifestyle?

Mini Investigation

Back and Forth

Skills: Predicting, Performing, Observing, Analyzing, Evaluating, Communicating

SKILLS HANDBOOK A1.1, A1.2

Many chemical changes, such as the rusting of steel on a car, cannot be reversed. However, as you will see in this investigation, some chemical changes can be reversed.

Equipment and Materials: chemical safety goggles; lab apron; test tube; test-tube rack or small beaker; dropper bottles containing dilute solutions of hydrochloric acid, HCl(aq) (0.1 mol/L), sodium hydroxide, NaOH(aq) (0.1 mol/L), and iron(II) sulfate, $FeSO_4$(aq) (0.1 mol/L)

The hydrochloric acid and sodium hydroxide solution used in this activity are irritants. Wash any spills on skin or clothing immediately with plenty of cold water. Report any spills to your teacher.

1. Put on your chemical safety goggles and lab apron.
2. Add a solution of iron(II) sulfate, $FeSO_4$(aq), to a depth of about 2 cm in a test tube.
3. Add the same volume of sodium hydroxide solution, NaOH(aq), to the same test tube. Record your observations.
4. Slowly add a small volume of hydrochloric acid, HCl(aq), to the test tube until the solution is clear again.
5. Predict how you could reverse the effects of Step 4. Test your prediction by trying it. Record your observations.

A. What evidence of chemical change did you observe? T/I

B. Evaluate your prediction in Step 5. T/I

C. Suggest an explanation for your observations in Step 5. T/I

4.1 Introduction to Chemical Reactions

Figure 1 Cooking involves chemical changes. Energy from the torch caramelizes sugar, producing new substances, tastes, and aromas.

Chemistry is the study of substances and the changes these substances undergo. These changes can be either physical or chemical.

During a physical change, the properties of a substance may change, but its chemical identity remains the same. For example, the rapid expansion of compressed butane gas as it is released from its container is an example of a physical change. The volume of the gas increases, but the gas is still butane. Other common physical changes are changes of state and dissolving.

Chemical changes result in the formation of new substances. For example, when butane reaches the hot nozzle of a blowtorch, it ignites, producing carbon dioxide gas and water vapour (**Figure 1**). It is no longer butane. This chemical change also releases the energy needed to caramelize the sugar in the dessert. The process in which one or more substances undergo a chemical change to produce one or more new substances is called a **chemical reaction**.

chemical reaction a process in which one or more substances change into one or more new substances

precipitate a solid produced as a result of the reaction of two solutions

Evidence of Chemical Reactions

All chemical reactions result in the formation of new substances. How can you tell that a new substance has been produced? Any one of the following clues may indicate that a chemical change has produced one or more new substances:

- There is an unexpected change in colour.
- Energy is released or absorbed.
- A gas is produced.
- A precipitate forms.

Two of these clues apply to the work of the chef in Figure 1. Butane is a colourless gas. However, as it burns, it produces an intense blue flame that radiates a great deal of thermal energy. Both of these observations suggest that the combustion of butane is a chemical reaction. Similarly, the darkening of the surface of the dessert indicates that caramelizing sugar is likely a chemical reaction.

A **precipitate** is a solid produced when two liquids are combined (**Figure 2**).

None of these four clues are conclusive proof that a chemical reaction has taken place. For example, heating water until it boils produces a gas, but no new substance is produced, so this is not a chemical change. The only way to be certain that a chemical change has occurred is to test the product to show that it is a new substance.

Figure 2 Adding yellow potassium chromate solution to colourless silver nitrate solution produces solid red silver chromate.

Atoms in Chemical Reactions

We can use an example to explore what happens at the molecular level during a chemical reaction. Hydrogen peroxide, H_2O_2, decomposes quickly in the presence of manganese(IV) oxide to produce water and a colourless gas (**Figure 3(a)**).

A test can help to identify the gas that bubbles out of the liquid. A glowing splint inserted in the gas relights, suggesting that it is oxygen.

Manganese(IV) oxide is not used up or produced in the reaction; it is neither a reactant nor a product. Instead, it is a catalyst. A **catalyst** is a substance that makes a reaction occur faster without being used up in the reaction.

For every two molecules of hydrogen peroxide that react, one molecule of oxygen and two molecules of water are produced (**Figure 3(b)**). For this to occur, the bonds within the hydrogen peroxide molecules must first break, allowing the atoms to rearrange and form new bonds. All the atoms that were present at the start of the reaction must also be present at the end. Therefore, the total mass of the reactant (hydrogen peroxide) must equal the total mass of the products (water and oxygen). This is true for all chemical reactions and is called the **law of conservation of mass**.

catalyst a substance that makes a chemical reaction occur faster without itself being consumed in the reaction

law of conservation of mass the statement that, during a chemical reaction, the total mass of reactants equals the total mass of products

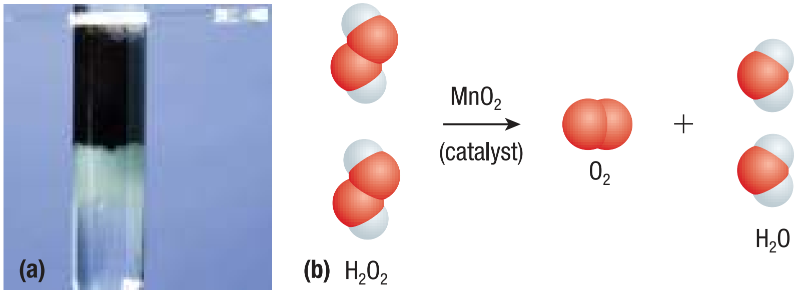


Figure 3 (a) Hydrogen peroxide breaks down in the presence of black manganese dioxide. The products are water and oxygen. (b) All the atoms in the reactants are accounted for in the products. Notice the subscripts, indicating the numbers of atoms in each molecule.

Mini Investigation

Elephant Toothpaste

Skills: Controlling Variables, Performing, Observing, Analyzing, Evaluating, Communicating

SKILLS HANDBOOK **A1, A2.2, A3.8**

Drugstores sell 3 % hydrogen peroxide as a disinfectant for cuts. Beauty supply stores sell 6 % and 12 % hydrogen peroxide for use with hair-colouring products. In this activity, you will explore the effect of the concentration of a hydrogen peroxide solution on its reactivity.

Equipment and Materials: chemical safety goggles; lab apron; 2 large narrow test tubes; 2 stirring rods; small beaker; scoopula; candle; 3 % hydrogen peroxide solution; 6 % hydrogen peroxide solution; masking tape; marker; liquid dish detergent; dry yeast; wooden splint

1. Put on your chemical safety goggles and lab apron.
2. Pour 3 % hydrogen peroxide into one test tube to a depth of about 3 cm and the same volume of 6 % hydrogen peroxide to the other test tube. Label each test tube appropriately.
3. Add about 5 drops of liquid dish detergent to each test tube.
4. Use a different stirring rod to mix the contents of each test tube.
5. Place the test tubes in the beaker.
6. Place the beaker and its contents in the sink to catch any spills.
7. Add enough yeast to cover the tip of the scoopula to each test tube.
8. Light your splint from the lit candle. Test the gas produced with a glowing splint.

An open flame is used. Tie back long hair and secure loose clothing. Never leave the flame unattended.

9. Follow your teacher's instructions for proper disposal.

A. What evidence of chemical reactions did you observe? T/I

B. What variables were controlled in this investigation? What variables were changed? T/I

C. Compare the reactivities of the two hydrogen peroxide solutions. T/I

D. Why do you think hair salons use hydrogen peroxide solutions that are more concentrated than 3 %? What precautions should the technicians in salons follow when using these products? T/I A

Describing Chemical Reactions

Chemists use both word equations and chemical equations to describe chemical reactions. Both types of equations list reactants on the left of an arrow and products on the right. A word equation gives the names of the reactants and products. A chemical equation, however, provides far more detail: it gives the chemical formulas of the reactants and products, their state (**Table 1**), and specific conditions required for the reaction to occur. The chemical equation also gives the ratio in which the chemicals react. This is done through coefficients placed before each chemical formula. A coefficient of "1" is implied if no coefficient is written. For example, the word and chemical equations for the decomposition of hydrogen peroxide are

$$\text{Word equation: hydrogen peroxide} \xrightarrow[\text{(catalyst)}]{MnO_2} \text{water} + \text{oxygen} + \text{energy}$$

$$\text{Chemical equation: } 2\ H_2O_2(aq) \xrightarrow[\text{(catalyst)}]{MnO_2} 2\ H_2O(l) + O_2(g) + \text{energy}$$

This chemical equation is "balanced" because the total number of atoms of each type in the reactants is the same as in the products. In particular, there are four atoms of hydrogen and four atoms of oxygen on both sides of the arrow.

Table 1 State Symbols

Symbol	Meaning
(s)	solid
(l)	liquid
(g)	gas
(aq)	aqueous (dissolved in water)

LEARNING TIP

Balanced Chemical Equations

A chemical equation is balanced if the total number of atoms of each type is the same on both sides of the equation.

Tutorial 1 Balancing Chemical Equations

Follow these useful strategies when balancing chemical equations. For simplicity, the state symbols have been omitted from these examples.

- Start by writing a skeleton equation. This is an equation that includes only the chemical formulas, without coefficients.
- Balance atoms that appear only once on each side of the equation by placing appropriate coefficients in front of the chemical formulas.
- Leave atoms that appear more than once to the end.
- Treat polyatomic ions as one unit, rather than as individual atoms, providing that they do not change during the reaction.
- Check that the final equation is balanced.

Sample Problem 1: Balancing a Chemical Equation

Figure 4 Sodium reacts vigorously in pure oxygen to form sodium oxide, Na_2O.

Write the balanced chemical equation for the reaction of sodium in oxygen to produce sodium oxide (**Figure 4**).

Step 1. Write a skeleton chemical equation for the reaction.

$Na + O_2 \rightarrow Na_2O$

Step 2. Count the number of atoms of each type on either side of the arrow.

$Na + O_2 \rightarrow Na_2O$

1 Na atom   2 Na atoms

2 O atoms   1 O atom

Step 3. Multiply the formulas by an appropriate coefficient until all the atoms are balanced. Keep checking whether the numbers of each type of atom on both sides are balanced.

Na_2O (on the right) must be multiplied by 2 to balance the two oxygen atoms on the left side. Na (on the left) must be multiplied by 4 to balance the four Na atoms on the right side.

$4\ Na + O_2 \rightarrow 2\ Na_2O$

4 Na atoms   4 Na atoms

2 O atoms   2 O atom

Step 4. Write the final chemical equation.

$4\ Na + O_2 \rightarrow 2\ Na_2O$

Sample Problem 2: Balancing an Equation with Polyatomic Ions

Write the balanced chemical equation for the reaction of copper in silver nitrate to produce copper(II) nitrate and silver.

Step 1. Write a skeleton chemical equation for the reaction.

$Cu + AgNO_3 \rightarrow Cu(NO_3)_2 + Ag$

Step 2. Count the number of atoms and polyatomic ions of each type on either side of the arrow.

$Cu + AgNO_3 \rightarrow Cu(NO_3)_2 + Ag$

1 Cu   1 Cu^{2+}

1 Ag^{+} 1 Ag

1 NO_3^{-} 2 NO_3^{-}

Step 3. Multiply the formulas by an appropriate coefficient until all the atoms and ions are balanced. Keep checking whether the numbers of each type of atom on both sides are balanced.

Since NO_3^- remains intact, it can be counted as a unit. $AgNO_3$ must be multiplied by 2 to balance the nitrate ions. Therefore, silver must also be multiplied by 2.

$Cu + 2\ AgNO_3 \rightarrow Cu(NO_3)_2 + 2\ Ag$

1 Cu	1 Cu^{2+}
2 Ag^+	2 Ag
2 NO_3^-	2 NO_3^-

Step 4. Write the final chemical equation.

$Cu + 2\ AgNO_3 \rightarrow Cu(NO_3)_2 + 2\ Ag$

Practice

1. Balance the following chemical equations: K/U C
   (a) $P + O_2 \rightarrow P_2O_5$
   (b) $K_2O + H_2O \rightarrow KOH$
   (c) $AlBr_3 + K_2SO_4 \rightarrow KBr + Al_2(SO_4)_3$
   (d) $FeCl_3 + NaOH \rightarrow Fe(OH)_3 + NaCl$
   (e) $AgNO_3 + H_2S \rightarrow Ag_2S + HNO_3$
   (f) $(NH_4)_2CO_3 \rightarrow NH_3 + H_2O + CO_2$

UNIT TASK BOOKMARK

You will need to balance chemical equations in the Unit Task, described on page 242.

4.1 Summary

- Evidence of a chemical reaction includes colour change; absorption or release of energy; production of a gas (except evaporating or boiling of a liquid); and formation of a precipitate.
- During a chemical reaction, reactant atoms rearrange to form products.
- Chemical reactions are described using word equations or chemical equations.
- A balanced chemical equation gives the correct proportions of chemicals in a chemical reaction. As a result, it obeys the law of conservation of mass.

4.1 Questions

1. (a) What evidence in **Figure 5(a)** suggests that a chemical reaction has occurred?
   (b) What evidence in **Figure 5(b)** suggests that an invisible gas is one of the reaction products? K/U T/I

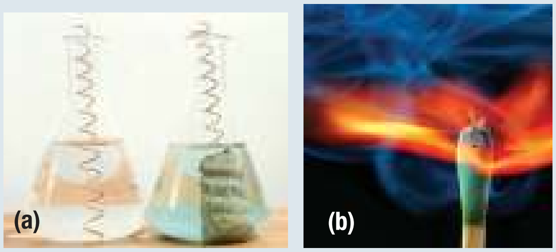


Figure 5 (a) Copper metal reacts in a silver nitrate solution to produce impure silver and dissolved copper nitrate. Dissolved copper(II) ions are blue. (b) A match flares as it catches fire.

2. Consider the statement, “In a chemical reaction, atoms are neither created nor destroyed, only rearranged.” Why is this a statement of the law of conservation of mass? K/U

3. Nitrogen dioxide, NO_2, is an air pollutant formed from automobile exhaust. This toxic gas can be prepared in the laboratory by adding copper metal to a solution of nitric acid, $HNO_3(aq)$. The other reaction products are water and a solution of copper(II) nitrate. K/U T/I C
   (a) Write a word equation for this reaction.
   (b) Write a balanced chemical equation, including all state symbols.
   (c) What is the subscript for oxygen in nitric acid?
   (d) What is the coefficient of nitric acid in the equation?
   (e) How many atoms of oxygen appear on the left-hand side of the equation?
   (f) Distinguish between the symbols (l) and (aq) used in this equation.

4. Why are the coefficients, but not the subscripts, sometimes changed when balancing a chemical equation? K/U

5. Balance the following chemical equations: K/U C
   (a) $S_8 + O_2 \rightarrow SO_2$
   (b) $N_2 + H_2 \rightarrow NH_3$
   (c) $Na + H_2O \rightarrow NaOH + H_2$
   (d) $Li + AlCl_3 \rightarrow LiCl + Al$
   (e) $C_4H_{10} + O_2 \rightarrow CO_2 + H_2O$
   (f) $N_2 + O_2 \rightarrow N_2O_5$
   (g) $Li + B_2O_3 \rightarrow Li_2O + B$
   (h) $Fe_2O_3 + H_2SO_4 \rightarrow Fe_2(SO_4)_3 + H_2O$
   (i) $H_3PO_4 + Ca(OH)_2 \rightarrow Ca_3(PO_4)_2 + H_2O$
   (j) $NH_3 + O_2 \rightarrow N_2 + H_2O$
   (k) $Ca_3(PO_4)_2 + SiO_2 + C \rightarrow CaSiO_3 + CO + P$
   (l) $C_6H_6 + O_2 \rightarrow CO_2 + H_2O$

4.2 Synthesis and Decomposition Reactions

In the previous unit, you learned that the periodic table is a useful tool to predict the properties of elements. Since the properties of elements are predictable, many of the chemical reactions that these elements undergo are predictable as well. For example, **Figure 1** shows the reaction of two metals—sodium and potassium—with chlorine. Because sodium and potassium are both alkali metals, they undergo similar chemical reactions to form chloride compounds. The chemical equations for these reactions are also similar: two elements react to form one compound.

LEARNING **TIP**

Diatomic Elements

Several non-metallic elements occur naturally as diatomic molecules. These include hydrogen and the elements of the fluorine "corner" of the periodic table:

H_2	N_2	O_2	F_2
			Cl_2
			Br_2
			I_2

(a) $2\ Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$ **(b)** $2\ K(s) + Cl_2(g) \rightarrow 2\ KCl(s)$

Figure 1 Alkali metals are highly reactive elements. They react with chlorine to form very stable compounds.

We can use this pattern to predict the products of other reactions involving a metal and chlorine. Magnesium, for example, also reacts with chlorine. The product of this reaction is the compound magnesium chloride. Magnesium has an ionic charge of +2, so the chemical formula of this compound is $MgCl_2$. Therefore, the balanced chemical equation for this reaction is

$$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$$

There are millions of known reactions. Chemists group them into patterns to organize them. Grouping reactions also makes the prediction of reaction products much simpler. In the next few sections, you will learn to recognize four patterns that chemists use to classify chemical reactions. Note, however, that there are many ways in which reactions can be classified. Using these patterns is just one of the ways.

Synthesis Reactions

synthesis reaction a reaction in which two reactants combine to make a larger or more complex product; general pattern: $A + B \rightarrow AB$

The reactions in Figure 1 are examples of synthesis reactions. In a **synthesis reaction**, two reactants combine to form one larger or more complex product. For this to occur, the reactants must first collide, break existing bonds between their atoms, and form new bonds. The chemical equations for synthesis reactions fit the general pattern:

$A + B \rightarrow AB$ (**Figure 2**)

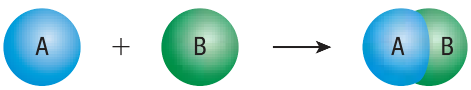


Figure 2 A model of a typical synthesis reaction

Can we use this pattern to predict the products of a synthesis reaction?

Tutorial 1 Predicting the Products of Synthesis Reactions

We can use the example of burning magnesium in a sparkler (**Figure 3**) to illustrate how to predict the product of a synthesis reaction. The reactants in this case are magnesium and oxygen. Oxygen is necessary for most combustion reactions to occur.

Figure 3 The shimmering flashes of a sparkler come from burning grains of metals like magnesium.

Sample Problem 1: Writing Chemical Equations for Synthesis Reactions

Write the balanced chemical equation for the reaction of magnesium with oxygen.

Step 1. Identify the reactants.

$Mg \qquad O_2$

Step 2. Identify the type of reaction.

These reactants fit the pattern of a synthesis reaction:

$A + B \rightarrow AB$

$Mg + O_2 \rightarrow Mg_xO_y$

Step 3. Use the ionic charges of the reactants to predict the formula of the product. Reduce to simplest ratios, if necessary.

Ionic charges: $Mg^{2+} \quad O^{2-}$

Product: $Mg_{\cancel{2}}O_{\cancel{2}}$

Skeleton equation: $Mg + O_2 \rightarrow MgO$

Step 4. Balance the chemical equation and write the balanced chemical equation, including state symbols.

$2\,Mg(s) + O_2(g) \rightarrow 2\,MgO(s)$

Practice

1. Write the balanced chemical equation for the reaction of these pairs of reactants:
   (a) calcium and bromine
   (b) aluminum and oxygen

All the reactions that we have considered so far involve metals reacting with non-metals to form ionic compounds. Some reactions, however, involve only non-metals.

Synthesis Reactions of Non-Metals

We can apply the pattern of synthesis reactions to this situation. Since the reactants are non-metals, the product will be a molecular compound.

SYNTHESIS REACTIONS INVOLVING HYDROGEN

Hydrogen is unusual for a Group 1 element. Most Group 1 elements form ionic compounds when they bond with other elements. Hydrogen, however, usually forms molecular compounds. Molecular compounds are made up of atoms, not ions. However, ionic charges are still useful in predicting the products of synthesis reactions involving hydrogen (**Figure 4**). For example, the reaction between hydrogen, H_2, and chlorine, Cl_2, follows the familiar pattern of synthesis reactions:

$$A \quad + \quad B \quad \rightarrow \quad AB$$

$$H_2(g) + Cl_2(g) \rightarrow H_xCl_y$$

If we apply the ionic charges of hydrogen (+1) and chlorine (–1), the product of the reaction is hydrogen chloride, HCl. Therefore, the skeleton equation and balanced chemical equation for this reaction are

$$H_2 + Cl_2 \rightarrow HCl \text{ and}$$

$$H_2(g) + Cl_2(g) \rightarrow 2\,HCl(g)$$

Figure 4 Colourless hydrogen gas from the tube reacts with yellow chlorine gas in the cylinder to produce toxic hydrogen chloride gas.

Figure 5 Some Canadian cities use hydrogen-powered buses for public transit.

WEB LINK

To find out more about using hydrogen as a fuel,

GO TO NELSON SCIENCE

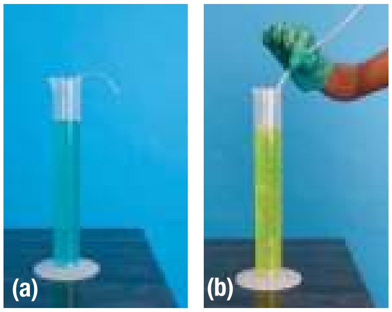


Figure 6 A colour change of bromothymol blue indicator from (a) blue to (b) yellow indicates that an acid is forming in the solution.

Figure 7 Acidification of the oceans threatens coral reefs.

The reaction between hydrogen and oxygen also follows the pattern of synthesis reactions:

$$A \quad + \quad B \quad \rightarrow \quad AB$$

$$H_2(g) + O_2(g) \rightarrow H_xO_y$$

If the ionic charges of hydrogen (+1) and oxygen (−2) are applied, we can see that two hydrogen atoms are needed to balance the ionic charge of one oxygen atom. The product of the reaction is H_2O. Therefore, the balanced chemical equation for this reaction is

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$$

The reaction of hydrogen with oxygen releases a great deal of energy, as well as water. This reaction has the potential to be an important source of "green energy." In fact, a fleet of hydrogen buses was used in Vancouver during the 2010 Winter Olympics (**Figure 5**). However, since hydrogen does not occur naturally as an element, it has to be extracted from compounds that contain hydrogen. This, of course, requires energy. Therefore, hydrogen fuel can truly be considered "green" only if it is generated from environmentally friendly raw materials using a renewable energy source. Water would be a suitable raw material. Solar power would be a suitable energy source.

SYNTHESIS REACTIONS NOT INVOLVING HYDROGEN

The products of synthesis reactions of non-metals other than hydrogen are difficult to predict. The products of these reactions often depend on the reaction conditions. For example, two products are possible for the reaction of carbon with oxygen, depending on the availability of oxygen:

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

$$2\ C(s) + O_2(g) \rightarrow 2\ CO(g)$$

In these cases, the only way to identify the reaction products is by conducting chemical tests.

Synthesis Reactions Involving Compounds

So far, we have only considered synthesis reactions involving elements as reactants. However, compounds also can participate in synthesis reactions. For example, when carbon dioxide is bubbled into water the two compounds react to form a new substance. The new substance is an acid called carbonic acid.

$$H_2O(l) + CO_2(g) \rightarrow H_2CO_3(aq)$$

We can detect the production of carbonic acid (or any acid) by adding an acid–base indicator to the mixture. Bromothymol blue is an example of an acid–base indicator. It is blue in a basic solution and yellow in an acidic solution. When carbon dioxide reacts with water containing bromothymol blue, the mixture turns from blue to yellow, showing that an acid has formed (**Figure 6**).

Air contains a small percentage of carbon dioxide. Whenever water is in contact with air, some of the carbon dioxide dissolves in the water and reacts to produce carbonic acid. The synthesis of carbonic acid makes "normal" rain slightly acidic. The same process also acidifies seawater. The recent increase in carbon dioxide in the atmosphere has produced a significant increase in the acidity of the world's oceans. This increased acidity could have devastating effects (**Figure 7**). Coral is made mostly of calcium carbonate. This calcium carbonate will react with the acid in the ocean to produce soluble compounds. Coral reefs will gradually dissolve and crumble.

Sulfur trioxide is another compound that is affecting the environment because of a synthesis reaction. It is produced during the combustion of fossil fuels. Water in the environment reacts with sulfur trioxide in a synthesis reaction to form sulfuric acid:

$$H_2O(l) + SO_3(g) \rightarrow H_2SO_4(aq)$$

The product of each of the previous two examples could be predicted simply by combining the two reactant chemical formulas. Unfortunately, we cannot predict the products of all synthesis reactions so simply. The products that actually form depend on the reaction conditions.

CAREER LINK

Environmental engineers are concerned with reducing pollution. To find out how you might join this career,

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Decomposition Reactions

Special effects are an important part of any action movie. Many effects are computer generated, but others are the result of chemical reactions. The reactants are explosives—placed in just the right spot to produce the desired effect—and oxygen (**Figure 8**). The science of handling explosives is called pyrotechnics and, as you would expect, requires highly specialized training. Explosives experts do not only work on Hollywood movie sets. Explosives are also used for clearing rock for road construction and mining.

CAREER LINK

If you like the idea of working safely with explosions, you might be interested in becoming a pyrotechnician. To investigate further,

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Figure 8 Carefully placed explosives produce spectacular movie effects such as vehicle rollovers. An explosion is a rapid chemical reaction.

Many explosives are molecular compounds of four non-metal elements: carbon, hydrogen, nitrogen, and oxygen. TNT, for example, is more correctly called trinitrotoluene, $C_7H_5(NO_2)_3$. It is a common industrial explosive. The detonation of this compound releases a great deal of energy and several reaction products:

$$2\ C_7H_5(NO_2)_3(s) \rightarrow 7\ C(s) + 7\ CO(g) + 3\ N_2(g) + 5\ H_2O(g) + \text{energy}$$

A chemical reaction in which one large compound breaks down or decomposes into two or more smaller substances is called a **decomposition reaction**. The general pattern for a decomposition reaction is

$AB \rightarrow A + B$ (**Figure 9**)

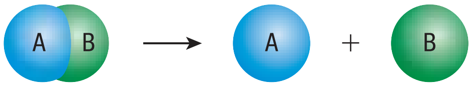

Figure 9 Decomposition reactions produce two or more smaller substances.

decomposition reaction a reaction in which a large or more complex compound breaks down to form two (or more) simpler products; general pattern: $AB \rightarrow A + B$

LEARNING TIP

Opposite Reactions
Note that the patterns of decomposition reactions and synthesis reactions are opposites of each other.

Decomposition reactions usually need energy to get started—even for reactions that release a lot of energy. In the case of TNT, a small electric current starts the reaction. It is not only the release of energy that makes TNT so dangerous, however. The source of TNT's destructive power becomes clear when you consider the states of its decomposition products. Most of these substances are gases. Gases occupy far more space than solids and expand very quickly when they are heated. The thermal energy released by exploding TNT makes these gases expand very quickly, creating a powerful destructive force a fraction of a second after detonation.

Predicting the Products of Decomposition Reactions

Simple ionic compounds such as potassium chloride, KCl, can be made to decompose into their elements. For example, potassium metal and chlorine gas are made industrially by passing electricity through molten potassium chloride:

$$2\ \mathrm{KCl(l)} \xrightarrow{\text{electricity}} 2\ \mathrm{K(s)} + \mathrm{Cl_2(g)}$$

However, predicting the decomposition products of most molecular compounds or ionic compounds involving polyatomic ions is difficult. For example, you might expect potassium chlorate, $KClO_3$, to decompose into its elements when heated (**Figure 10**). Instead, the decomposition products are potassium chloride and a colourless gas. A glowing splint relights when placed in the gas, identifying it as oxygen. The chemical equation for the decomposition of potassium chlorate is therefore

$$2\ \mathrm{KClO_3(s)} \xrightarrow{\text{heat}} 2\ \mathrm{KCl(s)} + 3\ \mathrm{O_2(g)}$$

Some decomposition reactions produce only compounds. For example, the intense heat of a Bunsen burner flame causes calcium carbonate, $CaCO_3$, to decompose into calcium oxide, CaO, and an invisible gas. Bubbling this gas into limewater produces a precipitate, indicating that the gas is carbon dioxide. (Recall that the limewater test is a common test for carbon dioxide gas.) Therefore, the chemical equation for the decomposition of calcium carbonate is

$$\mathrm{CaCO_3(s)} \xrightarrow{\text{heat}} \mathrm{CaO(s)} + \mathrm{CO_2(g)}$$

Calcium oxide, also known as lime, is a key ingredient in cement. Industrially, kilns operating above 1400 °C are used to decompose limestone ($CaCO_3$) to produce calcium oxide (**Figure 11**). Kilns consume huge quantities of fossil fuels to maintain this temperature. Fossil fuel combustion and the decomposition of calcium carbonate both generate carbon dioxide—a greenhouse gas. Not surprisingly, cement kilns are a significant source of greenhouse gas emissions.

Figure 10 The head of a match contains potassium chlorate, among other chemicals. Oxygen, a product of the decomposition reaction, makes the match burn faster.

Figure 11 Cement kilns must operate at high temperatures to decompose limestone (calcium carbonate).

4.2 Summary

- Chemical reactions can be grouped according to the pattern of reactants and products in their chemical equations.
- In a synthesis reaction, two simple reactants combine to make a larger or more complex compound and follow the general pattern

  $A + B \rightarrow AB$
- In a decomposition reaction, a complex compound breaks down into two or more simpler products and follows the general pattern

  $AB \rightarrow A + B$

4.2 Questions

1. Classify these reactions as synthesis or decomposition. Justify your choice. K/U
   (a) $2\ Al + 3\ Br_2 \rightarrow 2\ AlBr_3$
   (b) $2\ HCl \rightarrow H_2 + Cl_2$
   (c) $CaO + H_2O \rightarrow Ca(OH)_2$
   (d) $P_4 + 5\ O_2 \rightarrow 2\ P_2O_5$
2. Predict the products of these synthesis or decomposition reactions. Write a balanced chemical equation to represent each reaction. K/U C
   (a) $Z + S \rightarrow$
   (b) $CaCl_2 \rightarrow$
   (c) $NH_3 + HCl \rightarrow$
   (d) $K_2O \rightarrow$
   (e) $AlCl_3 \rightarrow$
   (f) $Mg(OH)_2 \rightarrow H_2O +$
3. Hydrogen peroxide forms gas bubbles when it is added to blood. The other reaction product is water. Inserting a glowing splint into a sample of this gas causes the splint to relight. K/U T/I
   (a) Identify the gas.
   (b) Classify the reaction.
   (c) Write a balanced chemical equation for this reaction.
4. Write a balanced chemical equation for each of these reactions: K/U C
   (a) Aluminum metal readily reacts in air to form a hard protective coating of aluminum oxide.
   (b) Copper(II) oxide and carbon dioxide are produced when copper(II) carbonate is heated.
   (c) Solid nitrogen triiodide is a shock-sensitive explosive that is stable when wet and explosive when dry. This compound decomposes rapidly to produce a gas when detonated.
5. Photosynthesis is a chemical process that occurs in green plants in which solar energy is converted into stored chemical energy. This process is the basis of life on Earth. The overall chemical equation for photosynthesis is

   $$6\ CO_2(g) + 6\ H_2O(l) \rightarrow C_6H_{12}O_6(s) + 6\ O_2(g)$$

   Compare photosynthesis to the synthesis reactions discussed in this section. K/U
6. Most fossil fuels contain traces of sulfur. When the fuel is burned, the sulfur impurities become a major cause of acid precipitation. Three chemical reactions are involved. Write a balanced chemical equation for each reaction. K/U C
   (a) Solid sulfur reacts with oxygen to form gaseous sulfur dioxide.
   (b) The product from (a) then further reacts with oxygen to form gaseous sulfur trioxide.
   (c) The product from (b) then reacts with droplets of water in the atmosphere to form sulfuric acid.
7. Sodium hydrogen carbonate (baking soda) easily decomposes to release sodium carbonate, water, and a gas. Cake batter rises because of the decomposition of sodium hydrogen carbonate. K/U C
   (a) What gas is produced? [Hint: It turns limewater cloudy.]
   (b) Write a balanced chemical equation for this reaction.
8. A red solid compound is heated in a test tube. The reaction produces a colourless gas that relights a glowing splint. After the test tube cools, a silver-coloured liquid metal element remains. K/U T/I C
   (a) What type of chemical reaction occurred in the test tube? Why?
   (b) Identify the gaseous product.
   (c) Identify the metal.
   (d) Name and write the chemical formula for the red solid compound. (Use the higher ionic charge of the metal).
   (e) Write a balanced chemical equation for this reaction.
9. The decomposition of solid sodium azide, NaN_3, is the chemical reaction that inflates automobile airbags during a collision. K/U T/I C
   (a) Write a chemical equation for this reaction. Assume that the first products formed are both gases because of the high temperature of the reaction.
   (b) Research the risks and benefits associated with airbags.
10. Cement kiln operators sometimes supplement the fuel used in cement kilns with old car tires. Research the potential environmental benefits and threats of this practice. Communicate your findings in a Pros and Cons chart. In your opinion, should the practice be continued or banned? K/U T/I C

GO TO NELSON SCIENCE

4.3 Explore an Issue in Chemical Reactions

SKILLS MENU

- Defining the Issue
- Researching
- Identifying Alternatives
- Analyzing the Issue
- Defending a Decision
- Communicating
- Evaluating

Garbage Gasification—A Heated Debate

Ontario's garbage is mounting. Despite aggressive recycling and composting campaigns, we are producing more garbage than we know what to do with. So far, most garbage that is not recycled is dumped in landfill sites (**Figure 1**). However, existing landfill sites are rapidly filling up and the opposition to opening new sites is intense. We need to find an alternative way of getting rid of waste.

Many municipalities are investigating gasification as a possible solution to their garbage dilemma. Gasification is not simply burning the garbage. Gasification involves heating waste to temperatures high enough to cause their molecules to decompose into simpler substances. This process is already being used to convert toxic industrial and hospital waste into safer substances.

Plasma gasification is a particular gasification technology that involves heating waste in a sealed chamber in which there is almost no oxygen (**Figure 2**). Under these conditions, large carbon-based molecules, such as those in plastics, decompose into "syngas." Syngas is a mixture composed mostly of hydrogen and carbon monoxide. Chemical treatment or "scrubbing" removes contaminants from the syngas.

Cleaned syngas is an energy-rich mixture that burns like a fossil fuel and is used to generate electricity. The amount of electricity that it generates far exceeds the energy requirements of the gasification plant. The excess energy would be available to be sold.

Alternatively, the components in syngas can be used to make other carbon-based industrial chemicals, such as ethanol, C_2H_5OH.

Figure 1 Hundreds of truckloads of Ontario garbage arrive at landfill sites each day. The dilemma of what to do with this mountain of mess remains a topic of hot debate in many municipalities.

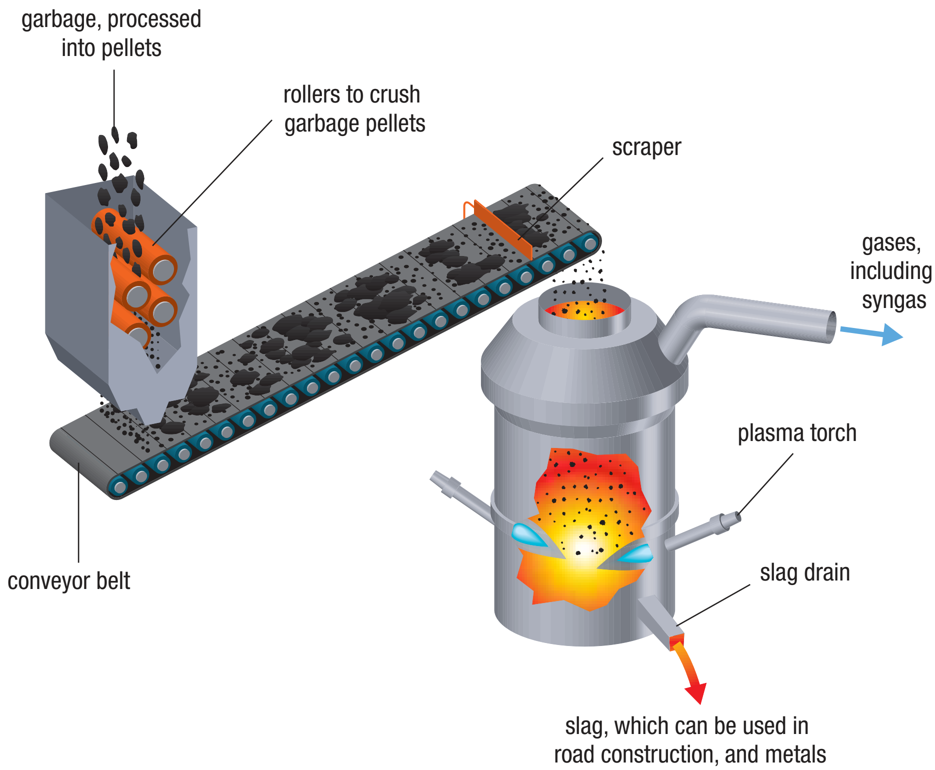

Figure 2 Plasma gasification furnace

Plasma gasification uses high-energy electrical sparks called plasmas to achieve temperatures as high as 10 000 °C—hotter than the surface of the Sun. The energy is so intense that it literally rips molecules apart and forces atoms to ionize. It is somewhat like placing waste into a continuous lightning bolt (**Figure 3**).

Figure 3 Plasma is a high-energy mixture of ionized particles and electrons. In plasma gasification, the energy of the plasma decomposes the molecules in waste into simpler substances.

Critics of gasification are concerned about its emissions. Previous attempts to burn waste produced toxic by-products that escaped into the environment. In particular, dioxin is a highly toxic compound produced when plastics containing chlorine are burned. Critics also argue that the convenience of "zapping away" garbage undermines progress Ontarians have made in recycling. The "buy today and destroy tomorrow" mentality is not sustainable, given that our planet has finite resources. Finally, critics maintain that there are always risks associated with implementing a new technology. Are the potential benefits worth it?

The Issue

SKILLS HANDBOOK A5

The Ontario city of Melton is considering adding a gasification plant to its waste management strategy. Recycling programs have been well received and are successful. However, every year the city still generates about 80 000 t of waste that cannot be composted or recycled. Furthermore, landfill sites are almost full. Attempts to incinerate waste 20 years ago failed miserably. The stench and fumes from that incinerator still linger in citizens' memories.

Your environmental consulting firm has been asked to prepare a report outlining the advantages and disadvantages of including gasification as part of the city's waste management strategy. Your report will be presented to Melton City Council.

Goal

To decide whether a gasification facility should be built in the city

Research

Work in pairs or in small groups to learn more about waste gasification. Research these questions:

- Compare plasma gasification to one other gasification technology.
- What toxic emissions or waste products does gasification generate?
- How successful are gasification technologies at treating their emissions and waste?
- What other disadvantages are associated with gasification?
- What other cities are currently using this technology and how successful has it been?

WEB LINK

To start your research into these questions,

GO TO NELSON SCIENCE

Identify Solutions

Is waste gasification a viable solution to deal with a portion of Melton's garbage? If so, which type of gasification process should be selected? If not, what alternative should be investigated?

Make a Decision

Which solution does your firm recommend? Outline the rationale you used to reach your decision.

Communicate

Complete a presentation of your findings that will be shared with City Council. State your recommendation and your reasons for it. Remember that only a few of the members of your target audience are scientists or engineers. However, these people will make the final decision on whether to proceed with your recommendation.

Plan for Action

Inform yourself about the garbage situation in your region. How does your municipality currently dispose of garbage? Is there a garbage crisis in your region? If so, what possible solutions have been proposed?

GO TO NELSON SCIENCE

4.4 Single Displacement Reactions

Many of us would not want to be without our cellphones. But for an increasing number of Canadians, cellphones are becoming literally a "pain in the ear." These people suffer from nickel contact dermatitis—a skin allergy to nickel (**Figure 1**). The source of the nickel is the shiny metal parts of the cellphone that are in frequent contact with the skin: the buttons and outer rim. This metal contains nickel. Nickel ions gradually leave the cellphone and transfer to the skin. Over time, constant exposure to nickel may make the skin sensitive to it. Once sensitized, the next exposure to nickel can result in an allergic reaction.

Figure 1 (a) and (b) Cellphones are sometimes responsible for serious skin problems. (c) A chemical test shows whether nickel ions are present.

All skin naturally produces secretions, such as sweat, that contain acidic compounds like lactic acid and amino acids. These compounds make sweat mildly acidic. The acidity of skin is sufficient to slowly corrode nickel from the cellphone. Let's consider the reaction of one of these acids, lactic acid, $HC_3H_5O_3(aq)$, with nickel:

$$\underset{\text{nickel}}{Ni(s)} + \underset{\text{lactic acid}}{2\,HC_3H_5O_3(aq)} \rightarrow \underset{\text{nickel lactate}}{Ni(C_3H_5O_3)_2(aq)} + \underset{\text{hydrogen}}{H_2(g)}$$

Elements Changing Places

The reaction of nickel with an acid follows the same pattern as other similar reactions. These reactions are called single displacement reactions. In a **single displacement reaction**, one element displaces or replaces an element in a compound. The general pattern for a single displacement reaction is

$A + BC \rightarrow AC + B$ (**Figure 2**)

single displacement reaction a reaction in which an element displaces another element in a compound, producing a new compound and a new element; general pattern: $A + BC \rightarrow AC + B$

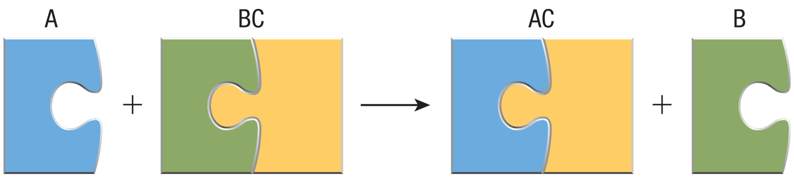

Figure 2 In a single displacement reaction, one element, A, displaces another element, B, from a compound, BC. The products are a new compound, AC, and the displaced element, B.

Single displacement reactions are useful when you need to recover a dissolved metal from solution. For example, copper(II) sulfate is a common chemical used in high school labs. Each year, schools generate a considerable volume of waste copper(II) sulfate. Unfortunately, dissolved copper ions are toxic and must be treated before disposal. One treatment option is to add a piece of reactive metal, such as zinc, to the solution. A single displacement reaction occurs in which zinc displaces copper from the solution (**Figure 3**). The equation for this reaction is

$$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$$

Figure 3 The reaction of zinc with a solution of copper(II) sulfate. The original blue colour of the copper(II) sulfate solution fades as copper ions are displaced out of the solution to form copper metal.

Zinc metal also reacts with solutions of many other metal compounds. In other words, zinc is a reactive metal. Gold, on the other hand, rarely reacts when placed in solutions. This property explains why gold is one of the few metals that are found as elements in nature. Most other metals exist only as compounds. Chemists use the word "activity" to describe the reactivity of a metal. The more reactive the metal, the greater its activity is. Zinc is a more reactive metal than gold.

The Activity Series of Metals

In Investigation 1.5.1, The Search for Patterns, you compared the activity of different metals. Chemists have used the results of investigations like this to develop an **activity series:** a ranking of the reactivity of metals relative to each other (**Figure 4**). The most reactive metals—lithium and potassium—are at the top of the activity series. The least reactive metals are at the bottom. Hydrogen is included in the series, even though it is not a metal, because it forms a positively charged ion like a metal. Including hydrogen also makes it easier to predict which metals react in acids and water. The activity series is based on two important generalizations:

1. One element can displace elements below it from compounds in solution but cannot displace elements above it.
2. The farther apart two elements are, the more likely it is that the displacement reaction will occur quickly.

To determine whether or not a reaction between an element and a compound will proceed, look at the relative positions of the two metals in the activity series. If the higher (more active) metal is the element, the reaction proceeds. If the higher metal is in the compound, no reaction occurs.

activity series a list of elements arranged in order of their observed reactivity in single displacement reactions

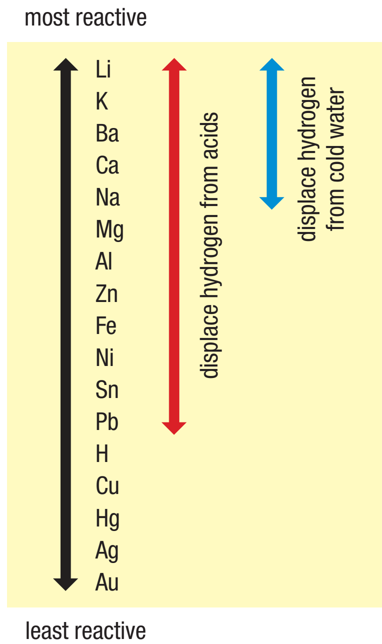

Figure 4 In the activity series of metals, the metal elements are arranged from most reactive to least reactive. Any metal will displace any other metal below it from a compound. Hydrogen is included in this series because, like the metals, it forms a positive ion.

Tutorial 1 Predicting Single Displacement Reactions

The following Sample Problems show how to use the activity series to predict whether a displacement reaction will occur between certain metals and solutions. Assume that the products of these reactions will be a dissolved metal compound and a solid metal.

Sample Problem 1: Reactions Involving a Metal and an Ionic Compound

Using the activity series, write a balanced chemical equation for the reaction that occurs when magnesium metal is placed in a solution of copper(II) sulfate. If you predict that no reaction occurs, write ***no reaction***.

Step 1. Write the chemical formulas of the reactants and identify the element and the compound.

Mg(s) $CuSO_4$(aq)
element   compound

$$A + BC \rightarrow AC + B$$

Step 2. Identify which metal is higher on the activity series.

The metal element, Mg, is higher than the metal in the compound, Cu. Therefore, a displacement reaction will occur.

Step 3. Determine the chemical formula of the new compound AC.

Mg^{2+} SO_4^{2-}
element   compound

$MgSO_4$

Step 4. Write the balanced chemical equation for the reaction.

$$Mg(s) + CuSO_4(aq) \rightarrow MgSO_4(aq) + Cu(s)$$

Sample Problem 2: Reactions Involving a Metal and an Ionic Compound

Using the activity series, write a balanced chemical equation for the reaction that occurs when lead is placed in a zinc nitrate solution. If you predict that no reaction occurs, write ***no reaction***.

WEB LINK

To investigate the activity series of metals using an online simulation,

GO TO NELSON SCIENCE

Step 1. Write the chemical formulas of the reactants and identify the element and the compound.

$Pb(s)$ $\quad$ $Zn(NO_3)_2(aq)$
element $\quad$ compound

$A \quad + \quad BC \rightarrow AC + B$

Step 2. Identify which metal is higher on the activity series. If the higher metal is the element, the reaction proceeds. If the higher metal is in the compound, no reaction occurs.

The metal in the compound, Zn, is higher than the metal element, Pb. Therefore, no displacement reaction occurs.

No reaction

Practice

1. Use the activity series to write a chemical equation for the following reactions that actually occur. Write ***no reaction*** if you predict that no reaction occurs. K/U C
   (a) $Mg(s) + AgNO_3(aq) \rightarrow$
   (b) $Zn(s) + FeCl_2(aq) \rightarrow$
   (c) $Ni(s) + Al(NO_3)_3 \rightarrow$

REACTIONS INVOLVING METALS AND WATER OR ACIDS

Not all single displacement reactions involve a metal in an aqueous solution of an ionic compound. Sometimes the reactants are a metal and a compound that contains hydrogen. Among the most familiar hydrogen compounds are water and acids (**Figure 5**). The same procedure outlined above can be used to predict displacement reactions involving acids and water. These tips will help:

Figure 5 Bubbles of hydrogen gas form on the surface of a metal as it reacts with an acid. In this case, magnesium reacts in hydrochloric acid:

$Mg(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

1. Treat acids as if they were ionic compounds, e.g.,

   $HCl(aq) = H^+\ Cl^-$

   $H_2SO_4(aq) = H^+\ H^+\ SO_4^{2-}$
2. Treat water as if it were an ionic compound:

   $H_2O = H^+\ OH^-$

Sample Problem 3: Reactions Involving a Metal and an Acid

Use the activity series to write a balanced chemical equation for the reaction that occurs when lithium is added to sulfuric acid. If you predict that no reaction occurs, write ***no reaction***.

Step 1. Write the chemical formulas of the reactants and, underneath, the atoms and ions involved in the reaction.

$Li(s)$ $\quad$ $H_2SO_4(aq)$

$Li(s)$ $\quad$ $H^+\ H^+\ SO_4^{2-}\ (aq)$

$A \quad + \quad BC \quad \rightarrow AC + B$

Step 2. Identify which is higher on the activity series: the element or hydrogen. If the element is higher, the reaction proceeds. If the hydrogen is higher, no reaction occurs.

Li is higher than H. Therefore, a displacement reaction does occur.

Step 3. Determine the chemical formula of the new compound AC. The remaining element is displaced.

Li^+ $\quad$ SO_4^{2-}

Li_2SO_4

Step 4. Write the balanced chemical equation for the reaction.

$Li(s) + H_2SO_4(aq) \rightarrow Li_2SO_4(aq) + H_2(g)$

Investigation 4.4.1

Hydrogen Blast-Off (p. 178)
The reaction of magnesium and hydrochloric acid is used to produce hydrogen in this investigation.

Sample Problem 4: Reactions Involving a Metal and Water

Use the activity series to write a balanced chemical equation for the reaction that occurs when sodium is added to water. If you predict that no reaction occurs, write ***no reaction***.

Step 1. Write the chemical formulas of the reactants and, underneath, the atoms and ions involved in the reaction.

$Na(s)$ $\quad H_2O(l)$

$Na(s)$ $\quad H^+ \; OH^-(l)$

$A \quad + \quad BC \quad \rightarrow AC + B$

Step 2. Identify which is higher on the activity series: the metal element or hydrogen. If the element is higher, the reaction proceeds. If the hydrogen is higher, no reaction occurs.

Na is higher than H. Therefore, a displacement reaction does occur (**Figure 6**).

Step 3. Determine the chemical formula of the new compound AC. The remaining element is displaced.

$Na^+ \quad OH^-$

NaOH

Step 4. Write the balanced chemical equation for the reaction.

$2\,Na(s) + 2\,H_2O(l) \rightarrow 2\,NaOH(aq) + H_2(g)$

Practice

2. Use the activity series to write a chemical equation for the following reactions that actually occur. Write ***no reaction*** if you predict that no reaction occurs. K/U C
   (a) $Ca(s) + HBr(aq) \rightarrow$
   (b) $Cu(s) + H_2O(l) \rightarrow$
   (c) $Ba(s) + H_2O(l) \rightarrow$
   (d) $Hg(l) + H_2SO_4(aq) \rightarrow$

Figure 6 The reaction of sodium metal (small ball in the centre) with water containing phenolphthalein indicator. The magenta colour of the indicator shows that hydroxide ions are being produced.

Figure 7 Chlorine gas is bubbled through the glass tube into a colourless solution of potassium iodide. The dark orange colour in the test tube is the characteristic colour of iodine, $I_2(s)$.

The Halogen Activity Series

The elements of the halogen family also have their own activity series based on displacement reactions. For example, **Figure 7** shows the reaction that occurs when chlorine gas is bubbled into a colourless solution of potassium iodide. The reddish colour indicates that elemental iodine was produced in the solution. The chemical equation for this reaction is

$$Cl_2(g) + 2\,KI(aq) \rightarrow 2\,KCl(aq) + I_2(aq)$$
$$A \quad + \quad BC \quad \rightarrow \quad BA \quad + \quad C$$

Note that the single displacement pattern for halogens is slightly different from the pattern used for metals. This is because a negative ion (anion) is being displaced rather than a positive ion (cation).

Bromine also reacts with potassium iodide solution. In this case, the balanced chemical equation is

$$Br_2(l) + 2\,KI(aq) \rightarrow 2\,KBr(aq) + I_2(aq)$$

Because of this evidence, you would expect iodine to be below both bromine and chlorine on a halogen activity series. You would therefore not expect a reaction when iodine is added to a potassium chloride solution.

$$I_2(s) + 2\,KCl(aq) \rightarrow \text{no reaction}$$

The evidence from similar investigations is the basis for the activity series of halogens (**Figure 8**). Just like the activity series of metals, it helps us to predict which elements will displace others from compounds and which will not. Fluorine, for example, displaces all other halogens from compounds.

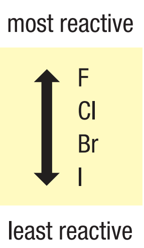

Figure 8 The halogen activity series. Notice that the trend in halogen reactivity is the same as the trend in electronegativity of these elements.

CAREER LINK

A metallurgical control analyst may work both outdoors and in the lab. If you would like to find out more about this career,

GO TO NELSON SCIENCE

Single Displacement Reactions of Solids

All the examples we have discussed so far involve aqueous solutions. Some industrial single displacement reactions involve only solid reactants. For example, making iron for steel involves reacting iron ore (mostly Fe_2O_3) with coke (carbon) in a blast furnace at 1000 °C. This is hot enough to melt iron. The chemical equation for this reaction is

$$\begin{array}{ccccccccc} 3\ C(s) & + & Fe_2O_3(s) & \rightarrow & 3\ CO(g) & + & 2\ Fe(l) \\ A & + & BC & \rightarrow & AC & + & B \end{array}$$

Mini Investigation

Ironing Out Rust

Skills: Performing, Observing, Analyzing, Communicating

SKILLS HANDBOOK A1.2, A3

Iron ore is made up of several compounds, including iron(III) oxide. In this activity you will recreate, on a small scale, the chemical reaction that converts iron(III) oxide to iron (**Figure 9**). Instead of a blast furnace, you will use a Bunsen burner.

Figure 9 A steel foundry heats coke (carbon) and other ingredients with iron ore (mostly Fe_2O_3) to convert it into molten iron.

Equipment and Materials: chemical safety goggles; lab apron; 2 Petri dishes; magnet; small, sealable plastic bag; heat-resistant test tube fitted with a one-hole stopper and angled glass tubing; scoopula; Bunsen burner clamped to a retort stand; spark lighter; 2 utility clamps; test tube; piece of paper, 6 cm × 6 cm; test-tube rack; vials of iron(III) oxide and powdered carbon; limewater

This investigation involves an open flame. Tie back long hair and secure loose clothing. Never leave the flame unattended.

The reaction of iron(III) oxide and carbon releases a lot of thermal energy. The test tube will become very hot. Handle it with caution.

1. Put on your chemical safety goggles and lab apron.
2. Place a small quantity (about one scoopula full) of iron(III) oxide powder in the Petri dish.
3. Seal the magnet in the plastic bag. Hold the magnet underneath the dish and test the powder for magnetism.
4. Fold the paper square in half, open it out, and place it on the bench. Tip the iron(III) oxide powder onto the paper and add a scoopula full of powdered carbon. Stir to mix.
5. Carefully tip the mixture into the heat-resistant test tube. Fit the stopper and tubing into the top.
6. Pour limewater into the second test tube until it is about one-third full.
7. Light a Bunsen burner and adjust the flame until it is blue.
8. Hold each of the test tubes with a utility clamp. Submerge the free end of the plastic tubing in the limewater.
9. Gently heat the iron(III) oxide powder in the flame for about 10 min. Record your observations of both test tubes.
10. When the reaction is complete, remove the tubing from the limewater. Make sure that there is no liquid in the tubing. Turn off the Bunsen burner. Place both test tubes in the test-tube rack.
11. Once the contents of the heated test tube have cooled, transfer them to the second Petri dish.
12. Test the contents of the dish from underneath with the magnet.
13. Dispose of the materials as directed by your teacher.

A. What evidence suggests that pure iron was produced in this reaction? T/I

B. Based on this investigation, where would carbon fit on an activity series, relative to iron? Why? K/U T/I

Investigation 4.4.2

An Activity Series of Ions (p. 180)
This investigation establishes an activity series for metal ions and compares it with the activity series of metals.

4.4 Summary

- In a single displacement reaction, an element displaces another element in a compound. The result is a new compound and a new element. General pattern: A + BC → AC + B
- In an activity series, the elements range from most reactive at the top to least reactive at the bottom.
- One element will displace an element lower on the series from a compound.
- Hydrogen is included in the metal activity series because it forms a positive ion.
- Hydrogen forms $H_2(g)$ when it is displaced from an acid or water.

4.4 Questions

1. Compare the types of reactants in synthesis and single displacement reactions. K/U
2. Use the activity series of metals to predict whether each of these reactions occurs. Write a balanced chemical equation for each reaction that does occur. Write ***no reaction*** for those that do not. K/U C
   (a) $Al(s) + AgNO_3(aq) \rightarrow$
   (b) $Zn(s) + Pb(NO_3)_2(aq) \rightarrow$
   (c) $Au(s) + H_2O(l) \rightarrow$
   (d) $Mg(s) + H_2SO_4(aq) \rightarrow$
   (e) $Ca(s) + H_2O(l) \rightarrow$
   (f) $Al(s) + HCl(aq) \rightarrow$
3. Use the halogen activity series to predict whether each of these reactions occurs. Write a balanced chemical equation for each reaction that does occur. Write ***no reaction*** for those that do not. K/U C
   (a) $Br_2(l) + NaI(aq) \rightarrow$
   (b) $Cl_2(g) + KF(aq) \rightarrow$
   (c) $F_2(g) + CaBr_2(aq) \rightarrow$
4. Use the activity series to rank these reactions in order from slowest to fastest. Justify your prediction. K/U T/I
   (a) $Zn(s) + AgNO_3(aq) \rightarrow$
   (b) $Mg(s) + AgNO_3(aq) \rightarrow$
   (c) $K(s) + AgNO_3(aq) \rightarrow$
   (d) $Au(s) + AgNO_3(aq) \rightarrow$
5. Some pieces of jewellery react with compounds in sweat. The reaction can leave stains on your skin. Not all jewellery has this effect. Use the activity series of metals to explain this observation. K/U T/I
6. **Figure 10** shows magnesium metal burning brightly inside a block of dry ice (solid carbon dioxide). One of the products of this reaction is elemental carbon. K/U C A

Figure 10 Magnesium burning in dry ice

   (a) Write a chemical equation for this reaction.
   (b) Fires involving reactive metals such as magnesium are difficult to extinguish. Why is pouring sand on the fire better than attempting to use a carbon dioxide fire extinguisher?
7. Some pipes supplying drinking water to old homes are made of lead. Use the activity series to explain why many homeowners are replacing these pipes with copper pipes. K/U A
8. **Table 1** summarizes the reaction of metals A, B, and C in different temperatures of water. Based on this evidence, describe the position of these metals relative to each other and to hydrogen on the activity series. Justify your prediction. K/U T/I

Table 1 Activity of Three Metals in Water of Different Temperatures

	Metals		
Temperature of water	**A**	**B**	**C**
cold	NR	reaction	NR
room temperature	NR	reaction	slight reaction
hot	NR	reaction	reaction

9. **Figure 11** shows three metals, A, B, and C, in hydrochloric acid. K/U T/I
   (a) Compare the relative position of these metals and hydrogen on the activity series.
   (b) Which reactive metal is highest in the activity series? Justify your answer.
   (c) If the three metals were copper, magnesium, and zinc, which letter would correspond to which metal?

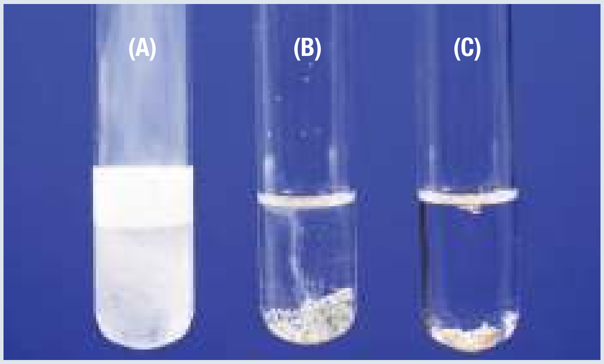

Figure 11 Three metals in acid

10. Most metals exist naturally as compounds rather than as elements. Extracting metals is far more complicated than simply decomposing their compounds. The Goldschmidt process was developed to extract important metals from their ores. Research this process and how it produces iron. Write a chemical equation for the reaction. T/I C A

GO TO NELSON SCIENCE

The Mystery of the Missing Mercury

SKILLS HANDBOOK A4

ABSTRACT

Mercury and ozone are two industrial pollutants that accumulate in the Arctic air during winter. Each spring, both substances mysteriously vanish. Mercury is a particular concern because of its toxicity and ability to accumulate in food chains. At first, these depletion events were thought to be unrelated. Recent discoveries suggest that they are caused by bromine released from seawater. The chemical reactions involved and the ultimate fate of the lost mercury are still the subject of ongoing research.

Introduction

Each spring, as the first rays of the Sun glisten across the Arctic ice and snow, a strange disappearing act occurs. All the mercury in the Arctic air suddenly vanishes. Several months later, as the dark of the Arctic winter returns, mercury reappears. Canadian scientists first noticed this peculiar cycling of mercury in 1982 and have been studying it ever since (**Figure 1**). Researchers led by Jan Bottenheim of Environment Canada want to get to the bottom of why mercury vanishes from the atmosphere. Perhaps even more importantly, they want to find out where it goes.

Figure 1 A team of scientists braves the frozen north to discover why levels of mercury cycle up and down with the seasons. Dr. Bottenheim (left) called this event "amazing and mind-boggling."

The questions are particularly important because mercury is such a dangerous element. It is a potent neurotoxin, capable of inflicting serious damage to the nervous system. The whereabouts of the missing element could have implications for both humans and the environment.

Sources of Mercury

Like gold, mercury is one of the few metals that are found in nature as pure elements. It is the only metal that is liquid at room temperature (**Figure 2**). Like any liquid, mercury gradually evaporates into the air, becoming a gas even at low temperatures. You may have seen mercury in thermometers, although this type of thermometer is being phased out because of the danger if it breaks. Mercury is also used in fluorescent lights.

Figure 2 Most schools have banned the use of mercury thermometers because mercury is so toxic.

Some mercury enters the atmosphere as a result of natural processes such as volcanic activity. However, as much as two-thirds of atmospheric mercury is released by human activities such as the combustion of coal. Canada and the United States burn vast quantities of coal to generate power, refine metals, and incinerate waste. Coal is mostly carbon, but it contains some impurities: traces of mercury, sulfur, and other elements.

Clues Emerge

The first clue to solving the missing mercury mystery came from an unlikely source. Ground-level ozone, $O_3(g)$, is another air pollutant. Like mercury, ozone is produced largely by human activities in the south and redistributed by air currents (wind) throughout the atmosphere. By coincidence, scientists noticed that ozone also vanishes from the Arctic air each spring. Is sunlight providing the energy needed for ozone and mercury to react and form new compounds? Or is another factor involved? Bottenheim's team knew that ozone depletion is usually caused by halogens or their compounds. For example, chlorofluorocarbons, or CFCs, were largely responsible for the depletion of the ozone layer in the 20th century. But CFCs have been banned for decades. Could there be a natural source of halogens in the Arctic?

A Salty Solution to the Mystery

Bottenheim accidentally found the answer to this question while examining data collected by another Arctic research team. Sure enough, Arctic air contains bromine—a halogen! Furthermore, the concentration of bromine in the air rises each spring, just when levels of ozone and mercury fall. The fluctuations appeared to be connected. According to Bottenheim, "this started the whole ball rolling."

The team soon determined that the source of the atmospheric bromine was seawater. Seawater contains dissolved bromide ions, $Br^-(aq)$. Bromide is a stable ion because it has eight valence electrons (**Figure 3(a)**). Scientists believe that sunlight causes bromide ions to lose one electron to form highly reactive bromine atoms, Br^0 (**Figure 3(b)**).

Figure 3 (a) Lewis symbol of a bromide ion (b) Lewis symbol of a bromine atom

Bromine atoms are very reactive, due to their incomplete valence electron orbits. Consequently, these atoms promptly react with oxygen in the air to form highly reactive compounds such as bromine monoxide, BrO (**Figure 4**).

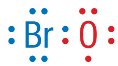

Figure 4 Lewis structure of a bromine monoxide molecule. This molecule is reactive because of the single electron on the oxygen atom.

Scientists theorize that both bromine and bromine monoxide react with elemental mercury in the air, pulling away two of its electrons. This leaves a mercury(II) ion, Hg^{2+}. The ion then attaches itself to snowflakes and falls to the ice surface. Thus, almost all of the mercury is removed from the air and trapped in the ice and snow.

The Arctic Sun fuels further photochemical reactions, causing the ice-bound mercury ions to regain two electrons. They now no longer have a positive charge and lose their attraction to the snow and ice crystals.

Some of this mercury is absorbed by micro-organisms and converted to compounds such as methyl mercury, CH_3Hg. This compound is dangerous because it readily accumulates in the fat tissue of larger organisms as they eat these micro-organisms. However, this process does not account for all the mercury that was in the air. The ultimate fate of the missing mercury remains unclear, and the research continues.

Further Reading

Abel, A. (2010). The case of the missing mercury. *Canadian Geographic*, 130(1), 56–61.

Sherman, L.S., Blum, J.D., Johnson, K.P., et al. (2010). Mass-independent fractionation of mercury isotopes in Arctic snow driven by sunlight. *Nature*, 355, 150–152.

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4.5 Questions

1. Mercury is an industrial pollutant. Why is mercury showing up in the Arctic where there is little or no heavy industry? K/U
2. Compare the reactivity of a bromide ion and a bromine atom. Why is there a difference? K/U
3. Why are carbon compounds that contain mercury a greater environmental concern than elemental mercury? K/U
4. Use the activity series of metals to justify why mercury can remain for one or two years in the atmosphere without reacting. K/U
5. High-temperature combustion has been used to safely dispose of some toxic industrial compounds. Why is mercury not "destroyed" when coal containing mercury impurities is burned? K/U
6. Give an example of the role of luck in Bottenheim's research. K/U
7. Why do you think mercury concentrations may be highest in organisms at the top of Arctic food chains? T/I A
8. Scientists predict that the area of the Arctic Ocean covered by sea ice will decrease in the near future. What impact might this have on how quickly mercury disappears in the spring? Why? A
9. Research in remote areas like the Arctic is expensive. Much of the funding comes from government sources. What is your opinion about funding for the research to solve the disappearing mercury problem? Is the funding warranted? Is it a waste of money? Explain. T/I

4.6 Double Displacement Reactions

Industrial processes produce unwanted by-products (**Figure 1**). Dissolved toxic metal ions—copper, mercury, and cadmium—are common leftovers in the wastewater. These toxins must be removed before the water can be released into the environment. In Section 4.4, you learned that some metal ions can be recovered as solid metals by allowing a solution of the ions to react with a more active metal.

Figure 1 Refining metals, such as these sheets of pure copper, produces toxic metal ions. These ions must be removed before the wastewater can be discharged.

Figure 2 Precipitates form when a solution of sodium hydroxide is added to solutions of metal compounds.

double displacement reaction a reaction in which elements in two compounds displace each other or trade places, producing two new compounds; general pattern:

$AB + CD \rightarrow AD + CB$

Another way to remove unwanted ions is to precipitate them from solution. A precipitate is the solid that forms as a result of the reaction of two solutions. Chemists have found that adding sodium hydroxide to a mixture of waste metal ions causes most of the toxic ions to form metal-hydroxide precipitates (**Figure 2**). The mixture is then filtered to remove the precipitates. Copper ions can be removed from solutions containing waste copper(II) sulfate by adding sodium hydroxide solution.

$$CuSO_4(aq) + 2\ NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$$

A closer look at the equation shows that sodium and copper ions in the reactants have traded places or displaced each other. A chemical reaction in which two elements from different compounds displace each other is called a **double displacement reaction**. The general equation for a double displacement reaction is

$AB + CD \rightarrow AD + CB$ (**Figure 3**)

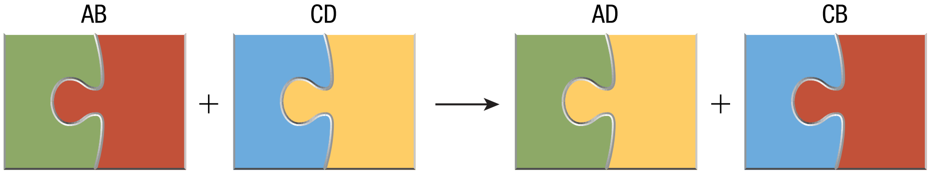


Figure 3 In a double displacement reaction, two elements, A and C, displace each other from their compounds. The products are two new compounds.

Types of Double Displacement Reactions

Double displacement reactions vary, resulting in different products. Some produce precipitates, some release gases, and others result in solutions that are more neutral than the reactant solutions. We can classify double displacement reactions according to the type of product they generate.

Precipitation Reactions

When two solutions react, positive ions (cations) in one solution attract and bond to negative ions (anions) of the other solution. If the new compound precipitates, it must be only slightly soluble in the liquid. Before we can predict which compound precipitates, we need to understand some key ideas related to solubility. A more detailed discussion of solubility is given in Unit 4: Solutions and Solubility.

SOLUBILITY

A **solution** is a homogeneous mixture of a solute dissolved in a solvent. For example, an aqueous salt solution is a mixture of salt (the **solute**) dissolved in water (the **solvent**). **Solubility** is defined as the quantity of solute that will dissolve in a given quantity of solvent. The solubility of a solute depends on factors such as the type of bonding within the solute, the states of both substances, and the temperature of the solvent.

All substances dissolve in water to some degree. Chemists describe a substance as being very soluble (or highly soluble) if a significant quantity of the substance dissolves. A substance that does not dissolve well is only slightly soluble. In the equation of a precipitation reaction, state symbols indicate which of the products forms a precipitate. The state symbol "(s)" indicates the precipitate; the symbol "(aq)" indicates the highly soluble substance that remains in solution.

solution a homogenous mixture of two or more substances.

solute the substance that dissolves in a solvent

solvent the substance that dissolves the solute

solubility the quantity of solute that dissolves in a given quantity of solvent at a given temperature

Figure 4 Combining solutions of sodium hydroxide and copper(II) sulfate produces a jellylike precipitate of copper(II) hydroxide.

PREDICTING PRECIPITATES

Selecting the correct reactant to precipitate metal ions from solution requires background knowledge of which metal compounds are soluble and which are not. This information is summarized in a solubility table (**Table 1**). This table is a summary of the combinations of cations and anions. It tells us the combinations of ions that produce compounds that are very soluble and compounds that are only slightly soluble. For example, the table states that all hydroxides of Group 1 elements are very soluble. Sodium hydroxide is a compound that includes a Group 1 element. Therefore, we would predict that sodium hydroxide dissolves readily in water. Conversely, most other cations form hydroxide compounds that are only slightly soluble. Therefore, mixing a solution of a soluble Cu^{2+} compound with a sodium hydroxide solution should produce a precipitate of copper(II) hydroxide (**Figure 4**).

Table 1 Solubility of Ionic Compounds at Room Temperature

Solubility	Ion	Exceptions
very soluble (aq) $\geq$ 0.1 mol/L	NO_3^-	none
	Cl^- and other halides	except with Cu^+, Ag^+, Hg_2^{2+}, Pb^{2+}
	SO_4^{2-}	except with Ca^{2+}, Ba^{2+}, Sr^{2+}, Hg^{2+}, Pb^{2+}, Ag^+
	$C_2H_3O_2^-$	Ag^+
	Na^+ and K^+	none
	NH_4^+	none
slightly soluble (s) $<$ 0.1 mol/L	CO_3^{2-}	except with Group 1 ions and NH_4^+
	PO_4^{3-}	except with Group 1 ions and NH_4^+
	OH^-	except with Group 1 ions, Ca^{2+}, Ba^{2+}, Sr^{2+}
	S^{2-}	except with Groups 1 and 2 ions and NH_4^+

The reactants do not both have to be in solution when they are combined. As long as they are both very soluble and are both added to water, they will dissolve, react, and form the precipitate.

SELECTING A REACTANT TO PRODUCE A PRECIPITATE

How could we treat water contaminated with toxic metal ions? We can precipitate the ions out of solution. The solubility table can help us choose an appropriate reactant. We want the reaction to produce a compound of the toxic metal that is only slightly soluble. The toxic ions then form a precipitate and can be filtered out. Sometimes two different reactions are used, one after the other.

To remove a range of metals, chemists select an anion that forms slightly soluble compounds with as many metals as possible. In practice, soluble sulfide compounds are usually used (**Figure 5**).

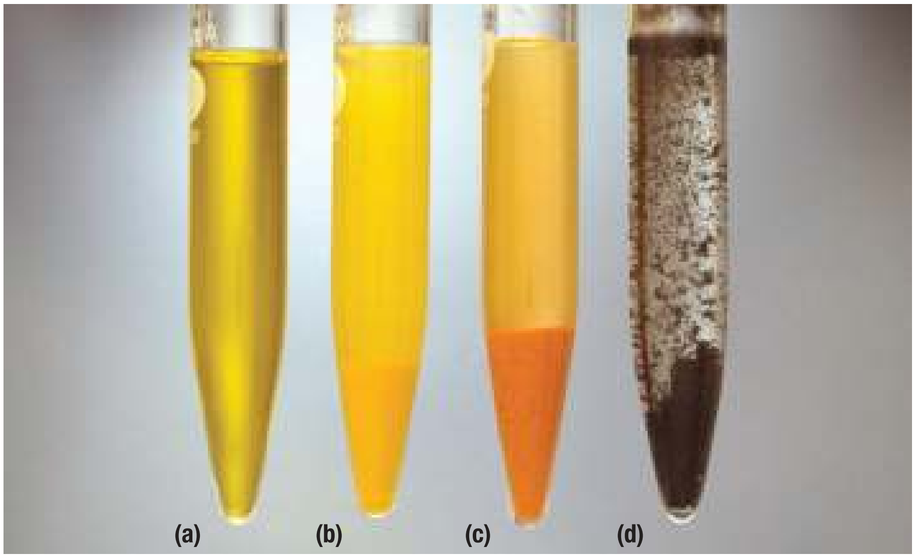


Figure 5 (a) The ammonium ion is one of the few cations whose sulfide, $(NH_4)_2S$, is very soluble. Sulfide compounds of most other cations are only slightly soluble. Examples include (b) cadmium sulfide, CdS, (c) antimony sulfide, Sb_2S_3, and (d) lead(II) sulfide, PbS.

Investigation **4.6.1**

Precipitate Patterns (p. 181)
In this investigation, you will discover which anion can precipitate the most metal cations from solution.

CAREER **LINK**

Research chemists work in many different fields, but all are concerned about the reactants and product of reactions. To find out more about the work of an analytical research chemist,

GO TO NELSON SCIENCE

Chemists can also selectively precipitate certain cations from a mixture by using differences in solubility. For example, the solubility table shows that silver chloride is slightly soluble, but copper(II) chloride is very soluble. Therefore, adding sodium chloride will precipitate only silver ions, $Ag^+(aq)$, from a solution containing both silver ions and copper(II) ions.

Tutorial 1 Predicting Double Displacement Reactions

The following examples illustrate how to predict whether or not a precipitate forms during a double displacement reaction, as well as how to write a balanced chemical equation for the reaction (**Figure 6**).

Sample Problem 1: Reactions Involving Two Ionic Compounds

Will a precipitate form when solutions of potassium chloride and silver nitrate are combined? If you predict that a precipitate will form, write a chemical equation for the reaction. If you predict that no reaction occurs, write ***no reaction***.

Step 1. Write chemical formulas of the reactants.

$KCl(aq)$ $AgNO_3(aq)$

Step 2. Separate the reactants into their ions (Figure 6).

K^+ Cl^- Ag^+ NO_3^-

Step 3. Combine cations and anions to form new compounds.

KNO_3 $AgCl$

Step 4. Check for low solubility.

AgCl is slightly soluble, while KNO_3 is very soluble. Therefore, AgCl should precipitate.

Step 5. Write a balanced chemical equation for the reaction. Label the precipitate with the state symbol (s) and soluble compounds with (aq).

$$KCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + KNO_3(aq)$$

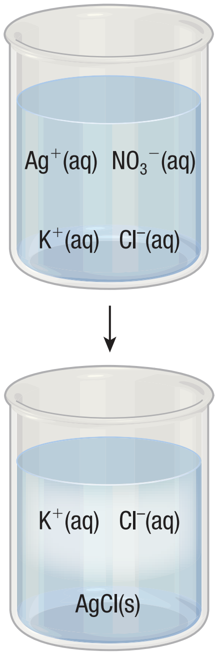


Figure 6 If the ions can combine to form a product that is only slightly soluble, the reaction will occur.

POLYATOMIC IONS

Compounds of polyatomic ions also vary in their solubility. This difference in solubility causes some polyatomic compounds to form precipitates in double displacement reactions. Sample Problem 2 shows how to approach reactions involving polyatomic ions.

Sample Problem 2: Reactions Involving Polyatomic Compounds

Will a precipitate form when solutions of potassium sulfate and iron(III) chloride are combined? If you predict that a precipitate will form, write a chemical equation for the reaction. If you predict that no reaction occurs, write ***no reaction***.

Step 1. Write chemical formulas of the reactants.

$K_2SO_4(aq)$ $\quad$ $FeCl_3(aq)$

Step 2. Separate the reactants into their ions.

K^+ K^+ SO_4^{2-} $\quad$ Fe^{3+} Cl^- Cl^- Cl^-

Step 3. Combine cations and anions to form new compounds.

KCl $\quad$ $Fe_2(SO_4)_3$

Step 4. Check for low solubility.

Both compounds are very soluble, so no precipitate forms: no reaction.

Practice

1. Which of the following combinations result in the formation of a precipitate? Write a balanced chemical equation if you predict that the reaction will occur. K/U C
   (a) $Na_2S(aq) + Pb(NO_3)_2(aq) \rightarrow$
   (b) $NH_4Cl(aq) + K_2SO_4(aq) \rightarrow$
   (c) $FeCl_3(aq) + Na_2CO_3(aq) \rightarrow$

Mini **Investigation**

Testing Water for Ions

Skills: Performing, Observing, Analyzing, Communicating

SKILLS HANDBOOK A1.1, A1.2

Water samples are routinely tested to identify the ions present. Information about the ions in your local drinking water is available from the Ministry of the Environment. The concentration of certain ions in bottled water is often provided on the product label.

A solution of silver nitrate is used to detect the presence of chloride ions in water. A white precipitate of silver chloride is a positive test for chloride ions. Similarly, calcium ions in water can be detected by adding a solution of sodium phosphate, $Na_3PO_4(aq)$. The result is a white calcium phosphate precipitate.

Equipment and Materials: chemical safety goggles; lab apron; safety gloves; test-tube rack; 6 test tubes; masking tape; marker; samples of distilled water, tap water, and bottled water; dropper bottles of dilute silver nitrate, $AgNO_3(aq)$(0.1 mol/L), and dilute sodium phosphate, $Na_3PO_4(aq)$ (0.1 mol/L)

Silver nitrate stains the skin.
Sodium phosphate is an irritant. Avoid skin and eye contact. If you spill these chemicals on your skin, wash the affected area with a lot of cool water and inform your teacher.

1. Put on your safety goggles, lab apron, and safety gloves.
2. Add distilled water, tap water, and bottled water to three test tubes to a depth of about 4 cm. Label the test tubes and place them in the test-tube rack.
3. Create another set of test tubes identical to those in Step 1.
4. Add 4 drops of silver nitrate to each test tube of the first set. Swirl to mix the contents of each test tube.
5. Add 4 drops of sodium phosphate to each test tube of the second set. Swirl to mix the contents of each test tube.

A. Compare the appearance of silver chloride precipitate in the first three test tubes. Suggest reasons to explain any differences you observed. T/I

B. Suggest a human activity that increases the concentration of chloride ions in water. A

C. Compare the appearance of calcium phosphate precipitate in the second three test tubes. T/I

D. Suggest a possible natural source of calcium ions in water. A

Figure 7 The reaction of magnesium carbonate and hydrochloric acid results in the production of a gas.

Reactions That Produce a Gas

Double displacement involving acids can also produce gases. This can occur in two ways. First, gases may be produced directly. For example, the addition of an acid to a sulfide compound like potassium sulfide, K_2S, results in a double displacement reaction that produces hydrogen sulfide gas:

$$K_2S(aq) + 2\ HCl(aq) \rightarrow H_2S(g) + 2\ KCl(aq)$$

Hydrogen sulfide is toxic and has a distinctive odour of rotten eggs. Spilling an acid onto a sulfide compound releases toxic hydrogen sulfide fumes. For this reason, the use of sulfide compounds has been banned in many schools.

Gases can also be produced when an unstable product of a double displacement reaction decomposes. For example, the pattern of double displacement reactions suggests that magnesium carbonate and sulfuric acid react to produce magnesium sulfate and carbonic acid:

$$MgCO_3(aq) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2CO_3(aq)$$

However, carbonic acid is unstable and immediately decomposes into water and bubbles of carbon dioxide (**Figure 7**):

$$H_2CO_3(aq) \rightarrow H_2O\ (l) + CO_2(g)$$

We can combine these two equations to show the overall reaction:

$$MgCO_3(aq) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2O(l) + CO_2(g)$$

Similar reactions occur when acids are added to solutions of sulfite compounds. For example, a solution of sodium sulfite reacts with hydrochloric acid according to the following chemical equation:

$$Na_2SO_3(aq) + 2\ HCl(aq) \rightarrow H_2SO_3(aq) + 2\ NaCl(aq)$$

The sulfite product, H_2SO_3, is sulfurous acid. However, bubbles of toxic sulfur dioxide appear as soon as the reactants are combined. This occurs because sulfurous acid is unstable and quickly decomposes into water and sulfur dioxide:

$$H_2SO_3(aq) \rightarrow H_2O\ (l) + SO_2(g)$$

Again, we can combine these two equations:

$$Na_2SO_3(aq) + 2\ HCl(aq) \rightarrow H_2O(l) + SO_2(g) + NaCl(aq)$$

The production of gases by double displacement reactions is summarized in **Table 2**.

Table 2 Compounds That React with Acids to Produce Gases

Compound	Gas produced
sulfides, e.g., Na_2S	hydrogen sulfide, H_2S
carbonates, e.g., Na_2CO_3	carbon dioxide, CO_2
sulfites, e.g., K_2SO_3	sulfur dioxide, SO_2

Neutralization Reactions

neutralization reaction the reaction of an acid and a base in which the resulting solution has a pH closer to 7 than either of the reactants; a type of double displacement reaction that produces water and an ionic compound

In earlier science courses, you learned that the reaction of an acid with a base is called a **neutralization reaction**. The end result is a mixture with a pH closer to a neutral pH of 7 than either of its reactants. For example, nitric acid, $HNO_3(aq)$, can be neutralized by the addition of a sodium hydroxide solution, NaOH(aq). The chemical equation for the reaction is

$$\begin{array}{ccccccc} HNO_3(aq) & + & NaOH(aq) & \rightarrow & H_2O(l) & + & NaNO_3(aq) \\ AB & + & CD & \rightarrow & AD & + & CB \end{array}$$

If you think of the water molecule as consisting of two parts, H—OH, then a neutralization reaction follows the pattern of double displacement reactions.

Consider another example. Antacids are medications that neutralize excess stomach acid. For example, magnesium hydroxide, $Mg(OH)_2$, is the active ingredient in some antacid medications (**Figure 8**). Magnesium hydroxide neutralizes excess hydrochloric acid in the stomach according to the equation

$$Mg(OH)_2(s) + 2\ HCl(aq) \rightarrow 2\ H_2O(l) + MgCl_2(aq)$$

Figure 8 Magnesium hydroxide has a low solubility in water. This explains why a mixture of magnesium hydroxide and water is a thick white paste rather than a clear solution.

Note that these reactions generally do not produce precipitates. We can detect that a reaction has occurred only by testing the pH of the resulting solution and comparing it with the pH of the reactant solutions.

4.6 Summary

- In a double displacement reaction, two elements trade places to form two new compounds. The net result is that two compounds react to form two new compounds.

  The pattern for these reactions is AB + CD → AD + CB.
- Double displacement reactions can produce a precipitate, a neutralized solution, or a gas.
- Most double displacement reactions produce precipitates that can be predicted using a solubility table.
- The reaction of an acid with a base produces a neutralized solution.

UNIT TASK BOOKMARK

You will be using neutralization reactions in the Unit Task, described on page 242.

4.6 Questions

1. How do single and double displacement reactions differ? K/U
2. Classify the reactions represented by these equations as either single or double displacement: K/U
   (a) $HI(aq) + AgNO_3(aq) \rightarrow AgI(s) + HNO_3(aq)$
   (b) $Fe(s) + CuSO_4(aq) \rightarrow FeSO_4(aq) + Cu(s)$
   (c) $ZnS(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2S(g)$
   (d) $Cl_2(g) + 2\ NH_4Br(aq) \rightarrow Br_2(l) + 2\ NH_4Cl(aq)$
3. Write the chemical formula for each of the following compounds. Predict the solubility of these compounds in water: K/U
   (a) lead sulfate (in car batteries)
   (b) ammonium phosphate (fertilizer)
   (c) calcium sulfate (a component of drywall)
   (d) aluminum sulfate (used in water purification)
   (e) calcium phosphate (in bones)
   (f) barium sulfate (used during stomach X-rays)
   (g) ammonium carbonate (smelling salts)
   (h) calcium carbonate (in shells)
4. Complete the chemical equations of the following reactions. Indicate the state of each compound. K/U C
   (a) $ZnCl_2(aq) + KOH(aq) \rightarrow$
   (b) $Ni(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow$
   (c) $Ba(OH)_2(aq) + K_2SO_4(aq) \rightarrow$
   (d) $FeSO_4(aq) + K_3PO_4(aq) \rightarrow$
   (e) $ZnS(s) + 2\ HCl(aq) \rightarrow$
   (f) $CaCO_3(s) + 2\ HNO_3(aq) \rightarrow$
   (g) $MgSO_3(aq) + HCl(aq) \rightarrow$
5. Will a precipitate form when the following solutions are combined? If you predict that a precipitate will form, write a balanced chemical equation for the reaction. Include states. If you predict that no reaction occurs, write ***no reaction***. K/U C
   (a) $AgNO_3 + K_2SO_4$
   (b) $NH_4Cl + Na_2S$
   (c) $Pb(NO_3)_2 + Na_3PO_4$
   (d) $BaCl_2 + Ca(OH)_2$
   (e) $CuSO_4 + K_2CO_3$
6. Write a balanced chemical equation (with state symbols) for each of the following double displacement reactions: K/U C
   (a) Barium carbonate is used in the ceramics industry as an ingredient in glazes. Write a chemical equation for the reaction of barium carbonate with sulfuric acid.
   (b) Hydrofluoric acid, HF(aq), is used to etch glass, which is mostly silicon dioxide, SiO_2. The products of this reaction are water and silicon tetrafluoride.
   (c) Most of the nickel ions in a nickel(II) nitrate solution can be precipitated by adding a solution of potassium carbonate.
   (d) Sulfuric acid removed from used car batteries is neutralized with sodium carbonate (soda ash).
   (e) Dropping hydrochloric acid on solid zinc sulfide produces a toxic gas with the odour of rotten eggs.
7. Silver compounds are expensive. Metal manufacturers therefore want to recover any waste silver ions from solution. One way to do this is to precipitate the silver as a compound and then separate the compound from the mixture using filtration. Use the solubility table to determine a way of precipitating only silver ions from a mixture of dissolved metal ions. K/U C
8. Research why soaking in vinegar is a convenient method of removing limescale from the heating coils of a kettle. Write and classify the chemical reactions involved. T/I C
9. "Black smokers" emerge from cracks in the ocean floor (**Figure 9**). These black plumes are created by superheated water passing through rock. Research the chemical reactions that cause these plumes to be black. Summarize your findings in a paragraph. K/U C

Figure 9 An ocean phenomenon

GO TO NELSON SCIENCE

Investigation 4.4.1 OBSERVATIONAL STUDY

SKILLS MENU

• Questioning	• Planning	• Observing
• Researching	• Controlling Variables	• Analyzing
• Hypothesizing		• Evaluating
• Predicting	• Performing	• Communicating

Hydrogen Blast-Off

Hydrogen is an ideal rocket fuel because it releases more energy per tonne when it burns than any other fuel. The chemical equation for the combustion of hydrogen is

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g) + \text{energy}$$

However, a successful launch requires the proportions of hydrogen to oxygen to be just right. In this investigation, you will produce hydrogen and oxygen using these chemical reactions:

$$Mg(s) + 2\ HCl(aq) \rightarrow H_2(g) + MgCl_2(aq)$$

$$2\ H_2O_2(aq) \xrightarrow{\text{yeast}} 2\ H_2O(l) + O_2(g)$$

You will prepare two different mixtures of hydrogen and oxygen. You will test the mixtures by squeezing them into a candle flame to see which produces the loudest "pop."

Purpose

SKILLS HANDBOOK A2.4

To discover which ratio of hydrogen to oxygen produces the most energy (the loudest "pop") when burned:
2 parts hydrogen to 1 part oxygen
or 1 part hydrogen to 2 parts oxygen?

Equipment and Materials

SKILLS HANDBOOK A1.1, A1.2

- chemical safety goggles
- lab apron
- scissors
- marker
- well plate
- one-holed stopper (that fits the wells of the well plate)
- candle or Bunsen burner clamped to a retort stand
- spark lighter
- toothpick
- dropper bottles of
  - 6 % hydrogen peroxide solution, $H_2O_2(aq)$
  - dilute hydrochloric acid, $HCl(aq)$ (0.1 mol/L)
- 2 plastic pipettes
- 1 cm strip of magnesium, $Mg(s)$
- yeast

An open flame is used. Tie back long hair and secure loose clothing. Never leave the flame unattended.

The hydrochloric acid and hydrogen peroxide solutions used in this activity are irritants. Wash any spills on skin or clothing immediately with plenty of cool water. Report any spills to your teacher.

Procedure

Part A: Preparing the Gas-Collecting Bulbs

1. Prepare two gas-collecting bulbs and one gas delivery stopper assembly, using **Figure 1** as a guide.

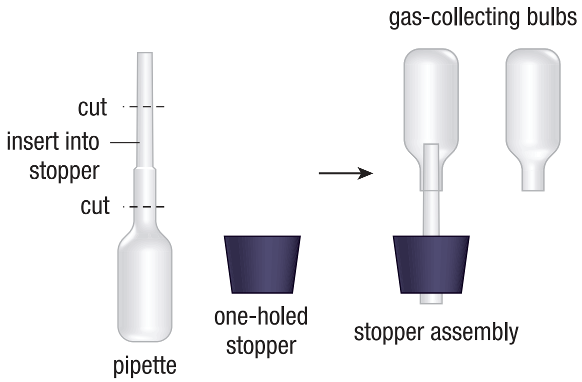


Figure 1 Preparing the gas-collecting bulbs and stopper assembly

2. Label one bulb as $1H_2{:}2O_2$ and the other as $2H_2{:}1O_2$. Draw two lines on each bulb, indicating one-third and two-thirds of the volume.
3. Fill the two bulbs with water. Stand them, open end upward, in the well plate.

Part B: Collecting Hydrogen

4. Put on your chemical safety goggles and lab apron.
5. Add 15 drops of hydrochloric acid to a clean, dry well of the well plate.
6. Add a 1 cm strip of magnesium to the same well.
7. Seal the well with the stopper assembly. Wait 5 s for the hydrogen gas being produced to displace the air in the well.

8. Place the "$1H_2$:$2O_2$" gas-collecting bulb full of water onto the stopper assembly (**Figure 2**). The fit should be loose to allow water to be displaced.

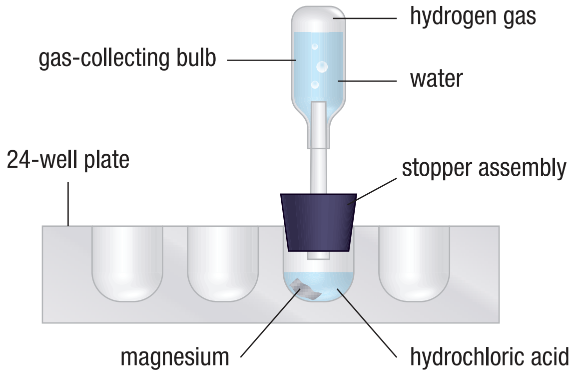


Figure 2 Setup for the collection of hydrogen gas. Leave an air space between the solution and the stopper to allow gas to go up the tube.

9. Allow one-third of the bulb to fill with hydrogen gas. (The remaining two-thirds will be water.)
10. Remove the bulb and stand it in a well plate. Always keep the opening of the bulb at the bottom; otherwise, the gas will escape.
11. Place the "$2H_2$:$1O_2$" gas-collecting bulb full of water onto the stopper assembly.
12. Allow two-thirds of this bulb to fill with hydrogen gas. (The remaining one-third will be water.)
13. Remove the bulb and stand it in the well plate, as in Step 10.

Part C: Collecting Oxygen Gas

14. Add 15 drops of hydrogen peroxide to a clean, dry well.
15. Use the wide end of the toothpick to add a few grains of yeast.
16. Seal the well with the gas delivery stopper (without a bulb). Ensure that the bottom of the tube is not blocked.
17. Wait 5 s for the oxygen gas being produced to displace the air in the well.
18. Place the "$2H_2$:$1O_2$" bulb that is two-thirds full of hydrogen gas onto the stopper assembly (**Figure 3**).
19. Allow the oxygen gas to displace all but about two drops of water in the bulb. These drops act as a plug to keep the collected gases in the bulb.
20. Return the bulb to its well.
21. Repeat Steps 18 to 20 with the "$1H_2$:$2O_2$" bulb.

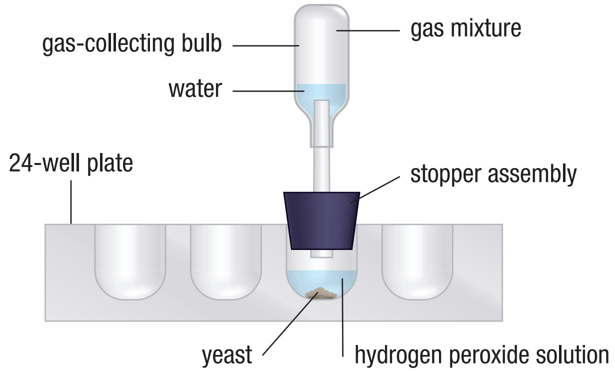


Figure 3 Setup for the collection of oxygen gas

Part D: Testing the Gas Mixtures

22. Light a candle or burner.
23. Without tilting it, bring the "$1H_2$:$2O_2$" bulb to within 2 cm of the flame (**Figure 4(a)**).
24. Quickly rotate the bulb and squirt its contents into the flame (**Figure 4(b)**). Record your observations.

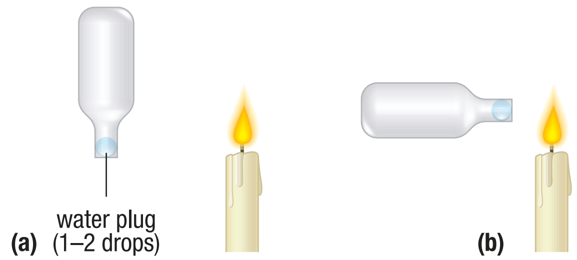


Figure 4 Keep the bulb pointing down until you are ready to push the gases into the flame.

25. Repeat Steps 23 and 24 with the other bulb. Record your observations.

Analyze and Evaluate

(a) Classify each of the three reactions (in Parts B, C, and D) that you observed in this activity. K/U T/I

(b) Which gas mixture would be better for launching a rocket? Support your answer with evidence. T/I

(c) How would your results differ if you did not wait a few seconds before collecting each gas? T/I

(d) What role does the yeast play in this investigation? T/I

(e) What was the most difficult challenge you faced during the investigation? How did you overcome that challenge? T/I

Apply and Extend

(f) Predict how the reaction of a 1:1 mixture of hydrogen to oxygen would compare with the reactions you tested in this investigation. Justify your prediction. T/I

Investigation 4.4.2 CONTROLLED EXPERIMENT

An Activity Series of Ions

Copper is above silver on the activity series (Figure 4 in Section 4.4). You would therefore expect that a single displacement reaction occurs when copper wire is placed into a solution of silver nitrate (**Figure 1**).

$$Cu(s) + 2\ AgNO_3(aq) \rightarrow 2\ Ag(s) + Cu(NO_3)_2(aq)$$

Silver nitrate is soluble in water, so the silver and nitrate ions separate as the compound dissolves. Therefore, a silver nitrate solution is actually a mixture of silver ions, $Ag^{+}(aq)$, and nitrate ions, $NO_3^{-}(aq)$, in water. So, the reaction is really between copper metal and silver ions. The nitrate ions just remain in solution. However, a similar reaction does not occur if a strip of silver is placed in a solution containing $Cu^{2+}(aq)$ ions. It seems that metal ions—as well as metal elements—differ in their reactivity.

Figure 1 The reaction of copper wire in a solution of silver nitrate. The blue colour of the solution is evidence that $Cu^{2+}(aq)$ ions are produced.

Testable Question

What is the trend of reactivity of metal ions?

Hypothesis

SKILLS HANDBOOK A2.2

Predict the trend in reactivity of the metal ions used in this experiment.

Variables

Identify all major controlled, manipulated, and responding variables in this experiment.

Equipment and Materials

SKILLS HANDBOOK A1.1, A1.2

- chemical safety goggles
- lab apron
- well plate
- sand paper
- clean strips of magnesium, Mg, zinc, Zn, iron, Fe, tin, Sn, and copper, Cu
- dropper bottles of dilute solutions of
  - magnesium sulfate, $MgSO_4$ (0.1 mol/L)
  - zinc sulfate, $ZnSO_4$ (0.1 mol/L)
  - iron(II) sulfate, $FeSO_4$ (0.1 mol/L)
  - tin(II) chloride, $SnCl_2$ (0.1 mol/L)
  - copper(II) sulfate, $CuSO_4$ (0.1 mol/L)

SKILLS MENU

• Questioning	• Planning	• Observing
• Researching	• Controlling Variables	• Analyzing
• Hypothesizing	• Performing	• Evaluating
• Predicting		• Communicating

Copper(II) sulfate and tin(II) chloride are both toxic and irritants. Iron(II) sulfate is also an irritant. Avoid skin and eye contact. If you spill these chemicals on your skin, wash the affected area with plenty of cool water and inform your teacher.

Experimental Design

You will place five metals in different metal ion solutions to see which combinations of metals and ions result in single displacement reactions.

Procedure

1. Design a procedure to determine which metal/metal ion solution combinations result in a single displacement reaction. Include safety precautions.
2. Proceed, with your teacher's approval.
3. Record that a reaction occurred only when a new solid forms. Disregard any gas bubbles that may form.
4. Dispose of the contents of the well plate as directed by your teacher.
5. Clean your workstation and wash your hands.

Observations

Organize your observations in a table similar to **Table 1**.

Table 1 Observation Table

Metal	Ion				
	$Mg^{2+}(aq)$	$Zn^{2+}(aq)$	$Fe^{2+}(aq)$	$Sn^{2+}(aq)$	$Cu^{2+}(aq)$
Mg(s)					

Analyze and Evaluate

(a) What variables were measured/recorded and/or manipulated in this investigation? What type of relationship was being tested? T/I

(b) Answer the Testable Question that was posed at the beginning of this investigation. Support your answer with evidence. T/I

(c) Compare the trends in reactivity of metals with trends in the reactivity of their ions. T/I

(d) Use your answer to question (c) and the activity series of metals to rank these ions in order from least to most reactive: Ni^{2+}(aq), Ba^{2+}(aq), Hg^{2+}(aq). Justify your prediction. T/I

(e) Write a chemical equation for each single displacement reaction recorded in your observation table. K/U C

Apply and Extend

(f) Some nuclear power plants use seawater as a coolant. Choosing the correct type of piping to carry seawater is critical because seawater contains trace amounts of dissolved silver and gold ions. Is copper piping a suitable choice? Why? A

Investigation 4.6.1 OBSERVATIONAL STUDY

SKILLS MENU

• Questioning	• Planning	• Observing
• Researching	• Controlling Variables	• Analyzing
• Hypothesizing		• Evaluating
• Predicting	• Performing	• Communicating

Precipitate Patterns

Wastewater from metal manufacturing processes often contains toxic metal cations. These ions must be removed before the water can be released into the environment. Many of these ions can be removed using precipitation reactions. For example, recycled car batteries contain toxic lead cations, Pb^{2+}(aq). These can be removed by adding a solution of potassium iodide, KI(aq) (**Figure 1**).

In this investigation, you will combine different pairs of solutions to discover which cation/anion combinations result in the formation of a precipitate.

Figure 1 Lead and iodide ions combine to form a precipitate of lead(II) iodide, PbI_2(s).

Purpose

SKILLS HANDBOOK A2.4

To identify which anion could be used to precipitate the most metal cations from an unknown mixture

Equipment and Materials

SKILLS HANDBOOK A1.1, A1.2

- chemical safety goggles
- lab apron
- safety gloves
- clear, colourless well plate
- dropper bottles of dilute solutions of
  - sodium sulfate, Na_2SO_4(aq) (0.20 mol/L)
  - sodium carbonate, Na_2CO_3(aq) (0.15 mol/L)
  - sodium chloride, NaCl(aq) (0.15 mol/L)
  - calcium nitrate, $Ca(NO_3)_2$(aq) (0.20 mol/L)
  - copper(II) nitrate, $Cu(NO_3)_2$(aq) (0.15 mol/L)
  - iron(III) nitrate, $Fe(NO_3)_3$(aq) (0.15 mol/L)
  - potassium nitrate, KNO_3(aq) (0.15 mol/L)
  - magnesium nitrate, $Mg(NO_3)_2$(aq) (0.15 mol/L)
  - silver nitrate, $AgNO_3$(aq) (0.15 mol/L)
  - zinc nitrate, $Zn(NO_3)_2$(aq) (0.15 mol/L)

Copper(II) nitrate is toxic and an irritant. Iron(III) nitrate is an irritant. Silver nitrate is toxic and stains the skin. Avoid skin and eye contact. If you spill these chemicals on your skin, wash the affected area with a lot of cool water.

Procedure

Your teacher may choose to demonstrate a reaction involving the silver ion.

1. Design a procedure to determine which anion precipitates the most cations. Include an observation table in your design. Also include safety precautions.
2. Proceed, with your teacher's approval.
3. Dispose of the contents of the well plate as directed.
4. Clean your workstation and wash your hands.

Analyze and Evaluate

(a) Evaluate the design of your investigation. How could it have been improved? T/I

(b) Based on your evidence, which anion could be used to precipitate most of the metal cations? T/I

(c) Which anion could be used to selectively remove silver ions from solution? Why? T/I

(d) What evidence suggests that nitrate compounds are soluble in water? T/I

(e) Write the chemical formula for each precipitate that formed. K/U C

Apply and Extend

(f) Write a balanced chemical equation for each precipitation reaction that occurred. K/U C

(g) Why is it necessary to use distilled water to prepare the solutions used in this investigation? T/I

(h) "Hard" water contains a high concentration of calcium ions. Suggest a way to make hard water "softer." T/I

Summary Questions

1. Create a question-and-answer game based on the Key Concepts listed in the margin on page 150. Divide the questions into four or five categories. The questions within each category should become increasingly complex. Try to anticipate questions that you think might appear on a test for this chapter.
2. Look back at the Starting Points questions on page 150. Answer these questions using what you have learned in this chapter. Compare your latest answers with those that you wrote at the beginning of the chapter. Note how your answers have changed.

Vocabulary

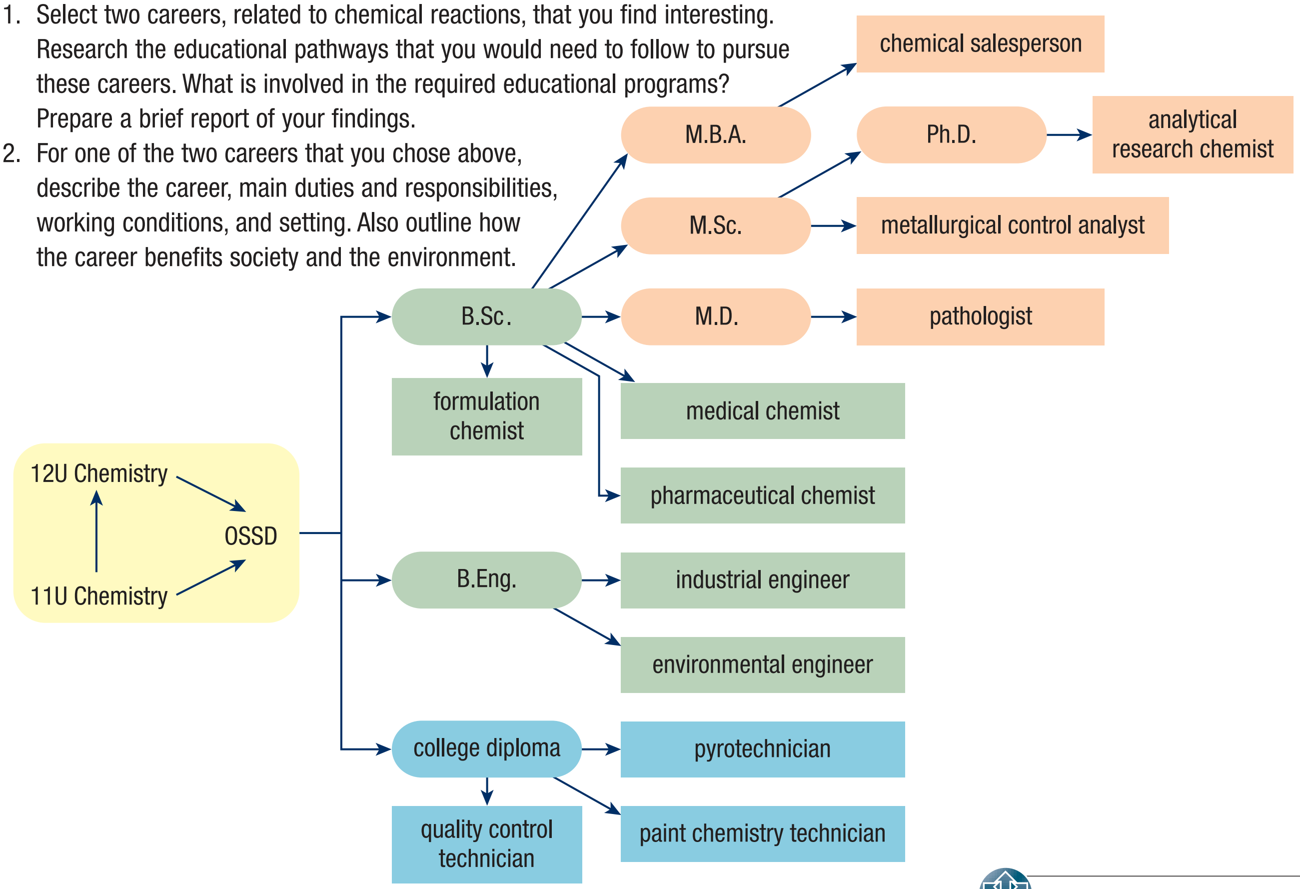
chemical reaction (p. 152)	synthesis reaction (p. 156)	activity series (p. 165)	solute (p. 173)
precipitate (p. 152)	decomposition reaction (p. 159)	double displacement reaction (p. 172)	solvent (p. 173)
catalyst (p. 152)	single displacement reaction (p. 164)	solution (p. 173)	solubility (p. 173)
law of conservation of mass (p. 152)			neutralization reaction (p. 176)

CAREER PATHWAYS

SKILLS HANDBOOK A7

Grade 11 Chemistry can lead to a wide range of careers. Some require a college diploma or a BSc degree. Others require specialized or postgraduate degrees. Here are just a few pathways to careers mentioned in this chapter.

1. Select two careers, related to chemical reactions, that you find interesting. Research the educational pathways that you would need to follow to pursue these careers. What is involved in the required educational programs? Prepare a brief report of your findings.
2. For one of the two careers that you chose above, describe the career, main duties and responsibilities, working conditions, and setting. Also outline how the career benefits society and the environment.



GO TO NELSON SCIENCE

SELF-QUIZ

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

For each question, select the best answer from the four alternatives.

1. Which of the following is a balanced chemical equation? (4.1) K/U
   (a) $Al(s) + Cl_2(g) \rightarrow Al_2Cl_6(s)$
   (b) $P_4(s) + 5\ O_2(g) \rightarrow P_4O_{10}(s)$
   (c) $3\ N_2(g) + 2\ H_2(g) \rightarrow 2\ NH_3(g)$
   (d) $H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$
2. What coefficient should be placed in front of HF to balance the chemical equation below?
   $UO_2(s) + HF(g) \rightarrow UF_4(s) + 2\ H_2O(l)$ (4.1) K/U
   (a) 1
   (b) 2
   (c) 3
   (d) 4
3. Which of the following equations shows the general pattern for a synthesis reaction? (4.2) K/U
   (a) $A + B \rightarrow AB$
   (b) $AB + CD \rightarrow AD + CB$
   (c) $AB \rightarrow A + B$
   (d) $AB + C \rightarrow A + CB$
4. Which of the following equations represents a decomposition reaction? (4.2) K/U
   (a) $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$
   (b) $Fe(s) + HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$
   (c) $2\ NH_3(g) \rightarrow N_2(g) + 3\ H_2(g)$
   (d) $S + O_2 \rightarrow SO_2$
5. Which of the following equations represents a single displacement reaction? (4.4) K/U
   (a) $4\ Fe(s) + 3\ O_2(aq) \rightarrow 2\ Fe_2O_3(s)$
   (b) $2\ H_2O(l) \rightarrow 2\ H_2(g) + O_2(g)$
   (c) $Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
   (d) $Ca(OH)_2(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + 2\ H_2O(l)$
6. Based on the activity series of metals, which of the following reactions occurs? (4.4) K/U
   (a) $Ni(s) + H_2SO_4(aq) \rightarrow$
   (b) $Cu(s) + H_2SO_4(aq) \rightarrow$
   (c) $Hg(s) + H_2SO_4(aq) \rightarrow$
   (d) $Au(s) + H_2SO_4(aq) \rightarrow$
7. Which halogen is most reactive at room temperature? (4.4) K/U
   (a) $Cl_2(g)$
   (b) $I_2(s)$
   (c) $Br_2(l)$
   (d) $F_2(g)$
8. What type of reaction is shown below?
   $Fe(s) + Cu(NO_3)_2(aq) \rightarrow Fe(NO_3)_2(aq) + Cu(s)$
   (4.2, 4.4, 4.6) K/U
   (a) single displacement
   (b) double displacement
   (c) synthesis
   (d) decomposition
9. What type of reaction occurs when hydrogen fluoride solution reacts with potassium hydroxide solution? (4.2, 4.4, 4.6) K/U
   (a) single displacement
   (b) neutralization
   (c) decomposition
   (d) synthesis
10. If lead metal and hydrogen bromide solution are combined, what type of reaction will occur? (4.2, 4.4, 4.6) K/U
    (a) double displacement
    (b) synthesis
    (c) decomposition
    (d) single displacement

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

11. A catalyst makes a reaction go slower. (4.1) K/U
12. A colour change is usually evidence that a chemical reaction has occurred. (4.1) K/U
13. In a synthesis reaction, a new element is produced from two simpler elements. (4.2) K/U
14. In a single displacement reaction, one element replaces another in a compound. (4.4) K/U
15. A neutralization reaction results in a solution that has a higher pH than the pH of the reactant solutions. (4.6) K/U
16. A precipitation reaction always has an acid and a base as reactants. (4.6) K/U
17. In a double displacement reaction, a highly soluble product forms a precipitate. (4.6) K/U

To do an online self-quiz,
GO TO NELSON SCIENCE

Knowledge

For each question, select the best answer from the four alternatives.

1. Which statement correctly describes the law of conservation of mass? (4.1) K/U
   (a) The mass of the products equals the mass of the reactants.
   (b) If a reaction has two products, their masses are equal.
   (c) The mass of the products is less than the mass of the reactants.
   (d) If a reaction has two reactants, their masses are equal.
2. Which of the following chemical equations is balanced? (4.1) K/U
   (a) $CS_2(l) + 3\ O_2(g) \rightarrow CO_2(g) + SO_2(g)$
   (b) $CS_2(l) + 3\ O_2(g) \rightarrow CO_2(g) + 2\ SO_2(g)$
   (c) $CS_2(l) + O_2(g) \rightarrow CO_2(g) + SO_2(g)$
   (d) $CS_2(l) + 3\ O_2(g) \rightarrow CO_2(g) + 3\ SO_2(g)$
3. Which of the following chemical equations is balanced? (4.1) K/U
   (a) $Fe_3O_4(s) + 2\ C(s) \rightarrow Fe(s) + 2\ CO_2(g)$
   (b) $AlCl_3(aq) + Ca(s) \rightarrow Al(s) + CaCl_2(aq)$
   (c) $NO(g) + Br_2(g) \rightarrow NOBr(g)$
   (d) $Ca(OH)_2(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + 2\ H_2O(l)$
4. Which of the following statements correctly describes a decomposition reaction? (4.2) K/U
   (a) A compound breaks down into smaller compounds or elements.
   (b) Two compounds combine to make a larger compound.
   (c) Two elements switch places between compounds.
   (d) A compound changes state.
5. Which equation most accurately represents the synthesis reaction between lithium and oxygen? (4.2) K/U
   (a) $Li(s) + O_2(g) \rightarrow LiO_2(s)$
   (b) $2\ LiO(s) \rightarrow 2\ Li(s) + O_2(g)$
   (c) $6\ Li(s) + O_2(g) \rightarrow 2\ Li_3O(s)$
   (d) $4\ Li(s) + O_2(g) \rightarrow 2\ Li_2O(s)$
6. Which of the following equations represents a single displacement reaction? (4.4) K/U
   (a) $HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(aq)$
   (b) $S(s) + O_2(g) \rightarrow SO_2(g)$
   (c) $Li(s) + KCl(aq) \rightarrow K(s) + LiCl(aq)$
   (d) $2\ H_2O_2(aq) \rightarrow 2\ H_2O(l) + O_2(g)$
7. The activity series of metals shows
   (a) the relative atomic numbers of various metals
   (b) the relative reactivity of various elements
   (c) whether or not a compound is soluble
   (d) the ionic charges of different metallic compounds (4.4) K/U
8. What type of reaction is represented by the equation below?
   $2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$ (4.2, 4.4, 4.6) K/U
   (a) single displacement
   (b) double displacement
   (c) neutralization
   (d) decomposition
9. What type of reaction is represented by the equation below?
   $LiCl(aq) + AgNO_3(aq) \rightarrow LiNO_3(aq) + AgCl(s)$
   (4.2, 4.4, 4.6) K/U
   (a) single displacement
   (b) double displacement
   (c) synthesis
   (d) decomposition
10. Which of the following equations represents a precipitation reaction? (4.6) K/U
    (a) $H_2O(g) + CO(g) \rightarrow H_2(g) + CO_2(g)$
    (b) $BaCO_3(s) \rightarrow BaO(s) + CO_2(g)$
    (c) $CaO(s) + SiO_2(s) \rightarrow CaSiO_3(s)$
    (d) $LiBr(aq) + AgNO_3(aq) \rightarrow LiNO_3(aq) + AgBr(s)$

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

11. If a substance changes colour, a chemical change must have occurred. (4.1) K/U
12. A precipitate speeds the rate of a chemical reaction without being used up. (4.1) K/U

13. Two simple reactants combine to form a more complex product in a decomposition reaction. (4.2) K/U
14. Calcium carbonate ($CaCO_3$) forms calcium oxide and oxygen gas when heated. (4.2) K/U
15. Plasma gasification uses electrical sparks to achieve extremely high temperatures. (4.3) K/U
16. We can use the activity series to determine whether a synthesis reaction will occur. (4.3) K/U
17. A single displacement reaction takes the form $A + BC \rightarrow ABC$. (4.4) K/U
18. Gold is the only metal that is liquid at room temperature. (4.5) K/U
19. A neutralization reaction between an acid and a base results in products that are more neutral than the reactants. (4.6) K/U

Match each chemical equation from the first list with the correct type of reaction from the second list.

20. (a) $MgO(aq) + 2\,HCl(aq) \rightarrow MgCl_2 + H_2O$
    (b) $2\,HgO(s) \rightarrow 2\,Hg(l) + O_2(g)$
    (c) $P_4(s) + 6\,Cl_2(g) \rightarrow 4\,PCl_3(l)$
    (d) $2\,KI(aq) + F_2(aq) \rightarrow 2\,KF(aq) + I_2(s)$
    (i) decomposition
    (ii) synthesis
    (iii) single displacement
    (iv) double displacement
    (4.2, 4.4, 4.6) K/U

Write a short answer to each question.

21. Describe four clues that indicate that a chemical reaction has taken place. (4.1) K/U
22. How does a catalyst affect a reaction? (4.1) K/U
23. What must be true for an equation to be balanced? (4.1) K/U
24. How are coefficients different from subscripts in a chemical equation? (4.1) K/U
25. Water and carbon dioxide combine to produce carbonic acid. What makes this change a synthesis reaction? (4.2) K/U
26. Write the general pattern for a decomposition reaction. (4.2) K/U
27. What is required for most decomposition reactions to proceed? (4.2) K/U
28. In what way are the synthesis and decomposition of water opposite reactions? (4.2) K/U
29. How is gasification different from burning garbage? (4.3) K/U
30. Describe one useful product of plasma gasification. (4.3) K/U
31. Describe three concerns about the gasification of landfill garbage. (4.3) K/U
32. What is the activity series of metals? (4.4) K/U
33. How can the relationship among metals in the activity series be used to predict whether or not displacement will occur? (4.4) K/U
34. What does the halogen activity series indicate about the role of fluorine in reactions with compounds of other halogens? (4.4) K/U
35. What are two ways that mercury is different from other metals? (4.5) K/U
36. Compare the general patterns for single and double displacement reactions. (4.4, 4.6) K/U
37. What is the state of a precipitate that results from a precipitation reaction? (4.6) K/U
38. What does the symbol (aq) indicate? (4.6) K/U

Understanding

39. Balance the following chemical equations: (4.1) K/U
    (a) $NaClO_3(s) \rightarrow NaCl(s) + O_2(g)$
    (b) $C_2H_2(g) + O_2(g) \rightarrow H_2O(g) + CO_2(g)$
    (c) $Na(s) + H_2O(l) \rightarrow NaOH(aq) + H_2(g)$
40. Balance the following chemical equations containing polyatomic ions: (4.1) K/U
    (a) $Ca(s) + HNO_3(aq) \rightarrow H_2(g) + Ca(NO_3)_2(aq)$
    (b) $Fe(NO_3)_3(aq) + (NH_4)_2CO_3(aq) \rightarrow Fe_2(CO_3)_3(s) + NH_4NO_3(aq)$
    (c) $AgNO_3(aq) + K_2CrO_4(aq) \rightarrow Ag_2CrO_4(s) + KNO_3(aq)$
41. During cellular respiration, an organism breaks down glucose to release a usable form of energy. Other chemicals are produced in the process. Balance the respiration equation shown below.
    $C_6H_{12}O_6(aq) + O_2(aq) \rightarrow CO_2(aq) + H_2O(l) + \text{energy}$ (4.1) K/U
42. (a) List three clues that a chemical reaction has occurred.
    (b) If you observed one of these clues, would that prove that a chemical reaction has occurred? Explain. (4.1) K/U
43. Classify these reactions as synthesis or decomposition: (4.2) K/U
    (a) $2\,H_2O_2\,(aq) \rightarrow 2\,H_2O(l) + O_2\,(g)$
    (b) $2\,Mg(s) + O_2\,(g) \rightarrow 2\,MgO(s)$
    (c) $2\,Al(s) + 3\,Cl_2\,(g) \rightarrow 2\,AlCl_3\,(s)$
    (d) $H_2CO_3\,(aq) \rightarrow H_2O(l) + CO_2\,(g)$ (4.2) K/U

44. Predict the products of these synthesis or decomposition reactions. Write a balanced chemical equation, without state symbols, to represent each reaction. (4.2) K/U C
    (a) $HgO \rightarrow$
    (b) $Na + O_2 \rightarrow$
    (c) $NH_3 \rightarrow$
45. When heated, silver carbonate decomposes into silver oxide and a gas. (4.2) K/U T/I C
    (a) Identify the gas.
    (b) Write a balanced chemical equation for this reaction.
46. Water can be forced to decompose into hydrogen and oxygen gases. Write the balanced equation for this decomposition reaction. (4.2) K/U
47. Some Canadian municipalities are considering gasification to deal with overcrowded landfills. (4.3) K/U
    (a) Which type of chemical reaction is involved in the process of gasification?
    (b) What conditions are required for plasma gasification?
    (c) Why must the product of plasma gasification be chemically treated before further use?
48. (a) Predict whether zinc will react with lead(II) sulfate.
    (b) What information did you use to make your prediction?
    (c) Name one metal (other than zinc) that, according to your information, would react with lead(II) sulfate.
    (d) Name one metal (other than zinc) that would not react with lead(II) sulfate. (4.4) K/U C
49. (a) Predict whether nickel will react with sodium chloride.
    (b) What information did you use to make your prediction? (4.4) K/U C
50. Use the activity series of metals to write a balanced chemical equation for each reaction that occurs. Write *no reaction* for any that do not. (4.4) K/U C
    (a) $Al(s) + H_2SO_4(aq) \rightarrow$
    (b) $Zn(s) + Cu(NO_3)_2(aq) \rightarrow$
    (c) $Au(s) + KNO_3(aq) \rightarrow$
51. Bubbling chlorine gas through a solution of potassium iodide produces elemental iodine and a solution of potassium chloride. (4.4) K/U C
    (a) Which type of reaction occurs?
    (b) Write a balanced chemical equation to describe the reaction.
52. A student placed a sample of steel wool into sulfuric acid, $H_2SO_4(aq)$. (4.4) K/U C
    (a) Which type of reaction occurs?
    (b) Which gas is produced?
    (c) Write a balanced chemical equation, including state symbols, to describe the reaction.
53. Every spring, mercury disappears from the Arctic atmosphere while the concentration of bromine increases. (4.5) K/U
    (a) What is the source of the bromine?
    (b) What do scientists believe happens to bromide ions in the presence of sunlight?
    (c) How might bromide be related to the disappearance of mercury?
54. Describe the solubility of each of the following compounds: (4.6) K/U
    (a) potassium bromide, $KBr(s)$
    (b) sodium carbonate, $Na_2CO_3(s)$
    (c) silver acetate, $AgC_2H_3O_2(s)$
55. (a) Rewrite the following chemical equation so that it includes the missing state symbols:

    $$CdSO_4(aq) + H_2S(aq) \rightarrow CdS(\;) + H_2SO_4(\;)$$

    (b) Predict any precipitates formed in this reaction.
    (c) Explain how you determined whether there would be any precipitates. (4.6) K/U C
56. A precipitate forms in the reaction between solutions of sodium hydroxide and nickel(II) chloride. (4.6) K/U C
    (a) Write the chemical formula for the precipitate.
    (b) Explain how you determined the identity of the precipitate.
57. Limestone consists mostly of calcium carbonate, $CaCO_3(s)$. One way to test if a mineral is limestone is to put a small sample of it in hydrochloric acid. If bubbles appear, the sample may be limestone. (4.1, 4.6) T/I
    (a) The reaction produces calcium chloride, carbon dioxide, and water. Write the balanced chemical equation, including state symbols, for the reaction.
    (b) What causes the bubbles?
    (c) How useful is this test for limestone? Explain.

58. Nitrous oxide, $N_2O(g)$, can be produced by heating solid ammonium nitrate. Another product of this reaction is water vapour. (4.2, 4.4, 4.6) T/I
   (a) Write the balanced chemical equation for the reaction that produces nitrous oxide.
   (b) What type of reaction is this?
   (c) This reaction can be dangerously explosive. Another way to produce nitrous oxide and liquid water is by mixing ammonia and oxygen gases. Write the balanced chemical equation for this alternative process.
   (d) What type of reaction is described in part (c)?
59. Hydrogen peroxide, $H_2O_2(l)$, can be produced by reacting oxygen gas and hydrogen gas. (4.2, 4.4, 4.6) T/I A
   (a) Write the balanced chemical equation for this process.
   (b) What type of reaction is this?
   (c) The reaction depends on the use of a catalyst. Describe the purpose of a catalyst.
   (d) What would be the economic effect if a hydrogen peroxide processing plant did not have the correct catalyst?
60. Predict which type of reaction is likely to take place between the following pairs of reactants. (4.2, 4.4, 4.6) K/U
   (a) calcium metal and oxygen gas
   (b) hydrocyanic acid and sodium hydroxide solution
61. Classify the reactions represented by the following equations. Explain how you arrived at your answers. (4.2, 4.4, 4.6) K/U C
   (a) $HF(aq) + KOH(aq) \rightarrow KF(aq) + H_2O(l)$
   (b) $2\ Al(s) + Cr_2O_3(s) \rightarrow Al_2O_3(s) + 2\ Cr(s)$
62. Complete and balance each of the following reaction equations and classify the reaction, giving reasons for your answer: (4.2, 4.4, 4.6) K/U C
   (a) $AlCl_3(aq) + AgNO_3(aq) \rightarrow AgCl(s) + Al(NO_3)_3(aq)$
   (b) $Li(s) + H_2SO_4(aq) \rightarrow Li_2SO_4(aq) + H_2(g)$
   (c) $H_2SO_4(aq) + NaOH(aq) \rightarrow H_2O(l) + Na_2SO_4(aq)$
63. An aqueous solution of lead(II) nitrate is mixed with aqueous sodium iodide. (4.6) T/I C
   (a) Predict whether these two compounds will react.
   (b) Describe how you made your prediction.
   (c) If you predict that a reaction will occur, write the balanced chemical reaction for the reaction. Include the states of all reactants and products.
   (d) List the precipitates, if any, that form in this reaction.

Analysis and Application

64. A chemist wants to determine whether a certain metal will act as a catalyst in a reaction between zinc and hydrochloric acid. T/I (4.1)
   (a) Write the balanced equation for the reaction between zinc and hydrochloric acid.
   (b) Write a design outlining an investigation to determine whether the unknown metal is a catalyst for this reaction.
65. Two aqueous solutions, hydrogen iodide and hydrogen peroxide, are mixed. It is unknown whether they give off a gaseous product when they react. Design an investigation to test for the release of a gas using the law of conservation of mass. (4.1) T/I A
66. Lead(II) nitrate reacts with potassium iodide in a double displacement reaction. (4.1) T/I C
   (a) Write the chemical formula for each reactant and product in this reaction.
   (b) Write the balanced equation for this reaction.
   (c) How could you use this reaction to demonstrate the law of conservation of mass? Write a design for the investigation, including any necessary safety precautions.
67. Mercury(II) oxide decomposes when heated in a Bunsen burner flame. (4.2) K/U A
   (a) Write the chemical formula for mercury(II) oxide.
   (b) Write the chemical formulas of the products of the decomposition reaction.
   (c) Write the balanced equation for the decomposition of mercury(II) oxide.
68. Hydrogen and bromine are both diatomic elements. (4.2) T/I
   (a) Write the formulas for both of these substances.
   (b) Predict the product of the synthesis reaction between them.
   (c) Write the balanced equation for the synthesis reaction between bromine and hydrogen.
69. A piece of sodium is added to a solution of potassium chloride. (4.4) T/I
   (a) Identify the products of this reaction.
   (b) Write the balanced equation for this reaction.
70. A chemist wants to know whether zinc reacts with sulfuric acid. (4.4) T/I
   (a) How could the chemist find the answer to this question without performing the test in the laboratory?
   (b) Predict whether zinc reacts with sulfuric acid.

71. The steel hull of a ship is made mostly of iron. Attaching a zinc plate to a ship's hull protects the hull by preventing iron from reacting with water. The zinc plate, however, reacts with water and must be replaced periodically. (4.4) A C
   (a) Use the activity series to suggest a possible explanation for why zinc "protects" the iron in the hull of the ship.
   (b) Would a silver plate attached to the ship's hull have the same effect? Explain.
   (c) Name one other metal that would be effective in slowing corrosion and one that would not. Explain your choices.
72. A researcher is studying a reaction in which the reactants are solutions of lithium chloride and silver nitrate. (4.6) A
   (a) Write the balanced equation for the reaction that occurs when solutions of lithium chloride and silver nitrate are combined. Include all state symbols.
   (b) Describe a chemical test that could be done to be certain that all of the silver ions present had precipitated. Why does your test work?
73. A student combines the following pairs of aqueous solutions. In each case, predict whether a reaction will occur and, if you predict that a reaction will occur, write the balanced equation for this reaction, including all state symbols. (4.6) T/I
   (a) $BaCl_2(aq)$ and $Na_2SO_4(aq)$
   (b) barium nitrate and potassium carbonate
   (c) magnesium chloride and sodium hydroxide
   (d) iron(II) nitrate and potassium sulfate
74. A chemist performs a laboratory test in which hydrochloric acid is mixed with magnesium hydroxide. (4.2, 4.4, 4.6) T/I
   (a) Write the balanced chemical equation for this reaction.
   (b) Which type of reaction is this?
75. Silicon carbide, SiC(s), is made by reacting solid silicon dioxide with solid carbon at very high temperatures. The other product of the reaction is carbon monoxide gas. (4.2, 4.4, 4.6) T/I
   (a) Write the balanced chemical equation for this reaction.
   (b) Which type of reaction is this?
76. Write the balanced equation for the neutralization reaction between nitric acid and potassium hydroxide. (4.6) T/I
77. Hydrofluoric acid is used to etch glass. Used hydrofluoric acid must be neutralized before disposal. (4.6) A
   (a) Name one compound that could be used to neutralize the acid.
   (b) Write the balanced equation for the neutralization reaction.

Evaluation

78. A chemical company has a large volume of waste silver nitrate solution. In order to reduce costs, they want to recover the silver from the solution as silver metal. One proposed recovery method is to add solid aluminum to the solution. (4.2, 4.4, 4.6) T/I C A
   (a) Using what you have learned about chemical reactions, evaluate this proposal.
   (b) If you think the proposal would work, write the balanced chemical equation for the reaction. If you think it would not work, suggest a different chemical reaction to retrieve the silver metal.
79. A water supply has been contaminated with an unknown mixture of toxic metal solutions leaking from a nearby factory. One way to treat the problem is to precipitate the metals from the water. Environmental chemists are considering using sodium carbonate and sodium sulfate for this purpose. Based on the solubility table, which compound would you recommend? (4.2, 4.4, 4.6) T/I A
80. The atmosphere of the planet Venus contains a great deal of sulfuric acid. Sulfuric acid falls as precipitation in much the same way as water falls on Earth. An engineer is designing a probe to send through the atmosphere of Venus. The engineer plans to use a thick plating of copper as the outer skin of the probe to protect it from the corrosive effects of the sulfuric acid. (4.2, 4.4, 4.6) T/I A
   (a) Given what you have learned about reactivity, evaluate this proposal.
   (b) List two alternative plating metals for this probe and explain the effectiveness of each of them.

Reflect on Your Learning

81. What did you learn in this chapter that you found most surprising? Explain. C

82. (a) What questions do you still have about chemical reactions?
    (b) What methods can you use to find answers to your questions? K/U C A

83. Has the information you have learned in this chapter changed your opinion about any environmental issues related to chemistry? Explain. C A

Research

GO TO NELSON SCIENCE

84. News stories often report on advances in hydrogen fuel cell technology for transportation. Research how hydrogen fuel cells are used to power vehicles. In your research, find out what problems remain to be solved for this technology. Create a chart that compares the advantages and disadvantages of using this technology. Then write a paragraph that explains your position on the question: Do you think hydrogen-powered cars are a useful technology for society to pursue? Support your position with evidence from your research. T/I C A

85. Gasification is not necessarily a new technology. The process was originally developed in the 1800s. Research the history and development of gasification processes. Find out about the different forms of gasification and determine where gasification is being used throughout Canada and other countries. Decide whether or not you think additional funds should be directed toward the further development of gasification processes. Support your position with evidence from your research. T/I C A

86. Mercury is a powerful neurotoxin. This makes its unexplained disappearance from Arctic air a concern. Research the properties of mercury and how it affects the human nervous system. Create a diagram of a food chain suggesting how mercury might find its way into the human body. Use what you learn to argue for or against continued research into the disappearance of mercury in the Arctic ecosystem. In addition to mercury, low-level ozone disappears from the Arctic air at the same time as mercury. Research the properties and dangers of low-level ozone, and use what you learn to suggest why its disappearance is also a concern. T/I C A

87. Acid precipitation damages both the natural environment and human-made structures. Research the causes of acid precipitation and what is being done to mitigate its effects. In your research, determine how human activity has affected acid precipitation. Prepare to present your response to the question: What should be done about acid precipitation? Include reasoned arguments to support your suggestions. T/I C A

88. Fritz Haber helped develop the Haber process for producing large quantities of ammonia. Research the Haber process to find out about the chemical reaction it uses. Write a short report that explains the importance of this process to our society and addresses the following questions:
    (a) What products would not be possible without the Haber process?
    (b) Do you think the Haber process has helped or hurt society? Explain your reasoning. T/I C A

CHAPTER 5 Chemical Processes

How Can We Minimize the Impact of Chemical Processes on the Environment?

KEY CONCEPTS

After completing this chapter you will be able to

- analyze the impacts of industrial processes on the environment and human health, and consider how we can minimize these impacts
- understand the chemical reactions involved in industrial processes
- assess how effective chemical reactions are in addressing social and environmental needs
- write balanced chemical equations to describe chemical processes, including combustion
- design an investigation to demonstrate the difference between complete and incomplete combustion
- plan and conduct an inquiry to compare solutions of non-metallic oxides and metallic oxides
- investigate neutralization reactions and the safest ways to neutralize an acid or a base

Have you ever found yourself mesmerized, staring at the soothing light of a candle flame? Candles have been around for centuries, so you would think that scientists would know exactly what is happening when a candle burns. Surprisingly, the chemistry of this common process is quite complex and not completely understood. For example, you may have noticed that the colour of a candle flame is not uniform. The base of the flame is almost invisible. Temperatures in this region reach only about 600 °C—relatively cool for a flame. Not much combustion occurs here. Instead, molecules of liquid wax vaporize and decompose into hydrogen gas and molecules made of carbon and other elements. These molecules clump together to form small visible particles of soot. Rising soot particles pass into the larger yellow region of the flame, where temperatures reach 1200 °C. This is hot enough to make the soot particles glow brightly like an incandescent light bulb before they burn. Some candle flames have a thin, blue, very hot outer layer. The temperature here—up to 1400 °C—is high enough to cause gas molecules in the flame to lose electrons and turn into ions.

The wick of the candle is made of an absorbent material that draws liquid wax upward, like a paper towel drawing up water. The size of the wick is one factor that controls how much soot a candle produces. A wick that is too thick may draw more wax than can be burned completely. The result is a sooty, polluting flame.

The soot from a candle is a minor inconvenience. The soot from many candles, however, can pose a health concern. As a result of burning candles and poor air circulation, the air quality in some enclosed places can be worse than in downtown Toronto during rush hour.

The burning of wax in a candle is an example of a chemical process. Several steps are required to process the starting materials (wax and oxygen) into products (soot, combustion gases, and energy). Chemical reactions are important parts of many industrial processes, such as making steel, paper, and consumer products. These processes all convert starting materials into finished products. In this chapter, you will examine the operation of these processes. As you learn about the processes, you will consider their effect on the environment and how this effect can be minimized.

STARTING POINTS

Answer the following questions using your current knowledge. You will have a chance to revisit these questions later, applying concepts and skills from the chapter.

1. What effect do you think the availability of oxygen may have on how a fuel burns?
2. Suppose that you are choosing a chemical to neutralize a spill of highly corrosive acid. What do you think are the most important characteristics of this chemical?
3. (a) What is meant when we describe a compound as being a non-metallic oxide? Give a possible example.
   (b) How do you think this compound could be made in the laboratory?
4. What are the potential effects of non-metallic oxides on human health and the environment?
5. How do industries reduce the environmental effect of processes that use chemicals?
6. Do you think it is possible for a company that manufactures or uses chemicals to operate in a green and sustainable way? Why?

Mini Investigation

The Burning Candle Riddle

Skills: Predicting, Performing, Observing, Analyzing, Communicating SKILLS HANDBOOK A1.2

In this investigation, you will observe how two candles respond differently to changes in their environment.

Equipment and Materials: chemical safety goggles; lab apron; Petri dish; gas jar; two birthday candles, one short and one tall

This activity involves flames. Tie back long hair and secure loose clothing.

1. Put on your chemical safety goggles and a lab apron.
2. Your teacher will light the taller candle for you. Allow two drops of wax to drip from the candle into the Petri dish. Stand the candle upright in one of the drops of wax.
3. Light the shorter candle from the first one and stand it upright in the other drop of wax (**Figure 1**).
4. Predict which candle will go out first when the jar is lowered, upside down, over them both.
5. Test your prediction by lowering the gas jar over the candles.

A. Compare your observations with those of your classmates. T/I

B. Suggest an explanation for your observations. T/I

Figure 1 Which candle goes out first?

5.1 The Combustion of Hydrocarbons

In January 2005, 120 people who had attended a hockey game at a Vancouver arena were treated in hospital for carbon monoxide poisoning. The source of the carbon monoxide was a faulty propane-powered ice-resurfacing machine.

Carbon monoxide is a toxic gas that can be produced during the combustion of a carbon-based fuel such as propane. Initial symptoms of carbon monoxide poisoning include drowsiness, impaired judgment, and unconsciousness. Prolonged exposure can be fatal. Fortunately, everyone survived the Vancouver incident.

combustion a chemical reaction in which a fuel burns in oxygen to produce combustion products and, often, a flame

greenhouse gas any of several gases that allow solar radiation to pass through the atmosphere but absorb radiation emitted by Earth, thereby trapping thermal energy and making Earth warmer

organic compound a molecular compound that contains one or more carbon–carbon bonds and, often, one or more carbon–hydrogen bonds

air pollution chemicals and particles in the atmosphere that harm living organisms or damage the environment

Combustion is a chemical reaction in which a fuel burns in oxygen, usually from air. Coal, propane, and gasoline are common fuels (**Figure 1**). The molecules of fuels—like all compounds—are a source of chemical energy. Combustion releases this chemical energy as thermal energy and light. The other products of combustion vary depending on the type of fuel and how it was burned. Common products are carbon dioxide, water, and sometimes carbon monoxide. Carbon dioxide is a **greenhouse gas** that contributes to climate change, as does water vapour.

Propane is an example of an **organic compound:** a molecular compound containing one or more carbon–carbon bonds. Many organic compounds also contain one or more carbon–hydrogen bonds. The majority of fuels used today, including gasoline, are composed of organic compounds.

Many fuels contain impurities such as sulfur compounds and heavy metals. Combustion releases these substances into the atmosphere as **air pollution,** unless they are removed by special pollution-control technologies.

Figure 1 Coal is one of Earth's most abundant fuels. More electricity is generated globally by the combustion of coal than by any other energy source. However, the combustion of coal is a major source of greenhouse gas emissions, soot, mercury, and other pollutants.

Investigation **5.1.1**

Combustion of Ethyne (p. 228)
In this investigation, you will test three different mixtures of ethyne and air to explore their combustion characteristics.

Many substances burn. However, this section deals with the combustion of a special group of organic compounds called hydrocarbons. As their name suggests, hydrocarbons are made up of only two elements: hydrogen and carbon. Propane, C_3H_8, and butane, C_4H_{10}, are common examples of hydrocarbons. Gasoline is a mixture of mostly liquid hydrocarbons. Most hydrocarbons are extracted from coal, oil, and natural gas. Although some hydrocarbons are used in the synthesis of other compounds, most are burned for their energy content. Much of this energy is used for transportation. Other important uses of this energy are the generation of electricity and home heating.

Types of Hydrocarbon Combustion

All hydrocarbons burn, provided they are hot enough and have sufficient oxygen. We classify the combustion of hydrocarbons as either complete or incomplete, depending on the products that form.

Complete Combustion

The **complete combustion of a hydrocarbon** occurs when the oxygen supply is plentiful. The products are energy, carbon dioxide, and water vapour. The general unbalanced chemical equation for the combustion of a hydrocarbon, C_xH_y, is

$$C_xH_y + O_2(g) \rightarrow CO_2(g) + H_2O(g) + \text{energy}$$

This general equation applies to all hydrocarbons. For example, natural gas is mostly methane, CH_4. The chemical equation for the complete combustion of methane is

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g) + \text{energy}$$

Complete combustion usually occurs under carefully controlled conditions in which excess (more than enough) oxygen is present. Complete combustion is the ideal way to burn a fuel because it releases the most energy from the fuel molecules. As a result, complete combustion produces the cleanest (least sooty) and hottest flames. The blue flame of a gas stove is evidence of complete combustion (**Figure 2**). Keeping a natural gas furnace well maintained ensures that the gas is burned as completely as possible and produces the most thermal energy. This helps to reduce the cost of home heating.

Ethyne, C_2H_2, is another well-known hydrocarbon fuel. It is more commonly known as acetylene. When ethyne burns with a good supply of oxygen, it generates the high temperatures needed to cut steel. An oxyacetylene torch is supplied by two gas cylinders: one containing acetylene and the other containing oxygen. The steelworker controls the temperature of the flame by adjusting the flow of oxygen to the torch nozzle. As more oxygen reaches the flame, combustion of acetylene becomes more complete. This results in a hotter flame. The temperature of an oxyacetylene torch flame can reach 3500 °C—hot enough to slice through steel like a knife through butter (**Figure 3**). The chemical equation for the complete combustion of ethyne (acetylene) is

$$2\ C_2H_2(g) + 5\ O_2(g) \rightarrow 4\ CO_2(g) + 2\ H_2O(g) + \text{energy}$$

complete combustion of a hydrocarbon the combustion (burning) of a hydrocarbon in a plentiful supply of oxygen to produce carbon dioxide, water, and energy

LEARNING TIP

Complete Combustion Reactions
When writing balanced chemical equations for complete combustion reactions, start by balancing the C atoms, then O, then H. If you have to add a coefficient in front of the hydrocarbon formula, you will have to start again, balancing C, O, and H.

Figure 2 The blue colour of a natural gas flame indicates that complete combustion is occurring.

Figure 3 The two pipes in an oxyacetylene torch carry oxygen and the fuel from storage tanks to the nozzle.

Incomplete Combustion

Incomplete combustion of hydrocarbons occurs when the supply of oxygen is limited. This situation can occur while burning a hydrocarbon in an enclosed space, such as when operating a car or barbeque in a garage. Incomplete combustion often occurs in uncontrolled or less than ideal conditions. The mixture of air and fuel that results

incomplete combustion of a hydrocarbon the combustion (burning) of a hydrocarbon in a limited supply of oxygen; products may include carbon dioxide, carbon monoxide, soot, water, and energy

Figure 4 The combustion of excess hydrocarbons at an oil refinery produces yellow flames and soot. These are both evidence of incomplete combustion.

Figure 5 A cold gasoline-powered mower is easier to start if the engine is initially adjusted to favour incomplete combustion.

in incomplete combustion is said to be "fuel rich." A log in a wood-burning stove is an example of a fuel-rich condition: there is lots of fuel (wood) and limited oxygen. The lack of oxygen prevents "ideal" or complete combustion from occurring, limiting the quantity of energy released.

Flames resulting from incomplete combustion are often sooty, yellow, and cooler than flames from complete combustion (**Figure 4**).

Incomplete combustion is usually undesirable. However, it can be useful in some applications. For example, incomplete combustion is used to start gasoline-powered devices such as lawn mowers (**Figure 5**). A cold lawn mower is much easier to start if a fuel-rich mixture enters its engine than if there is less fuel and more air. The volume of air entering the engine is controlled by a valve on the lawn mower called the choke. Before starting a cold lawn mower, adjust the choke so that it is in the closed position. This restricts the air flow into the engine. The result is a fuel-rich mixture that is easier to ignite. Once the engine starts, you might see sooty exhaust produced. This is evidence of incomplete combustion. After warming up the lawn mower for a few seconds, open the choke, allowing more air to mix with the fuel. The result is combustion that is more complete.

Incomplete combustion produces a greater variety of combustion products than complete combustion. Carbon dioxide, water, and energy can be produced. However, incomplete combustion also releases either carbon monoxide or soot, or both. Soot is a solid black ash observed when a carbon-based fuel burns. Soot is actually a mixture of solid carbon-rich molecules. This is why soot is often represented in chemical equations as C(s). Because so many reaction products are possible, incomplete combustion cannot be represented by a single chemical equation. For example, both of the following chemical equations represent the incomplete combustion of heptane, a hydrocarbon in gasoline:

$$C_7H_{16}(l) + 7\ O_2(g) \rightarrow 3\ C(s) + 2\ CO(g) + 2\ CO_2(g) + 8\ H_2O(g) + \text{energy}$$

$$2\ C_7H_{16}(l) + 11\ O_2(g) \rightarrow 14\ CO(g) + 8\ H_2O(g) + \text{energy}$$

By comparison, there is only one chemical equation for the complete combustion of heptane. This reaction does not produce either soot or carbon monoxide:

$$C_7H_{16}(l) + 11\ O_2(g) \rightarrow 7\ CO_2(g) + 8\ H_2O(g) + \text{energy}$$

Note that more molecules of oxygen are required per molecule of heptane for complete combustion to occur.

CONCERNS RELATED TO INCOMPLETE COMBUSTION

Incomplete combustion of hydrocarbons can be a problem in enclosed places where there is poor ventilation. Even the candles burning in places of worship can pose a number of health concerns. Incomplete combustion in such places, and in gas-powered engines, is undesirable for a variety of reasons.

1. Incomplete combustion releases only a portion of the energy that may be obtained from hydrocarbon fuels such as gasoline. Automobile engines rarely burn gasoline completely. However, the combustion of gasoline is more complete while travelling at highway speeds than when driving in the city. The most incomplete combustion of gasoline occurs while a vehicle is idling. This is because an idling engine operates below its optimal temperature. As a result, gasoline is burned less completely than when the engine operates at higher temperatures. Also, fuel residue may form on engine components, making them work less efficiently.
2. Soot particles from incomplete combustion are an inhalation hazard. Many of the chemicals in soot are toxic. Furthermore, soot particles are often too small to be filtered by the upper respiratory tract. As a result, these particles may penetrate deep into the lungs and irritate sensitive tissue. This can lead to respiratory problems, including asthma. Soot can also foul moving engine parts, making the engine even less efficient.

3. Carbon monoxide produced during incomplete combustion is also an inhalation hazard. This toxic gas is often called the "silent killer" because it has no colour, taste, or odour. Carbon monoxide is dangerous because it binds tightly with the hemoglobin in blood. This prevents hemoglobin from binding with inhaled oxygen. During a rush-hour traffic jam, the carbon monoxide concentration in the air can be high enough to cause headaches—a symptom of mild carbon monoxide poisoning. Carbon monoxide poisoning can also occur in the home. According to the Canadian Safety Council, carbon monoxide is the leading cause of fatal poisonings in North America. Each year, several Canadians die needlessly due to faulty home heating equipment. Many of these deaths could have been prevented if a carbon monoxide detector had been installed.

WEB LINK

To read the Canadian Safety Council's information on carbon monoxide,

GO TO NELSON SCIENCE

Mini Investigation

Exploring Bunsen Burner Combustion

Skills: Performing, Observing, Communicating

SKILLS HANDBOOK A1.2, A2.2, A3.1

The operation of a Bunsen burner illustrates the role of oxygen in determining the type of combustion that occurs. If you limit the oxygen supply that reaches the flame, incomplete combustion will occur. Maximizing the oxygen supply results in complete combustion.

Equipment and Materials: chemical safety goggles; Bunsen burner clamped to a retort stand; spark lighter; 2 clean metal spoons; tongs

This activity involves flames. Tie back long hair and secure loose clothing.

1. Put on your chemical safety goggles.
2. Light your Bunsen burner with the spark lighter (**Figure 6**). Adjust the flame so that it shows yellow/orange.
3. Use tongs to hold a clean metal spoon in the flame for about 30 s. Record what you see.
4. Carefully cool the spoon in a stream of tap water.
5. Adjust the burner so that the flame becomes blue.
6. Repeat Steps 3 and 4 using the blue flame and a clean spoon.
7. Dispose of your materials and clean up your workstation as directed. Wash your hands.

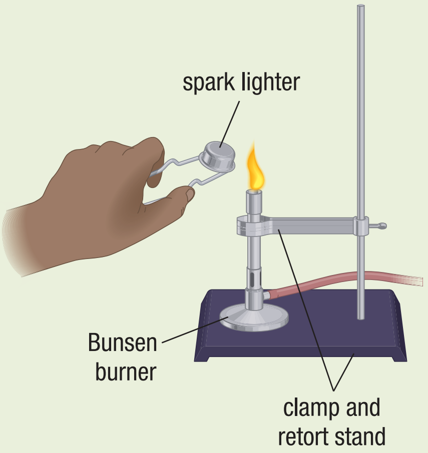


Figure 6 Lighting the Bunsen burner

A. What evidence of incomplete combustion did you observe? T/I
B. What evidence of complete combustion did you observe? T/I
C. Explain how the adjustments you made to the Bunsen burner changed the characteristics of the flame. T/I
D. Which type of combustion is more desirable from the perspective of a car owner? Support your answer with at least two reasons. A

OTHER COMBUSTION HAZARDS

Firefighting has always been a risky profession. However, the increased use of synthetic materials in recent decades has made the job even more hazardous. Many veteran firefighters describe domestic fires as being hotter and more toxic than ever before.

Synthetic materials have unusual combustion properties. As synthetic materials burn, they can produce toxic gases and particulate matter. Firefighters wear sealed breathing apparatus to avoid inhaling these toxins. However, studies have shown that some of the chemicals may be absorbed through the skin. Not surprisingly, fire departments in many countries are seeing a rise in the incidence of certain types of cancer among their workers and retirees.

CAREER LINK

Firefighting has a very long and honourable history. To find out about this career, and the training necessary to enter it,

GO TO NELSON SCIENCE

Even natural substances can produce potentially hazardous chemicals when they burn. For example, conservation areas routinely conduct controlled burns of vegetation. Before they do, notices are often sent to nearby residents warning people with asthma to avoid prolonged exposure to the smoke and that the smoke may carry small droplets of poison ivy oil.

Forest fires can affect air quality on a much larger scale. In a recent study of the effect of forest fires on air quality, scientists found that burning pine forests release large quantities of chemicals called alkaloids into the air. Alkaloids are naturally occurring substances produced by plants. Some, like morphine, are beneficial to humans. In small doses, pine tree alkaloids are not a risk. However, in larger doses, the alkaloids produced from burning pine trees can be potent toxins. As this research continues, it could change our perceptions of the health risks faced by communities downwind from a forest fire.

WEB LINK

To find out more about the possible health effects of forest fires,

GO TO NELSON SCIENCE

Research This

Clean Coal—A Contradiction?

Skills: Researching, Analyzing, Evaluating, Communicating, Defending a Decision

SKILLS HANDBOOK A5.1

Coal is one of the most abundant fuels on Earth. It is also the dirtiest. The combustion of coal is a major source of sulfur oxides, one of the pollutants that causes acid precipitation. In addition, burning coal produces more carbon dioxide per unit of energy than any other fuel. Coal is relatively cheap and readily available, so many countries rely on it to power their economic growth. For example, about half of the electricity in the United States is generated by burning coal. Without a cheaper alternative, the use of coal is expected to increase.

Recent advances in coal-processing technology have significantly cleaned up the industry (**Figure 7**). Some stakeholders claim that "clean" coal-powered electricity generation is now a reality. As a result, millions of dollars have been invested to further develop this technology.

1. Research how "clean coal" technology works.
2. Explore two opposing expert opinions on "clean coal" technology. You might choose to focus on the health and safety of local populations, for example.

A. Evaluate the two opinions. Prepare a two-paragraph summary that outlines the strengths and weaknesses of the technology. T/I C

B. In your opinion, is further investment in this technology warranted? Why? T/I

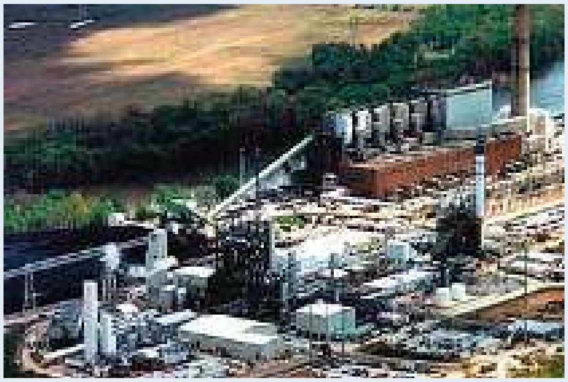


Figure 7 Coal gasification plant

GO TO NELSON SCIENCE

5.1 Summary

- Hydrocarbons are molecular compounds of the elements carbon and hydrogen.
- The combustion of hydrocarbons can be either complete or incomplete, depending on the availability of oxygen.
- Complete combustion occurs when oxygen is plentiful. It releases the most energy and produces the cleanest flames. The general equation for the complete combustion of a hydrocarbon is
  $C_xH_y + O_2(g) \rightarrow CO_2(g) + H_2O(g) + \text{energy}$
- Incomplete combustion occurs when the oxygen supply is limited. The products of incomplete combustion include soot (mostly carbon), carbon monoxide, carbon dioxide, and water.
- Incomplete combustion cannot be represented by a single balanced chemical equation.

5.1 Questions

1. Explain, in your own words, the difference between a complete combustion reaction and an incomplete combustion reaction involving a hydrocarbon fuel. K/U
2. Cylinders of butane, C_4H_{10}, are used in some types of camping stoves. The butane in these cylinders is liquid because it is under pressure. When the stove is turned on, butane leaves the cylinder and quickly evaporates before it ignites. K/U C A
   (a) Write the chemical equation for the complete combustion of butane.
   (b) Why is it not advisable to use a butane stove inside a tent?
3. Use the following information to determine the chemical formula of the hydrocarbon in each case: T/I
   (a) $C_xH_y + 8\ O_2(g) \rightarrow 5\ CO_2(g) + 6\ H_2O(g)$
   (b) $C_xH_y + 14\ O_2(g) \rightarrow 9\ CO_2(g) + 10\ H_2O(g)$
   (c) $2\ C_xH_y + 19\ O_2(g) \rightarrow 12\ CO_2(g) + 14\ H_2O(g)$
4. Hexane, C_6H_{14}, is a component of gasoline. Write a balanced chemical equation for
   (a) the complete combustion of hexane
   (b) the incomplete combustion of hexane, assuming that the two oxides of carbon are present in equal proportions and that no soot is produced K/U C
5. What evidence of incomplete combustion would each of the following devices show? K/U A
   (a) a Bunsen burner
   (b) automobile exhaust tailpipe
   (c) a gas fireplace
6. Suggest an explanation for each of the following events: K/U A
   (a) Cottagers died in their sleep while heating their cottage using portable propane heaters designed for use on patios.
   (b) A cloud of black smoke belches from the exhaust stack of a truck as the truck starts to move.
   (c) A Bunsen burner flame changes from yellow to blue as the burner is adjusted.
   (d) A candle flame produces more smoke as an inverted jar is lowered over the candle.
7. The air/fuel mixture in a combustion reaction can be described as being either "fuel-lean" or "fuel-rich." K/U A
   (a) Compare the relative quantities of air and fuel in lean and rich mixtures.
   (b) Which mixture is more likely to undergo complete combustion? Which describes incomplete combustion? Why?
8. Most municipalities ban the practice of burning household garbage in backyard bonfires, even in rural areas. Give two possible reasons for this ban. K/U A
9. Most of the hydrocarbon fuels used in transportation are derived from oil. The price of oil could rise from about $85 per barrel to $150–$200 per barrel as supplies of easily accessible oil are used up. Predict the effect this may have on each of the following things. Justify your prediction in each case. A
   (a) your purchasing power
   (b) the design of new neighbourhoods
   (c) the availability of locally produced goods, compared to the availability of goods produced in other countries where labour is cheaper
10. In November 1984, a controversial project to clean the frescoes on the ceiling of the Sistine Chapel in Rome began. The frescoes were painted by the Italian painter Michelangelo (**Figure 8**). Research what had caused these masterpieces to darken. Also, find out why this project was so controversial. Prepare a brief illustrated report on your research. T/I A

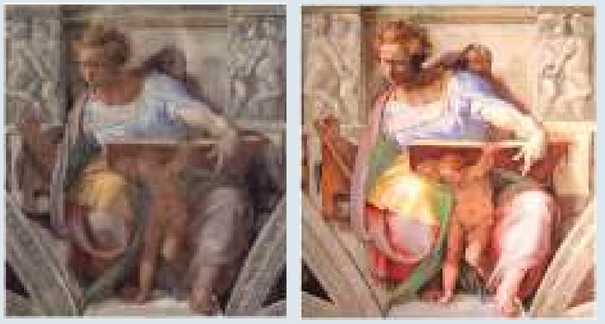


Figure 8 Part of this fresco in the Sistine Chapel has been cleaned.

11. An idling car produces a great deal of unnecessary pollution, as anyone who has worked at a drive-through window can attest. Research the facts about idling and determine which of the following statements are true and which are myths: T/I A
    (a) It is necessary to warm up a car for a few minutes in the winter before it is driven.
    (b) Restarting the car frequently is hard on the engine and uses more gas.
    (c) A block heater is a more efficient way of warming up a car than idling.
    (d) Idling is only a problem in the winter.
    (e) Idling for more than 10 s uses more fuel than restarting the engine.
    (f) An idling engine is not operating at peak efficiency and burns fuel incompletely.
12. Research the link between cancer and chemical exposure. Why is it difficult to say for certain that a specific chemical causes a certain type of cancer? T/I A

GO TO NELSON SCIENCE

The Role of Chance in Discovery

SKILLS HANDBOOK A4

ABSTRACT

Not all important discoveries are the result of carefully designed experiments. Some happened by chance or accident. Four substances that were invented or discovered accidentally are Teflon (the non-stick coating on cookware), Scotchgard (a fabric treatment), Super Glue, and the element iodine.

Introduction

Most scientific discoveries are the result of years of research and experimentation (**Figure 1**). Scientists conduct carefully designed experiments to test hypotheses or to see if one variable influences another. They plan procedures and combine reactants, usually with a fairly good idea of what will happen.

Figure 1 There are generally few surprises in routine chemical research.

Unexpected observations in an experiment usually mean that something went wrong. When this happens, scientists usually dismiss the observations as "experimental error" or "bad luck" and start over. However, the history of chemistry contains several examples of unexpected results leading to important discoveries. Fortunately, the chemists involved had the insight to recognize the significance of what they had stumbled upon.

Teflon

In the late 1930s, Roy J. Plunkett was a research chemist with the DuPont chemical company of New Jersey. DuPont had developed a new refrigerant gas called tetrafluoroethylene (TFE). Plunkett had been studying chemical reactions involving TFE.

One day, Plunkett's lab assistant, Jack Rebok, opened the valve of a pressurized steel gas cylinder that had been freshly filled with TFE the day before. Much to the researchers' surprise, no gas came out. What was even more puzzling was that the cylinder seemed heavier than an empty cylinder should be. They cut open the cylinder to see what went wrong. Inside they found a mysterious white powder. Plunkett reasoned that the TFE gas had undergone a chemical reaction to produce this powder.

Intrigued by this new material, Plunkett immediately transferred his attention to the white powder and its properties. He found that the substance was chemically inert (unreactive), had a high melting point, and was more slippery than any other solid known at the time. Eventually, Plunkett determined that the gas inside the cylinder had undergone a chemical reaction called polymerization: many TFE molecules had joined together to become extremely long molecules. Soon he figured out how to reproduce this reaction in the laboratory. Plunkett had discovered Teflon, the slippery, non-stick material that is now coated onto a host of products, including cookware.

There is more to this unique substance than its non-stick surface, however. The fact that Teflon® is so chemically inert and flexible makes it useful for insulating telecommunications cables. Teflon-coated fibreglass is even used in architecture. Plunkett and Rebok probably never imagined that their "solidified gas" would be so versatile and useful (**Figure 2**).

Figure 2 (a) The roof of the Millennium Dome in London, England and (b) the non-stick surface of cookware are both coated with Teflon.

Scotchgard

In 1952, Patsy Sherman was one of the few female professional chemists working for a major chemical company. Together with Sam Smith, a colleague at 3M, Sherman was involved in a project to develop a new type of synthetic rubber. This product would be used for jet aircraft fuel lines. One day, Sherman and Smith saw a lab assistant accidentally drop a bottle containing liquid synthetic latex that Sherman had just made. Some of the fluid splashed on the assistant's shoes. Sherman and Smith were surprised to find that the liquid could not be washed off. Solvents also could not remove it, and it resisted being soiled. The chemical that Sherman and Smith had invented eventually came to be the key ingredient in Scotchgard, a stain-resistant fabric treatment that is still in use today.

Super Glue

In 1942, while working for Kodak Research Laboratories, Harry Coover was exploring the properties of chemicals called cyanoacrylates. He was hoping to develop a clear plastic to make gun sights. Coover quickly found that cyanoacrylates were too sticky to work with, and gave up on them. Later, in 1951, Coover decided to try these materials in a different application. He was developing a new heat-resistant plastic for aircraft. This time Coover realized that the "super stickiness" of cyanoacrylates had huge potential. As a result, the first cyanoacrylate adhesive was sold in stores in 1958. Super Glue is now used in areas as diverse as carpentry and surgery (**Figure 3**).

Figure 3 Surgeons use cyanoacrylate adhesive to repair bones cut during major surgery.

Iodine

The story of the discovery of iodine begins at the start of the nineteenth century. France was at war with England. A British naval blockade prevented the French from receiving many imports, including the ingredients to produce gunpowder. The French chemist Bernard Courtois (1777–1838) developed a new process that required large quantities of sodium carbonate. Normally, sodium carbonate was produced from wood ashes, but wood was scarce. Courtois did, however, have an ample supply of seaweed, so he tried to obtain sodium carbonate from the ashes of burned seaweed.

One day, while cleaning seaweed pans with hot concentrated sulfuric acid, Courtois noticed a violet-coloured vapour rising above the pans. When the rising vapour touched a cool metal surface it solidified, forming dark purple crystals (**Figure 4**).

Figure 4 Iodine is a violet-black solid that gives off corrosive iodine vapour when heated. Iodine was originally extracted from seaweed.

Courtois was too involved in the war to publish his discovery. Instead, he shared samples of the purple solid with a number of prominent chemists for further study. Joseph Gay-Lussac and Humphrey Davy both announced that the substance was a new element with properties similar to those of chlorine.

Further Reading

Coffey, Patrick. (2008). *Cathedrals of Science: The Personalities and Rivalries That Made Modern Chemistry*. New York: Oxford University Press.

Levere, Trevor H. (2001). *Transforming Matter: A History of Chemistry from Alchemy to the Buckyball*. Baltimore: The Johns Hopkins University Press.

GO TO NELSON SCIENCE

5.2 Questions

1. What characteristics do you think a research scientist should have? Why? A
2. Consult a materials safety data sheet (MSDS) for iodine. Identify the hazards that Courtois faced when he made his discovery. T/I A
3. Major chemical companies like 3M and Kodak have both chemists and chemical engineers on staff. Research how the roles of a chemist and chemical engineer in a company may differ. T/I
4. Many other scientific discoveries have been made "by accident." Chance played a role in Marie Curie's discovery of radiation, Edward Jenner's observations about smallpox, and Louis Pasteur's famous bacterial cultures. Explore one of these discoveries, or another that interests you, and prepare a presentation for your classmates. T/I C

5.3 Elements and Their Oxides

Metals and non-metals differ significantly in their properties. Most metals are shiny, malleable, and good conductors of electricity, while solid non-metals tend to be dull, brittle, and poor conductors of electricity. In addition, metals and non-metals form quite different oxides.

oxide a compound composed of oxygen and one other element

An **oxide** is a compound of any element combined with oxygen. Oxides can form when an element reacts with oxygen, often in air. This reaction can be rapid with the release of a great deal of energy, as in the combustion of magnesium. Alternatively it can be slow, like the gradual tarnishing of a freshly cleaned piece of magnesium. These two reactions are actually the same:

$$2\,Mg(s) + O_2(g) \rightarrow 2\,MgO(s)$$

acid a compound that produces hydrogen ions when mixed with water, forming a solution that conducts electricity, tastes sour, turns blue litmus red, and neutralizes bases

base a compound that produces hydroxide ions when mixed with water, forming a solution that conducts electricity, tastes bitter, turns red litmus blue, and neutralizes acids

Review of Acids and Bases

Before explaining the properties of oxides, we should review acids and bases. (You will study acids and bases in more detail in Unit 3. We will start here with simple definitions and introduce more complex concepts later.) An **acid** is a substance that produces hydrogen ions in water. These hydrogen ions cause the acidic properties of the solution. Hydrogen chloride is a gaseous molecular compound. When hydrogen chloride is bubbled into water, it forms hydrochloric acid, HCl(aq). This compound breaks down in water to produce hydrogen ions and chloride ions:

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

A **base** is a substance that produces hydroxide ions in water. The properties of basic solutions are due to the presence of hydroxide ions. Sodium hydroxide, NaOH(s), is a solid, ionic compound used in drain-opening products. It is a typical base. Sodium hydroxide produces sodium ions and hydroxide ions when it dissolves in water:

$$NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$$

The pH Scale

The pH scale is a numerical scale ranging from 0 to 14 (**Figure 1**). It is used to compare the acidity of solutions. One solution is more acidic than another because the dissolved compound produces more hydrogen ions. A solution with a pH of 7 is neutral. Solutions with a pH above 7 are basic; solutions with a pH below 7 are acidic.

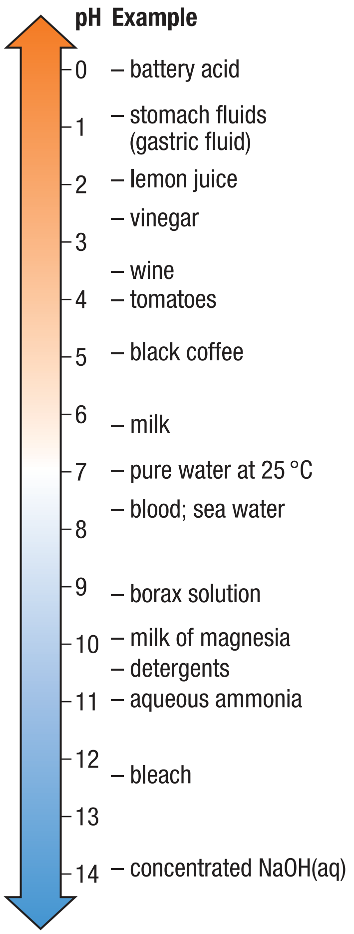


Figure 1 The pH scale is used to compare the acidity of a broad range of solutions.

Properties of Oxides

Non-metals, such as sulfur and nitrogen, may react to produce non-metallic oxides (**Figure 2**). Most non-metallic oxides such as sulfur dioxide or nitrogen monoxide are gases at room temperature and produce acidic solutions when dissolved in water. Oxides that form acidic solutions are called **acidic oxides**.

Metals, such as magnesium and calcium, react to produce metallic oxides (**Figure 3**). Metallic oxides tend to be ionic solids that produce basic solutions when dissolved in water. Oxides producing basic solutions are called **basic oxides**.

acidic oxide an oxide that forms an acidic solution when dissolved in water; a non-metallic oxide

basic oxide an oxide that forms a basic solution when dissolved in water; a metallic oxide

Figure 2 Sulfur is a solid, yellow non-metal that burns with a blue flame to produce colourless sulfur dioxide gas. Liquid sulfur is dripping into the beaker and solidifying.

Figure 3 Magnesium metal burns with a white-hot flame to produce magnesium oxide, a white powder.

Acidic Oxides

In earlier chemistry courses, you learned that acids produce hydrogen ions in water. These hydrogen ions give acids their acidic properties. A non-metallic oxide like carbon dioxide, CO_2, does not contain any hydrogen. Yet bubbling carbon dioxide into water can lower the pH of water from 7 to 5.6. Therefore, the reaction of carbon dioxide with water must result in a substance that produces hydrogen ions. In the following examples, we will examine how the oxides of carbon, sulfur, and nitrogen form acids.

CARBON DIOXIDE

Chemists think that the acidification of water by carbon dioxide most likely occurs in a two-step process:

1. Carbon dioxide undergoes a synthesis reaction with water to form carbonic acid:

   $$CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$$

2. Carbonic acid immediately breaks down in water, producing a hydrogen ion:

   $$H_2CO_3(aq) \rightarrow H^+(aq) + HCO_3^-(aq)$$

The reaction of carbon dioxide with water droplets in the atmosphere makes rain naturally mildly acidic. As a result, "normal rain" has a pH of about 5.6 rather than a neutral pH of 7. Oceanographers have evidence that unusually high levels of carbon dioxide in the atmosphere are affecting the chemistry of the oceans. In addition to contributing to climate change, carbon dioxide appears to be making the oceans more acidic. Increasing acidity hampers the growth of the shells of many marine organisms (**Figure 4**). At a certain pH, some animals may not be able to make shells at all.

Figure 4 This species of snail lives in Arctic waters. Like other organisms that form shells of calcium carbonate, it is threatened by the increasing acidity of the world's oceans.

NITROGEN OXIDES

Other non-metallic elements also form acidic oxides. Oxides of nitrogen contribute to the acidification of rain below its normal pH of 5.6. Some of the nitrogen oxides in the atmosphere form as a result of natural processes like lightning strikes. Most, however, are the by-product of the combustion of fuels.

Normally, most of the nitrogen in the atmosphere is made up of the diatomic N_2 molecule. This molecule is stable and unreactive because of its triple bond (**Figure 5**).

However, nitrogen makes up much of the air that is drawn into cars' engines. The high temperatures in these engines cause nitrogen to undergo an synthesis reaction with oxygen. The product is nitrogen monoxide (also called nitric oxide):

$$N_2(g) + O_2(g) \rightarrow 2\ NO(g)$$

LEARNING TIP

Generic Oxides
If an element forms more than one compound with oxygen, we can refer to these compounds by the generic term "oxides." For example, nitrogen forms at least two different oxides: nitrogen monoxide and nitrogen dioxide.

Investigation 5.3.1

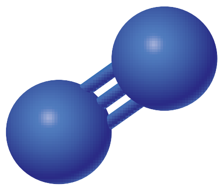
Properties of Oxides (p. 230)
What are the common properties of metallic oxides and non-metallic oxides? In this investigation, you will synthesize and compare some oxides.

CAREER LINK

Oceanographers have to understand the chemistry of marine environments. To find out more about preparing for this career,

GO TO NELSON SCIENCE

$:N{\equiv}N:$



(a) (b)

Figure 5 A nitrogen molecule can be represented by (a) a Lewis structure or (b) a molecular model. The triple bond joining the atoms makes the molecule very stable.

Nitrogen monoxide is released into the atmosphere through the car's exhaust pipe. It mixes with the air and undergoes a further reaction with oxygen to form nitrogen dioxide:

$2\ NO(g) + O_2(g) \rightarrow 2\ NO_2(g)$

Nitrogen dioxide is a toxic respiratory irritant and is involved in the formation of smog. It also combines with water droplets in the atmosphere to form nitric acid, $HNO_3(aq)$, and nitrous acid, $HNO_2(aq)$:

$2\ NO_2(g) + H_2O(l) \rightarrow HNO_3(aq) + HNO_2(aq)$

Fortunately, modern vehicles are equipped with a device called a catalytic converter that prevents most of the nitrogen oxides and other pollutants from being discharged in the exhaust (**Figure 6**). The interior of the converter is coated with beads of platinum, palladium, or rhodium catalyst. As nitrogen oxides come in contact with these beads, they decompose into harmless nitrogen and oxygen:

$2\ NO(g) \xrightarrow{\text{catalyst}} N_2(g) + O_2(g)$

Figure 6 (a) The catalytic converter, situated between the engine and the tailpipe, is part of a car's emission control system. (b) The ceramic matrix inside is covered with beads of platinum, palladium, and rhodium.

A catalyst is ideal for this purpose because, unlike a reactant, it is not consumed (used up) during the reaction. A well-maintained catalytic converter can often last the life of the automobile. Despite the catalytic converter, however, some pollutants are still released into the atmosphere. In Unit 5, you will learn how nitrogen oxides and unburned hydrocarbons from car exhaust contribute to the formation of smog. You will also learn how air pollutants, including smog, affect us.

SULFUR OXIDES

Sulfur oxides are a major contributor to acid precipitation. Some sulfur oxides are produced naturally as a result of volcanic activity. However, the bulk of sulfur oxide in the atmosphere results from the combustion of fossil fuels. Sulfur is a common impurity in many fossil fuels like coal and crude oil. Sulfur dioxide is produced when these fuels are burned:

$S(s) + O_2(g) \rightarrow SO_2(g)$

Once released into the atmosphere, sulfur dioxide reacts with additional oxygen to produce sulfur trioxide:

$2\ SO_2(g) + O_2(g) \rightarrow 2\ SO_3(g)$

Finally, sulfur trioxide combines with water in the atmosphere to form droplets of aqueous sulfuric acid. This mixture falls to Earth as acid precipitation:

$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

As you will see in later sections, many industries have devised clever technological solutions to treat sulfur-containing emissions before they become an environmental problem. Some industries follow the principles of green chemistry and capture the "waste" sulfur for use in other products.

CAREER LINK

The oil and gas industry has developed technology to capture sulfur. To find out more about careers in the oil and gas processing industry, such as a petroleum engineer,

GO TO NELSON SCIENCE

Basic Oxides

Most metallic elements form basic metallic oxides. Metallic oxides generally react with water to form metal hydroxides. A considerable quantity of thermal energy is often released as well. The word equation for this reaction is

metallic oxide + water → metallic hydroxide + thermal energy

Many metals form oxides simply by being exposed to the air. A freshly cut piece of sodium, for example, tarnishes in a matter of seconds (**Figure 7**) to form a coating of sodium oxide:

$4\ Na(s) + O_2(g) \rightarrow 2\ Na_2O(s) + \text{energy}$

Figure 7 A cylinder of sodium has just been cut, exposing the shiny metal. Within seconds, a layer of dull sodium oxide will cover this shiny surface.

Adding sodium oxide to water produces sodium hydroxide, a base (**Figure 8**):

$Na_2O(s) + H_2O(l) \rightarrow 2\ NaOH(aq) + \text{energy}$

Sodium hydroxide is one of the most widely used industrial chemicals. Its applications include the production of detergents and paper.

Figure 8 (a) The dark blue colour of the universal indicator test strip shows that sodium hydroxide solution is very basic. (b) Drain cleaners containing sodium hydroxide are very corrosive.

Calcium is another metal whose oxide, CaO(s), is an important industrial chemical (**Figure 9**). Calcium oxide is more commonly known as "lime." Calcium oxide reacts with water to form calcium hydroxide, $Ca(OH)_2$, also known as "slaked lime:"

$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq) + \text{energy}$

Figure 9 (a) Calcium burns to produce calcium oxide, CaO(s). (b) Adding water to calcium oxide produces calcium hydroxide and so much energy that steam rises from the mixture.

Calcium oxide and calcium hydroxide are very useful substances. When a lake is affected by acid precipitation, either compound can be added to raise the water's pH. Farmers use the compounds to treat acid soil, and hazmat (hazardous materials) teams use them to neutralize acid spills. Calcium oxide and calcium hydroxide are also important ingredients in the production of glass, cement, and paper.

UNIT TASK **BOOKMARK**

The Unit Task, described on page 242, involves the removal of copper(II) oxide from soil. You might find the information in this section useful as you plan to carry out the Unit Task.

5.3 Summary

- Non-metallic oxides are called acidic oxides because they form acidic solutions with water.
- Metallic oxides are called basic oxides because they form basic solutions with water.
- The acidity of carbon, nitrogen, and sulfur oxides causes serious environmental problems.
- A catalytic converter prevents many automotive pollutants from being emitted into the atmosphere.
- The basicity of metallic oxides can be used to neutralize unwanted acidity.

5.3 Questions

1. Draw a Venn diagram to compare the properties of acidic and basic oxides. K/U C
2. What effect would the addition of a small quantity of an acid to a basic solution have on the pH of the solution? Why? K/U
3. What type of solution will each of these compounds form when it is mixed with water? K/U
   (a) barium oxide
   (b) potassium oxide
   (c) chlorine dioxide
   (d) arsenic trioxide
4. Copy and complete the following pairs of chemical equations. Note that, in each pair, the product of the first reaction is a reactant in the second reaction. T/I C
   (a) $4\ K(s) + O_2(g) \rightarrow$ ______
   ______ $+ H_2O(l) \rightarrow 2\ KOH(aq)$
   (b) ______ + ______ $\rightarrow$ ______
   $Cl_2O(g) +$ ______ $\rightarrow 2\ HClO(aq)$
   (c) ______ + ______ $\rightarrow 2\ P_2O_5(s)$
   ______ + ______ $\rightarrow 2\ H_3PO_4(aq)$
5. Adding potassium oxide to water causes the pH of the water to rise to 10. Write a chemical equation to explain this observation. T/I C
6. Calcium oxide is a key ingredient in cement. What safety precautions do you think should be observed while mixing cement? Why? T/I A
7. Nitrogen oxide emissions from a coal-burning electricity generating station can be neutralized by either bubbling them through a mixture of calcium oxide and water or passing them through a large catalytic converter. Which option do you think is more sustainable? Why? T/I A
8. Plants generally do not grow near volcanoes that have recently erupted. Use what you know about oxides to explain this observation. K/U A
9. (a) Describe the reaction that would occur if a small piece of calcium oxide were placed in a dish of water.
   (b) Write a chemical equation for the reaction in (a).
   (c) What additional observation would you expect to make if the water contained phenolphthalein indicator? T/I C
10. In 1940, an example of cave art was discovered in southern France at Lascaux (**Figure 10**). Experts estimate that these cave drawings are about 15 000 years old. After the caves had been opened to tourists, the art began to deteriorate. T/I
    (a) Visit Lascaux online and view the cave art.
    (b) What caused the art to deteriorate? What was done to the save the art?

Figure 10 This painting of animals was created many thousands of years ago.

11. Find out more about the effects of ocean acidification. What effects might marine ecosystems experience as a result of this phenomenon? T/I A

GO TO NELSON SCIENCE

Neutralization Reactions

5.4

According to industry representatives, 97 % of discarded car batteries in Ontario are recycled (**Figure 1**). This process prevents the hazardous contents of car batteries from contaminating the environment. An automotive lead–acid battery consists of plates of lead and lead(IV) oxide immersed in a corrosive fluid containing concentrated sulfuric acid. Lead is toxic to a number of organs and systems in the body, including the heart and the nervous system.

Every part of a discarded battery is recycled into a useful product (**Figure 2**). The plastic cases of the batteries are cleaned, ground into pellets, and reused to manufacture new battery cases. Lead is recovered, purified, and used to manufacture new batteries. The sulfuric acid drained from the batteries is treated in two ways. Some is cleaned and reused in new batteries. The remainder is neutralized with sodium carbonate (also called soda ash), Na_2CO_3:

$$H_2SO_4(aq) + Na_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g) + Na_2SO_4(aq)$$

Sodium sulfate collected from this process is a key ingredient in the manufacture of glass and detergents.

Figure 1 Old car batteries contain lead, a toxic metal, and corrosive sulfuric acid.

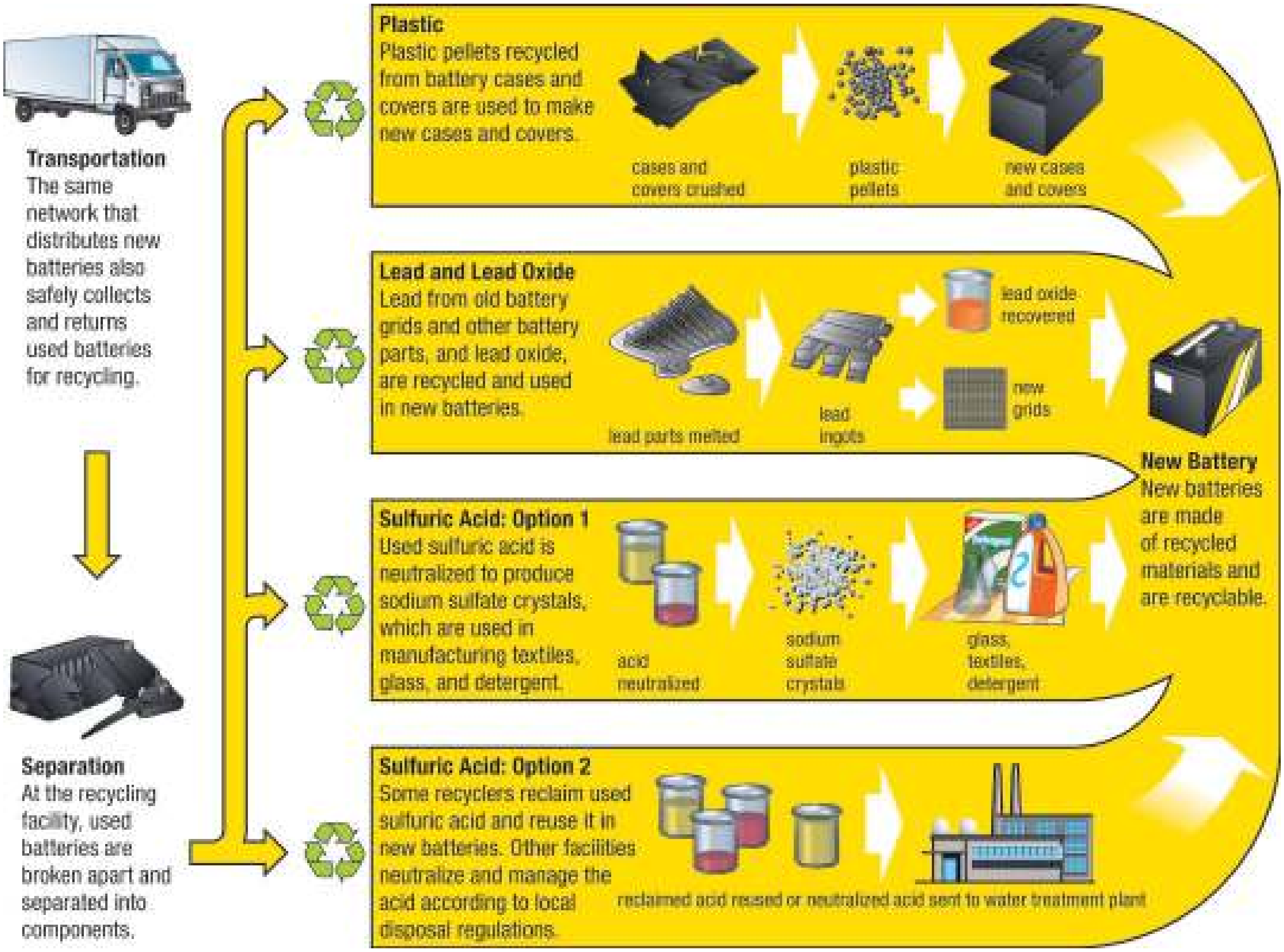


Figure 2 The recycling of car batteries is an excellent example of how hazardous substances can be reused, rather than discarded.

Investigation 5.4.1

Neutralizing Pain Relievers (p. 232)
In this investigation you will use a neutralization reaction to test which of two pain reliever tablets is the most acidic.

Types of Neutralization Reactions

In Section 4.6, you learned that neutralization reactions are a special case of double displacement reactions. During a neutralization reaction, an acid reacts with a base to produce a solution with a pH closer to 7 than its reactants.

Neutralizing an Acid

Imagine that you wanted to neutralize an acidic solution. Hydroxide compounds (bases) and carbonate compounds are both effective at neutralizing acids. We will look at these two reactions, and at some situations in which we want to neutralize an acid.

NEUTRALIZATION WITH A HYDROXIDE COMPOUND

The reaction of hydrochloric acid, HCl(aq), with sodium hydroxide, NaOH(aq), is a typical neutralization reaction. The chemical equation for this reaction is

$$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$$

This type of neutralization reaction follows the general pattern:

acid + base → water + ionic compound

You will discover in future chemistry courses than not all neutralization reactions follow this general equation. For most neutralization reactions, however, the hydrogen ions from an acid react with the hydroxide ions from a base to produce water (**Figure 3**). The ionic compound, which is sometimes called a "salt," is made up of the remaining ions.

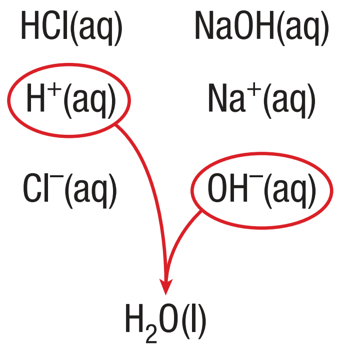


Figure 3 The hydrogen ion from the acid and the hydroxide ion from the base react to form water.

NEUTRALIZATION WITH A CARBONATE COMPOUND

Acids can also be neutralized by carbonate compounds. Calcium carbonate is the major component in limestone rock. It also makes up the shells of aquatic animals, such as clams, snails, and corals. Calcium carbonate reacts with sulfuric acid to produce carbonic acid, $H_2CO_3(aq)$. The chemical equation for this reaction is

$$H_2SO_4(aq) + CaCO_3(s) \rightarrow H_2CO_3(aq) + CaSO_4(s)$$

However, carbonic acid immediately decomposes into water and carbon dioxide:

$$H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$$

Therefore, the chemical equation for the net reaction of calcium carbonate with sulfuric acid is

$$H_2SO_4(aq) + CaCO_3(s) \rightarrow H_2O(l) + CO_2(g) + CaSO_4(s)$$ (**Figure 4**)

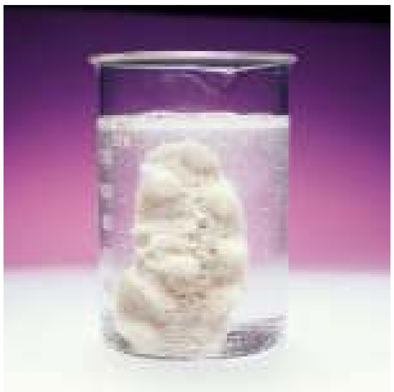


Figure 4 Coral contains mostly calcium carbonate. This compound neutralizes sulfuric acid. Bubbles of carbon dioxide appear as the reaction proceeds.

Carbonate compounds react similarly with other acids. In general, the reaction of acid with carbonate compound yields water, carbon dioxide, and an ionic compound:

acid + carbonate → water + carbon dioxide + ionic compound

Mini Investigation

Explaining Disappearing Ink

Skills: Performing, Observing, Analyzing, Communicating SKILLS HANDBOOK A1.1, A1.2

Phenolphthalein is an acid–base indicator that is magenta in basic solutions but colourless in acidic or neutral solutions. In this mini investigation, a neutralization reaction will occur to change the colour of the indicator. Your task is to determine the source of the acidic substance that makes this reaction occur.

Equipment and Materials: chemical safety goggles; lab apron; dropper bottles of dilute sodium hydroxide solution, NaOH(aq) (0.1 mol/L), and phenolphthalein indicator; beaker; dropper; paper towel

Sodium hydroxide solution is corrosive and may cause blindness if it comes into contact with the eyes. Any spills on skin or clothing should be washed immediately with plenty of cold water. Report any spills to your teacher.

Phenolphthalein indicator solution is flammable. Keep it away from open flames.

1. Put on your chemical safety goggles and lab apron.
2. Add 20 drops of the sodium hydroxide solution to the beaker.
3. Add 5 drops of phenolphthalein indicator to the beaker. Record your observations.
4. Swirl to mix the contents of the beaker.
5. Draw some of the mixture into the dropper.
6. Use the dropper to write a letter or symbol on the paper towel.
7. Hang the paper towel from a cabinet or desk. Observe how the intensity of the colour changes over time.

A. Describe the colour change that you observed as the paper dried. T/I

B. Did the pH of the solution on the paper increase, decrease, or stay the same over the course of the investigation? What is your evidence? T/I

C. Suggest an explanation for the colour change. Justify your explanation. T/I

D. Write a chemical equation for the reaction that changed the colour of the ink. T/I C

Neutralization Reactions in Action

You now know the chemistry behind neutralization reactions. Acids and bases are some of the most widely used industrial chemicals. Many useful household products also contain acids and bases. The neutralization of these compounds has important consumer and environmental applications.

NEUTRALIZING STOMACH ACIDS

The digestive fluids in your stomach have a pH of about 1.5—acidic enough to slowly "eat" through a strip of zinc metal. The source of this acidity is hydrochloric acid produced by specialized cells in the stomach lining. Hydrochloric acid aids in the digestion of proteins and suppresses the growth of unwanted bacteria. Since your digestive fluids cannot distinguish between food and stomach tissue, they could easily digest the stomach itself. Fortunately, the stomach has defences against being digested: other cells in the stomach lining produce mucus that coats the inside of the stomach, keeping the acid away from the living cells. Another defence is the ability of the cells of the stomach lining to replace themselves very quickly: at a rate of about half a million per minute.

Stomach acid can become a problem if it leaks upward out of the stomach, through a valve-like structure and into the esophagus. Irritation of the esophagus by stomach acid is called gastroesophageal reflux, acid reflux, or "heartburn." This condition can be caused by a number of factors, including overeating and spicy foods. A short-term fix for heartburn is to consume an antacid medication that will neutralize the excess acid in the stomach (**Figure 5**).

The active ingredient in an antacid must be effective at neutralizing excess stomach acid and yet it must not irritate tissue on its way to the stomach. Ideally, the neutralizing compound would produce only enough hydroxide ions to react with the acid. Weakly basic substances such as magnesium hydroxide and aluminum hydroxide are used in some antacid products. Because they are only slightly soluble, neither compound produces enough hydroxide ions to irritate the mouth or throat as the tablet is

LEARNING TIP

Acid Formulas and Hydrogen
Ethanoic acid (acetic acid) molecules produce one hydrogen ion in solution. That is why one hydrogen atom is written first in the chemical formula: $HC_2H_3O_2$, rather than $C_2H_4O_2$. The chemical formula for tartaric acid is $H_2C_4H_4O_6$, indicating that this compound can produce two hydrogen ions in solution.

Figure 5 Antacids contain hydroxide or carbonate compounds. The recommended dosage is intended to neutralize some, but not all, of the stomach acid.

chewed and swallowed. The chemical equation for the neutralization of hydrochloric acid with aluminum hydroxide is

$$3\ HCl(aq) + Al(OH)_3(s) \rightarrow AlCl_3(aq) + 3\ H_2O(l)$$

Carbonate compounds are also safe yet effective ingredients in antacid products. Sodium hydrogen carbonate, $NaHCO_3$, and calcium carbonate, $CaCO_3$, are commonly used. The equations for the neutralization of hydrochloric acid with these compounds are

$$HCl(aq) + NaHCO_3(s) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$$

$$2\ HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

Frequent consumption of sodium hydrogen carbonate is not recommended because it adds to the sodium ions we already consume by salting our food. Excess sodium ions have been linked to hypertension and stroke. Too much calcium carbonate may increase the risk of kidney stones.

Figure 6 (a) A lake with a low pH is a stark contrast to (b) a lake with a normal pH.

REHABILITATING LAKES

Emissions of sulfur oxide and nitrogen oxide are largely responsible for the acidification of Ontario's lakes. These oxides are acidic, and cause acid precipitation. Despite considerable progress in curbing these emissions over the last 30 years, the problem persists. Aquatic ecosystems generally can only tolerate small reductions in the pH of their water. Organisms fail to reproduce and populations crash (**Figure 6**). As pH drops below 5, only the most acid-tolerant organisms survive.

Lakes differ in their ability to resist changes in acidity. One factor that helps offset the effects of acid precipitation is the type of rock beneath the lake. Much of the surface rock of southern Ontario is limestone (mostly calcium carbonate, $CaCO_3$). An important characteristic of limestone is that it reacts slowly with acids (**Figure 7**). Limestone in southern Ontario lakes acts as a natural antacid, helping to neutralize acidified lake water. Unfortunately, the surface geology of northern Ontario is mostly granite. Granite is a hard, impervious rock that is less reactive than limestone. As a result, the lakes in this region have little natural capacity to neutralize acid precipitation. Consequently, these northern lakes are the most prone to environmental damage from acid precipitation.

One way to assist lakes threatened by acid precipitation is to neutralize the acidity by adding calcium oxide or lime (**Figure 8**). However, "liming" a lake only solves the acidity problem in the short term. The long-term solution is to eliminate the pollutants at their source. In later sections of this chapter, you will learn about new technologies that help industries reduce their acidic oxide emissions.

Figure 7 Limestone reacts with acid. Regions where the underlying rock is limestone have some protection against acid precipitation because of this reaction.

Figure 8 The addition of calcium oxide (lime) to a lake acidified by acid precipitation

BAKING

Substances that make bread dough rise are called leavening agents. Leavening agents produce bubbles of carbon dioxide that are trapped within the dough. As the bubbles expand, they push the elastic dough upward.

The carbon dioxide bubbles could be produced either by baker's yeast or by a neutralization reaction involving a carbonate compound. Baking soda and baking powder are both leavening agents (**Figure 9**).

Baking soda is pure sodium hydrogen carbonate. The leavening action of baking soda is activated when it is mixed with an acidic ingredient like fruit juice, vinegar, or buttermilk. The chemical equation for the reaction of the citric acid, $H_3C_6H_5O_7(aq)$, in fruit juice with baking soda is

$$H_3C_6H_5O_7(aq) + 3\ NaHCO_3(s) \rightarrow Na_3C_6H_5O_7(aq) + 3\ H_2O(l) + 3\ CO_2(g)$$

Bubbles of carbon dioxide gas from this reaction cause the dough to rise. The warmth of the oven makes the bubbles expand. Because baking soda is unstable at high temperatures, it is generally used when a short baking time is required. That is why you will see baking soda listed as an ingredient in many cookie recipes.

Baking powder, however, is a mixture of several ingredients. Two essential ingredients in most types of baking powder are baking soda and a dry acid such as tartaric acid, $H_2C_4H_4O_6$. Baking powder is activated by moisture. As solid tartaric acid dissolves and produces hydrogen ions, sodium hydrogen carbonate neutralizes its acidity. The resulting products are carbon dioxide (which makes the dough rise), water, and sodium tartrate. The chemical equation for this reaction is

$$H_2C_4H_4O_6(aq) + 2\ NaHCO_3(s) \rightarrow Na_2C_4H_4O_6(aq) + 2\ H_2O(l) + 2\ CO_2(g)$$

Baking powders can be fast acting and slow acting, depending on the acid they contain. The acids in fast-acting baking powders (usually tartaric acid) react immediately with baking soda to produce carbon dioxide. Slow-acting baking powders contain acids that only begin to react when heated in the oven. Slow-acting baking powders are used when longer baking times are required. The recipes for cakes usually specify baking powder as the leavening agent.

Figure 9 Neutralization reactions involving baking powder produce bubbles of carbon dioxide gas, giving cakes and muffins their spongy appearance. A similar result can be achieved by using baker's yeast.

Research This

Neutralizing Hair Care

Skills: Researching, Analyzing, Evaluating, Communicating

SKILLS HANDBOOK A5.1

Human hair grows from tiny openings in the skin called follicles. The pigment that gives hair its colour is located in the inner core of each hair strand. The visible layer of a hair strand is called the cuticle. The cuticle consists of tiny, flat plates that overlap each other much like the scales of a snake's skin, or shingles on a roof (**Figure 10**). Small changes in pH can have a profound effect on the cuticle and, as a result, on the properties of your hair.

Figure 10 A hair growing out of a hair follicle

1. Research the effect of changes in pH on the scales of the cuticle.
2. Research the role that pH and neutralization reactions play in permanent hair-colouring treatments.

A. Based on your research, summarize the possible damaging effects of permanent hair dyes. T/I

B. Which damaging effect is, in your opinion, the most significant? Why? What precautions, if any, can be taken to avoid this effect? A

C. Now that you have explored the topic of permanent hair dyes, how do you feel about using them? Explain your reasons. T/I A

D. Present your findings in a novel way intended for a group of younger high school students who are considering permanently colouring their hair for the first time. C A

GO TO NELSON SCIENCE

Choice of Neutralizing Reactant

Neutralization reactions are important parts of many chemical processes. Some, like the recycling of car batteries, are large industrial processes in which corrosive substances must be neutralized before they can be discharged into the environment. Other reactions involve consumer products such as antacids and hair dyes. In each case, the effectiveness of the process depends on the properties of the reactants used.

Chemists consider a number of factors when selecting an appropriate neutralizing reactant. Safety is always an important factor, along with the reactant's cost and chemical properties. For example, solutions of sulfuric acid are very acidic because they produce a lot of hydrogen ions in solution. Sulfuric acid would be effective at neutralizing a base, but it is not always the best choice. Sulfuric acid is hazardous to work with and can corrode other materials on-site, because of all the hydrogen ions in solution. There could also be leftover sulfuric acid after the neutralization reaction. This residual acid could make the site highly acidic, which would defeat the purpose of neutralization. Moreover, reactions involving concentrated sulfuric acid can generate enough thermal energy to make the reaction mixture boil and splatter. It is often safer to use a less acidic acid, such as ethanoic acid, $HC_2H_3O_2(aq)$, the acid in vinegar. Ethanoic acid can effectively neutralize the base with minimal risk to workers or materials.

Figure 11 Skin contact with sodium hydroxide can result in severe chemical burns.

Similarly, a very basic substance such as sodium hydroxide is not always the wisest choice for neutralizing an acid. Sodium hydroxide is a white crystalline solid that is highly soluble in water. When it dissolves, sodium hydroxide produces a great many hydroxide ions into the solution. This makes sodium hydroxide very basic and highly corrosive to both workers and materials (**Figure 11**). Neutralization reactions involving sodium hydroxide can also release a lot of thermal energy, which can cause the solution to boil and splash. Any sodium hydroxide remaining after the acid has been neutralized may remain on the site, making the soil highly basic.

A better choice for neutralizing a waste acid might be a less basic substance such as calcium hydroxide, $Ca(OH)_2$. Calcium hydroxide is less basic than sodium hydroxide because of differences in their solubility. According to the solubility table (Section 4.6, Table 1), calcium hydroxide is only slightly soluble in water. Therefore, this base produces fewer hydroxide ions than the equivalent amount of sodium hydroxide. This makes calcium hydroxide less corrosive and safer to handle and transport than the more basic hydroxides.

Research This

Neutralizing a Base

Skills: Researching, Analyzing, Evaluating, Communicating, Defending a Decision

SKILLS HANDBOOK A5.1

Sodium hydroxide (also called caustic soda) is one of the most widely used industrial chemicals. It is also one of the most corrosive. Sodium hydroxide is the active ingredient in many types of drain openers. Consequently, an accidental spill of a drain cleaner is one of the most hazardous chemical spills a homeowner could encounter.

1. Research two different substances that could be used to neutralize a drain opener spill.

A. Write the chemical equation for the neutralization of sodium hydroxide with each substance. K/U C

B. What safety precautions should be followed when cleaning up the spill? T/I

C. Based on your research, which substance would you recommend using to clean up the spill? Why? T/I A

GO TO NELSON SCIENCE

A Versatile Neutralizing Agent

Sodium hydrogen carbonate, $NaHCO_3$, or baking soda, is unusual because it can neutralize both acids and bases. It is therefore a very versatile neutralizing compound. It is also one of the safest, since it is not corrosive at all. In fact, it is safe enough to be used in food and consumer products such as toothpaste. Acids react with sodium hydrogen carbonate following the general pattern for carbonate reactions. For example, sodium hydrogen carbonate reacts with hydrochloric acid as follows:

$$HCl(aq) + NaHCO_3(s) \rightarrow H_2O(l) + CO_2(g) + NaCl(aq)$$

Sodium hydrogen carbonate reacts with bases to produce water and sodium carbonate. For example:

$$NaOH(aq) + NaHCO_3(s) \rightarrow H_2O(l) + Na_2CO_3(aq)$$

5.4 Summary

- Acids can be neutralized by either a hydroxide compound or a carbonate compound.
- General neutralization reactions:
  acid + base → water + ionic compound
  acid + carbonate → water + carbon dioxide + ionic compound
- Very acidic substances produce more hydrogen ions than do mildly acidic substances.
- Very basic substances produce more hydroxide ions than do mildly basic substances.
- Neutralization reactions are useful in many industrial, consumer, and environmental applications.
- Reactants are selected to be as safe and inexpensive as possible, and to have the desired results.

5.4 Questions

1. Complete and balance the following equations: T/I C
   (a) $HNO_3(aq) + Ca(OH)_2(aq) \rightarrow$
   (b) $HNO_3(aq) + K_2CO_3(aq) \rightarrow$
   (c) $HC_2H_3O_2(aq) + Al(OH)_3(s) \rightarrow$
   (d) $H_3PO_4(aq) + NaHCO_3(aq) \rightarrow$
   (e) $H_3PO_4(aq) + Ca(OH)_2(aq) \rightarrow$
   (f) $H_2SO_4(aq) + Ca(HCO_3)_2(aq) \rightarrow$
   (g) $HClO_3(aq) + CaCO_3(s) \rightarrow$
2. Barium sulfate can be made using two different neutralization reactions, using the following substances:
   $Ba(OH)_2$, $Ba(HCO_3)_2$, H_2SO_4
   Write chemical equations for these two reactions. Each reaction should have only two reactants. T/I C
3. Carbonated drinks are made by bubbling pressurized carbon dioxide gas into them. K/U C A
   (a) What acid forms as a result of this addition?
   (b) Sodium hydroxide is a common ingredient in drain openers. Write the chemical equation for the neutralization of dissolved sodium hydroxide in a drain opener with a solution of the acid in (a).
   (c) What do you think are the advantages of using pop instead of hydrochloric acid, HCl(aq), to neutralize a spill of drain cleaner?
4. Sometimes cooked fish has a distinctive fishy smell because of basic compounds called amines. Describe why a squirt of lemon juice can eliminate these odours. K/U A
5. In 2007, a train derailment spilled a large volume of concentrated sulfuric acid into the Blanche River near Englehart, Ontario. During the cleanup, calcium oxide (lime) was added to the river upstream from the site of the accident. How did this action help restore the river? K/U A
6. Industrial acid spill kits contain, among other things, an absorbent material to soak the spill and a neutralizing reactant. Often, the neutralizing agent is incorporated into the absorbent material. Why is this advantageous to the workers cleaning up a spill? K/U A
7. Acidosis and alkylosis are two medical conditions caused by acidity imbalances in the body. Research the causes of these conditions and how they are treated. In your report, include a description of how these treatments work. T/I A

GO TO NELSON SCIENCE

5.5 Mining, Metallurgy, and the Environment

Figure 1 Copper is an important component in computer systems because it is such a good conductor of electricity.

Copper is one of those substances that you rarely think about but could not do without. What would your life be like if all the copper suddenly disappeared? You may have no running water, because the water pipes in your home are probably made of copper. You also might have no electricity, since most of the wires that distribute electricity are made of copper. Furthermore, all of your electronic devices could not function without their copper components (**Figure 1**).

What Is Mining?

mineral a naturally occurring solid that has a definite crystal structure and chemical composition

ore rock containing a relatively high proportion of a desirable mineral

Copper rarely occurs in nature as a pure element. Instead, it is found in minerals. A **mineral** is a solid substance found in Earth; usually a crystalline ionic compound, or mixture of compounds (**Figure 2**). Occasionally, however, a mineral can be made up of a single element such as gold, silver, or even copper. The chemical composition of every mineral is unique. Azurite, $Cu_3(CO_3)_2{\bullet}(OH)_2$, and bornite, Cu_5FeS_4, are examples of minerals that contain copper ions.

Rock that contains enough of a valued resource that it can profitably be mined is called an **ore**. Highly desired metals such as gold and silver are known as "precious metals." Less valuable metals—copper, nickel, and zinc—are sometimes called "base metals."

LEARNING **TIP**

Base Metals

Do not let the term "base metals" confuse you. Their name has nothing to do with the bases that neutralize acids. Rather, the word "base" has a very ancient origin, and means "common" or "of little value."

To access the ore, mining companies must first excavate it out of the ground. This usually involves either surface mining or underground mining. Surface mining involves opening a large pit to expose mineral deposits (**Figure 3**). Surface or open-pit mines extract ores that are close to Earth's surface. Vast quantities of material must be removed to obtain the minerals in the ore. Returning the site to its natural state once mining operations cease is a huge environmental challenge.

In an underground mine, miners drill shafts deep into Earth to reach the ore.

Figure 2 Clockwise from lower left: native copper, bornite, azurite, and a mixture of malachite and azurite. Only native copper contains copper atoms. The other minerals contain either Cu^+ or Cu^{2+} ions.

Figure 3 The Diavik diamond mine, 300 km northeast of Yellowknife, Northwest Territories, is Canada's largest diamond mine.

What Is Metallurgy?

metallurgy the science and technology of separating and refining metals from their ores and subsequent processing

Once an ore is mined it must be processed to extract its valued components. **Metallurgy** involves extracting metals from their ores and processing them into a useful form. Metallurgy often involves a number of chemical and physical processes. Sometimes the extraction and processing procedures are done where the ore is mined. In other cases, the ore is shipped elsewhere for processing.

A Typical Metallurgical Process: Extracting Copper

A typical ore sample from an operating mine contains surprisingly little metal. Copper ore, for example, can contain as little as 0.5 % copper by mass. This means that every 1000 t of ore might contain only 5 t of copper. A huge quantity of rock needs to be processed and discarded to extract the "good stuff." The discarded materials from this process are called tailings. The challenge for the metallurgical engineer is to design a profitable process for separating copper from its ore, changing it to metallic copper, and then purifying it. **Figure 4** summarizes the key steps of processing a metal such as copper from its ore into a finished product.

CAREER **LINK**

What do metallurgical engineers do? Where do they work? How do they use chemistry? To find answers to these questions, and to research this career,

GO TO NELSON SCIENCE

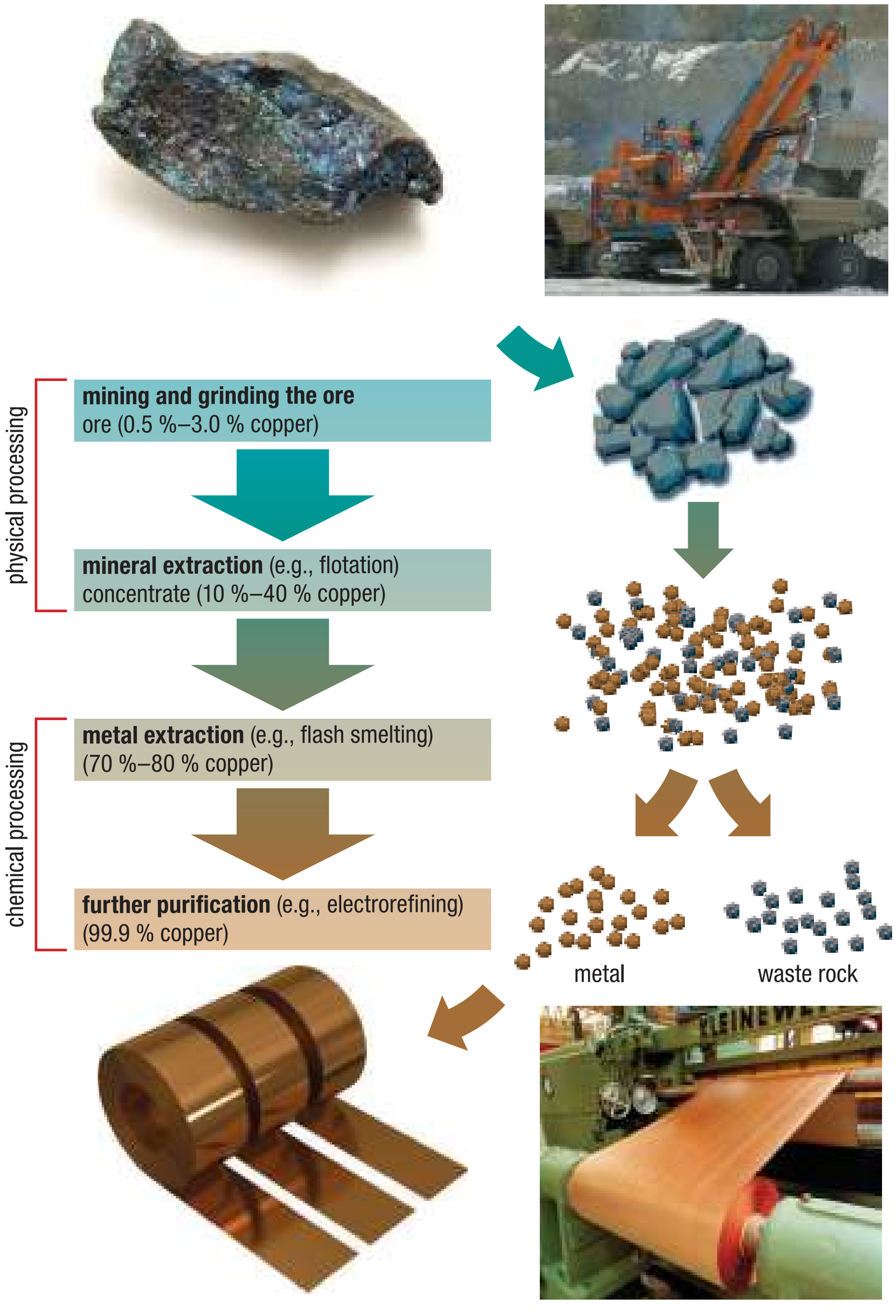


Figure 4 Extracting and purifying copper metal from its ore involves both physical and chemical processing, in several stages. A great deal of rock must be processed to extract a small quantity of metal.

PHYSICAL PROCESSING

Ore fragments extracted from the mine can be as large as a small car. Huge grinding machines crush them into small pieces. The ore is then concentrated by an ingenious technology called flotation (**Figure 5**, next page). During flotation the crushed ore is mixed with water, air bubbles, and a detergent-like ionic compound. This compound breaks down into ions when it dissolves. The resulting anion is negatively charged at one end and non-polar at the other. The negatively charged end is hydrophilic (attracted to water); the non-polar end is hydrophobic (repelled by water).

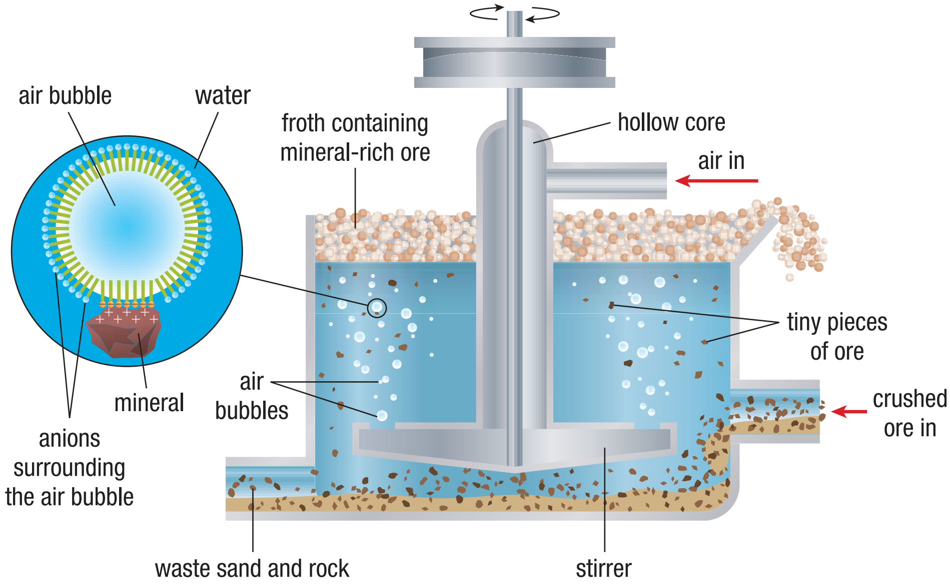


Figure 5 In a flotation cell, particles of metal-bearing ore attach to bubbles and float to the surface, where they are skimmed off.

Figure 6 A flotation cell separates powdered ore with high metal content from worthless rock.

The charged hydrophilic ends of these anions attach themselves to copper-bearing ore particles. The hydrophobic end enters the "skin" of air bubbles. Rising air bubbles pull ore particles to the surface of the tank, where they can be skimmed off (**Figure 6**). The concentrate collected from flotation contains from 30 % to 40 % copper.

CHEMICAL PROCESSING

smelting the chemical process that extracts a metal from its ore using heat and chemicals

The goal of chemical processing is to extract copper from the ore concentrate and convert it into pure copper metal. In the past, this was done using a process called smelting. **Smelting** involves heating ore in air, together with other chemicals, to over 1000 °C in a furnace. The chemical reactions in the smelter produce liquid copper metal. The copper sinks to the bottom of the smelter, where it is separated. Waste materials are skimmed off and discarded.

In the process, any sulfide that may have been present in the ore is converted to sulfur dioxide gas. In the early days of smelting, sulfur dioxide gas went up the furnace chimney untreated and into the air. The acid rain caused by these emissions devastated ecosystems in the surrounding area. Later, the installation of devices called scrubbers prevented much of the sulfur dioxide emissions from escaping.

flash smelting a fairly new technology for separating a metal from its ore by heating ore in an atmosphere of almost pure oxygen

A new technology called flash smelting is becoming the preferred method of processing copper, particularly from sulfide-containing ores. Like traditional smelting, **flash smelting** involves heating ore with other chemicals to high temperatures. However, rather than using air, flash smelting is done in almost pure oxygen. As a result, the chemical reactions in the smelter occur more vigorously. So much thermal energy is released by these reactions that very little additional fuel is needed to keep the furnace hot. This benefits the company in two ways:

- reduced fuel costs
- less pollution

Flash smelting produces more concentrated sulfur dioxide than traditional smelting because pure oxygen is used. This makes it economically feasible for the company to convert sulfur dioxide into sulfuric acid, which they can sell to help offset costs.

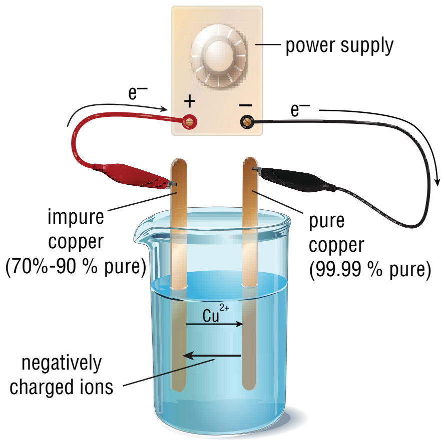


Figure 7 Electricity can be used to produce very pure copper. Industrial-scale versions of this setup purify many large sheets of copper at a time.

FURTHER PURIFICATION

Flash smelting can produce copper that is up to 90 % pure. This may be adequate for water pipes, but not for electrical applications. Impurities decrease copper's electrical conductivity, so copper used in the electrical industry must be over 99 % pure. In order to achieve this standard, copper is further purified, often using an electrical process called electrolysis (**Figure 7**). In this process, a large sheet of impure copper slowly breaks down into ions with the aid of an electric current. The pure copper ions in the solution travel through the solution and collect on a thin sheet of pure copper. The impurities fall to the bottom of the container.

Mini Investigation

Recovering Copper

Skills: Predicting, Planning, Performing, Observing, Analyzing, Communicating

SKILLS HANDBOOK A1.1, A1.2

Electrolysis is used to produce highly pure metals. It can also be used to extract toxic metal ions from wastewater. In this investigation, you will use electrolysis to extract copper metal from a solution containing copper(II) ions, $Cu^{2+}(aq)$.

Equipment and Materials: chemical safety goggles; lab apron; 250 mL beaker; 2 electrical leads with clips; power supply; 2 pencils with both ends sharpened; 100 mL of dilute copper(II) sulfate solution, $CuSO_4(aq)$ (0.1 mol/L)

Copper(II) sulfate is toxic. It is also a skin and eye irritant. Wash your hands thoroughly after pouring the solution.

When disconnecting the power supply, pull on the plug rather than the cord.

1. Put on your chemical safety goggles and lab apron.
2. Connect one end of one electrical lead to the end of one pencil. Connect one end of the other electrical lead to the end of the other pencil.
3. Connect the other ends of the leads to the terminals of the power supply.
4. Pour about 100 mL of the copper(II) sulfate solution into the beaker.
5. Place the free ends of the pencils into the copper sulfate solution so that their tips do not touch.
6. Turn on the power supply and sent it to 6 V.
7. After 1 to 2 min, turn off the power supply and remove the pencils from the solution. Observe their tips.
8. Dispose of your materials and clean up your workstation as directed. Wash your hands.

A. What solid was produced at one of the tips? T/I
B. What was the source of the solid? T/I
C. Predict what changes you would observe in the blue colour of the solution if you left the power on for a longer period of time. T/I
D. How would you know that all of the copper ions have been removed from the solution? T/I
E. Design a process that will remove zinc ions from a zinc nitrate solution. If time permits, and with the approval of your teacher, proceed with your plating procedure. Record your observations. T/I

UNIT TASK BOOKMARK

You might find the technique outlined in this Mini Investigation useful as you plan for the Unit Task described on page 242.

The Impact of Mining on the Environment

Mining has been responsible for a large portion of Canada's income over the years. However, the mining and mineral processing industry also has a long history of environmental mishaps and human tragedy. Mining engineers, supervisors, and employees must constantly be on the alert for potential problems. For example, large quantities of sodium cyanide, NaCN, are used to extract gold out of ore. Because cyanide is so toxic, accidental

WEB LINK

To find out more about the effects of mining—both positive and negative—on Canada,

GO TO NELSON SCIENCE

discharges can be devastating to local wildlife. Acidic oxide emissions from smelting operations kill vegetation and make the surrounding regions barren.

Environmental disasters were common in the past, when there was little regulation of the mining industry. The situation is quite different today—at least in North America. Mining companies must fund detailed environmental assessments before they get permission to open a new mine. Moreover, plans for the closure of the mine and rehabilitation of the site must be in place before the first shovel enters the ground. Recent technological advances, together with stricter government regulations, have greatly reduced the environmental impact of the mining industry. Despite these advances, the risk of contamination of our air and water remains.

Not all countries have such strict mining regulations. In 2010, an environmental disaster occurred near an aluminum mine in Hungary. Aluminum is obtained from a mineral called bauxite. When bauxite is processed, it is crushed and mixed with a solution of sodium hydroxide. The resulting basic mixture is commonly called "red mud." After the aluminum compounds are removed, the red mud is sometimes stored in holding ponds. In October, 2010, a holding pond in Hungary burst its banks and released a flood of corrosive sludge into nearby towns and villages (**Figure 8**). The short-term effects were serious—at least eight people died—but the long-term effects of heavy metal ions in the environment may be even more problematic.

Figure 8 The red mud that flowed out of the holding pond in Hungary had a very high pH and contained potentially dangerous metal ions.

Air Pollution

Mining operations are a source of many air pollutants. Surface mining, in particular, is a major source of dust. Digging, blasting, hauling, and crushing the large quantities of dirt and rock all produce rock dust. Wind sweeping across the overturned rubble can spread this dust over a wide area.

Emissions of sulfur dioxide remain the single most important air pollutant of the mining industry worldwide. A significant portion of acid precipitation in Canada results from sulfur dioxide emissions from the United States. Thanks primarily to improvements in technology and stricter emission standards, sulfur dioxide emissions are declining (**Figure 9**). New, stricter emission standards will ensure that they continue to decline into the future. This is not necessarily the case in other countries, however.

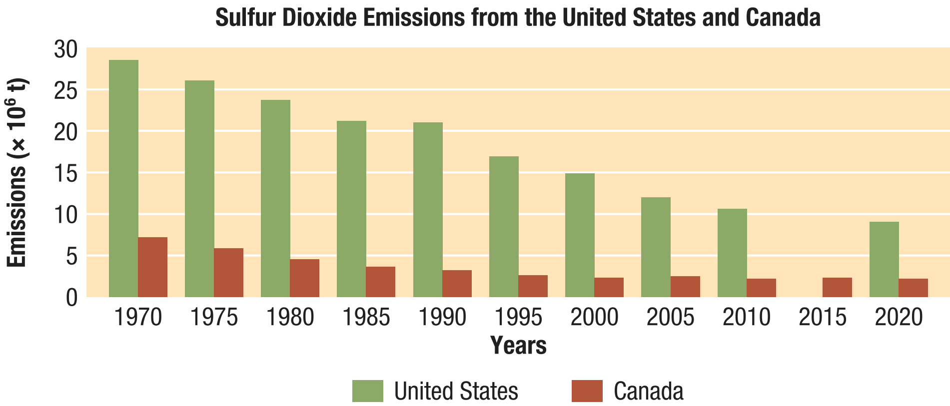


Figure 9 Sulfur dioxide emissions have fallen significantly in both Canada and the United States since 1970. The data for 2010, 2015, and 2020 are predictions.

Water Pollution

Mining uncovers rock that has been buried deep in the ground for much of Earth's history. Freshly exposed rock is prone to erosion. This can result in a water quality problem called acid mine drainage. Acid mine drainage is the outflow of unusually acidic water from a mine or tailings dump site. The excess acidity is caused by the reaction of sulfide minerals such as pyrite, FeS_2, reacting with oxygen and water. The chemical equation for this reaction is

$$4\ FeS_2(s) + 15\ O_2(g) + 2\ H_2O(l) \rightarrow 4\ Fe^{3+}(aq) + 8\ SO_4^{2-}(aq) + 4\ H^+(aq)$$

The hydrogen ions produced in the reaction make the water acidic. Iron(III) ions, and the precipitates they form, give the polluted water a rusty colour (**Figure 10**). Acidified groundwater can easily dissolve toxic metals, such as aluminum, from the rock. These contaminants can harm aquatic life in nearby bodies of water.

One approach to dealing with acid mine drainage is to neutralize the acidity with calcium carbonate (in limestone). This raises the pH of the water to more normal levels. It also promotes the precipitation of metal ions, such as Fe^{3+}(aq), as hydroxides.

Figure 10 The rusty colour of this stream is caused by iron(III) hydroxide precipitates.

5.5 Summary

- Mining removes minerals from the ground.
- Metallurgy is the technology of obtaining and refining metals.
- Flotation is used to concentrate ore.
- Smelting is the chemical process used in metallurgy to extract a metal from its ore.
- Flash smelting is more efficient than traditional smelting. Much of the energy comes from its chemical reactions, and the by-products of flash smelting can be captured and sold as valuable raw materials for other chemical processes.
- Acidic mine drainage can contaminate groundwater and aquatic ecosystems.

5.5 Questions

1. Why is it necessary to concentrate ore before it is chemically processed? K/U
2. Does flotation rely on a physical or chemical property of the ore? Explain. K/U
3. Identify two ways in which flash smelting is a greener technology than traditional smelting. K/U
4. How can groundwater passing through mine tailings be a threat to a nearby aquatic ecosystem? K/U
5. The reaction of the mineral pyrite, FeS_2, with air and water is a natural process that can acidify groundwater in undisturbed ground. Why do you think mining speeds up this process? K/U T/I
6. The acidity caused by acid mine drainage can leach toxic aluminum ions, Al^{3+}(aq), out of rock. One way to treat this problem is to add calcium oxide, CaO, or lime, to the water. K/U T/I A
   (a) Explain the effect this has on the pH of the water.
   (b) Use the solubility table (Section 4.6, Table 1) to explain how adding lime removes aluminum ions from water.
7. Future legislation will require further reductions in sulfur dioxide emissions. To meet these targets, some mining companies will have to update their processes. Some companies may decide to close their mines or smelters, rather than make costly upgrades. A mine closure can devastate the economies of nearby towns. In your opinion, should a government subsidy be used to keep these mines operating? Give one argument for and one against. T/I A
8. The carbonyl process is used to produce highly pure nickel. It occurs in two steps. First, impure nickel reacts with carbon monoxide gas to produce a gas called nickel carbonyl:

   $$Ni(s) + 4\,CO(g) \xrightarrow{50\ ^\circ C} Ni(CO)_4(g)$$

   Nickel carbonyl vapour is then purified and injected into a reaction vessel containing pellets of hot solid nickel at about 200 °C. At this temperature, nickel carbonyl readily breaks down:

   $$Ni(CO)_4(g) \xrightarrow{200\ ^\circ C} Ni(s) + 4\,CO(g)$$ K/U T/I

   (a) Classify each of these reactions as synthesis, decomposition, or displacement.
   (b) Research the hazards of this process. Why must it be done in sealed vessels?
9. In 2007, Canada's provincial ministers responsible for mines encouraged the mining industry and Canada's Aboriginal peoples to develop a stronger working relationship. Research specific examples that show how this would benefit both groups. T/I A

GO TO NELSON SCIENCE

5.6 Detox for Contaminated Land

Chemical waste dumps, tailings ponds, slag heaps—these are all legacies of our industrial past: places where waste has been deposited with no concern for human health or environmental safety. Dump sites like these are not the only examples of land contaminated with hazardous substances. Other sites include abandoned businesses, landfills, warehouses, and military bases that used or stored chemicals.

Environment Canada reports that there are tens of thousands of sites across Canada contaminated with hazardous substances. Some of the contamination occurred at a time when the consequences of releasing chemicals into the environment were not completely understood. Other contaminated sites are the result of indiscriminate dumping of chemicals during a period of lax or non-existent environmental laws. Perhaps one of the most infamous contaminated sites in Canada is the Sydney tar ponds in Sydney, Nova Scotia.

WEB LINK

To see photos of the Sydney tar ponds, before and after cleanup,

GO TO NELSON SCIENCE

What Is Contaminated Land?

Contaminated land is land that contains hazardous substances in sufficient concentration to pose a threat to human health or to the environment. These substances may be solids, liquids, or gases. Once in the ground, they can interact with groundwater and soil. For example, soluble solid materials such as metal compounds dissolve and contaminate groundwater. Organic substances, including oil and solvents, bind to particles in the soil. Understanding the properties of the contaminants is the first key step in deciding how best to deal with them. The process of decontaminating land or water so that it can safely be used again is called **remediation**.

remediation the process of treating contaminated land or water so that it is safe for use again

Remediation versus Total Cleanup

The remediation of contaminated land is not necessarily the same as a total cleanup. A total cleanup implies the complete removal of hazardous substances (**Figure 1**). A total cleanup, however, is not always practical or even necessary in order to make contaminated land useful again. Much depends on the intended use of the land. For example, land that will become a children's playground needs far more aggressive remediation than land intended to become a parking lot. Given the number of contaminated sites in Canada and the limited resources with which to clean them, it is important to choose the most appropriate and cost-effective strategy to rehabilitate the site.

Figure 1 The process of excavating and treating large tracts of contaminated land is very costly and disruptive.

There are four remediation strategies to consider when attempting to make contaminated land useful again:

1. Do nothing and restrict land use in the area.
2. Cover or encase the contaminated soil; for example, capping the area with concrete.
3. Excavate the entire site and replace the contaminated soil with clean soil. The contaminated soil is dumped elsewhere, in a landfill site licensed for such a purpose.
4. Treat the soil either at the site or off-site, and return the cleaned soil.

Treating Contaminated Soil

The remediation technologies used to treat contaminated soil can be classified as being physical, chemical, or biological.

PHYSICAL REMEDIATION

One example of a physical process to remediate soil is soil flushing. This involves pumping large volumes of a fluid through the soil in order to flush out and collect the contaminants. The fluid used for this process can be water, acids or bases, solvents, or detergents. The flushing fluid then must be captured and processed to remove the contaminant. A large volume of flushing fluid is required to clean a contaminated site. The remediation company must be particularly careful that none of the flushing fluid escapes into the environment.

CHEMICAL REMEDIATION

Three chemical methods used to remediate soil include stabilization and solidification, chemical oxidation, and electrolysis.

Stabilization and Solidification

This treatment method first uses chemical reactions to convert a contaminant into a safe form. For example, dissolved metal ions in groundwater can be precipitated by adding a strong base. The precipitate can then be collected, mixed with a binding agent such as concrete, and buried in the ground. Sometimes, the binding agent is pumped directly into the ground (**Figure 2**). This process prevents dissolved ions from moving with the groundwater and contaminating other water sources.

Figure 2 Stabilization and solidification involves using binding agents such as cement to help immobilize contaminants.

Electrolysis

Electrolysis is sometimes used to remove toxic metal ions dissolved in groundwater. This technology involves inserting electrical conductors (electrodes) into the ground and applying a voltage across them (**Figure 3**). This is similar to the use of electrolysis to purify metals described in Section 5.5.

Provided that the ground is moist, positively charged metal ions migrate to and attach onto the negatively charged electrode. The metal builds up on the electrode, and can later be safely removed and even recycled.

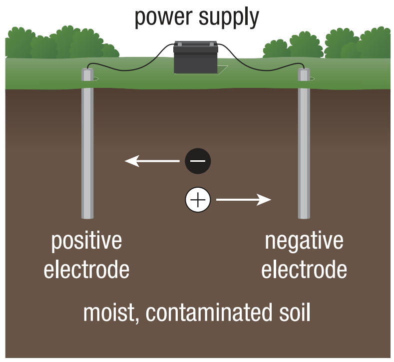


Figure 3 During electrolysis, positive metal ions flow toward the negative electrode. When they reach the electrode they acquire electrons and become metals.

Chemical Oxidation

The goal of chemical oxidation is to use chemical reactions to convert the contaminant into a less hazardous substance. The contaminant might be organic compounds, ionic compounds, or metal ions. The chemicals used for this process are worked into the ground using large drills or augers. Liquids such as hydrogen peroxide can be injected into the ground under pressure.

BIOREMEDIATION

Physical and chemical remediation technologies are often expensive, energy intensive, and disruptive, and not always effective. Fortunately, scientists have discovered that certain plants and micro-organisms can consume or absorb toxins from the environment. The use of living things to remove pollutants from a contaminated site is called **bioremediation**. In many cases, bioremediation is more cost-effective and less disruptive to the area than chemical or physical remediation methods. It may, however, take longer.

bioremediation the process of using organisms to treat contaminated land or water

LEARNING TIP

Bioremediation
"Bio" means living things.
"Remediation" means "to fix."
Bioremediation means using living things to fix the environment.

WEB LINK

To see videos, photos, and maps of the 2010 oil spill in the Gulf of Mexico,

GO TO NELSON SCIENCE

Figure 4 Airplanes sprayed huge volumes of dispersants onto the oil slick in the Gulf of Mexico in 2010.

The organisms used in bioremediation vary from tiny micro-organisms to large trees. The organism used depends on both the nature of the site and its contaminants. In some cases, it is appropriate to allow the micro-organisms that occur naturally in the soil to decompose the contaminants. For example, some soil micro-organisms naturally decompose organic pollutants, such as gasoline and oil. The products are relatively harmless, but include carbon dioxide, which is a greenhouse gas.

In other cases, nature requires a helping hand to allow natural decomposing processes to proceed. For example, during the oil spill in the Gulf of Mexico in 2010, chemical dispersants were used to break up the oil into small droplets. This helps naturally occurring micro-organisms to decompose the oil more quickly. You will learn more about the use of dispersants in Unit 4.

In other situations, micro-organisms can be pumped into the ground, where they find and decompose contaminants (**Figure 5**).

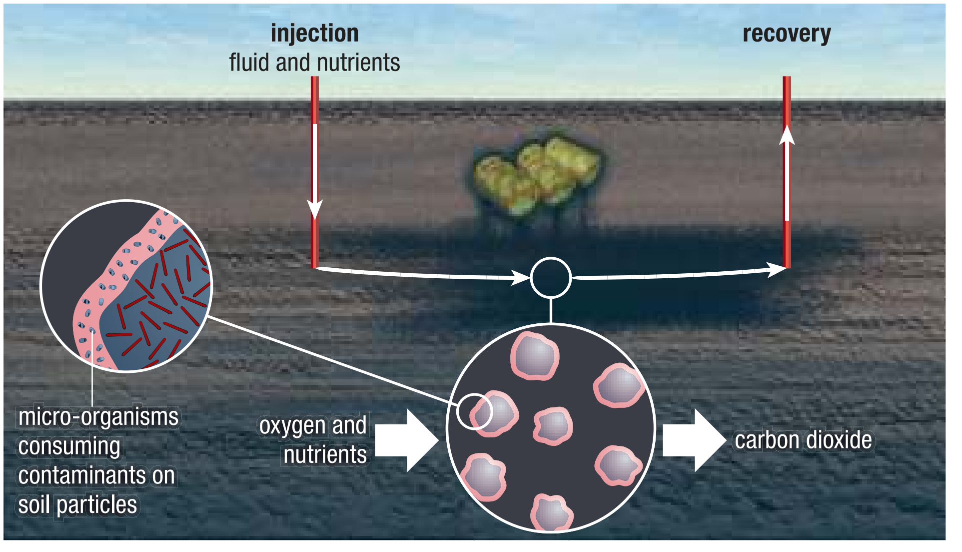


Figure 5 We can speed up bioremediation by pumping a mixture of liquid nutrients—and sometimes oxygen—into the soil.

Phytoremediation

Plants are surprisingly effective at cleaning up a broad range of pollutants. The use of plants to remediate soil or water is called **phytoremediation**. A healthy population of plants is a natural water purification system. Plants use their root systems to draw up groundwater and its contaminants. Some contaminants are decomposed by micro-organisms in the roots. Other contaminants are decomposed within the plant itself. The products are then either utilized by the plant or released to the atmosphere.

phytoremediation the process of using plants to treat contaminated land or water

Substances that cannot be degraded by the plant, such as metals, accumulate in the plant (**Figure 6**). Some plants, called "hyperaccumulators," have an unusually high ability to take in and trap metals. These plants can absorb as much as 1 % of their mass in metals. Sunflowers, geraniums, and Indian mustard are examples of hyperaccumulating plants. Metals can then be "mined" by harvesting the plants and burning them. The metals can be recovered chemically from the ashes of the plants.

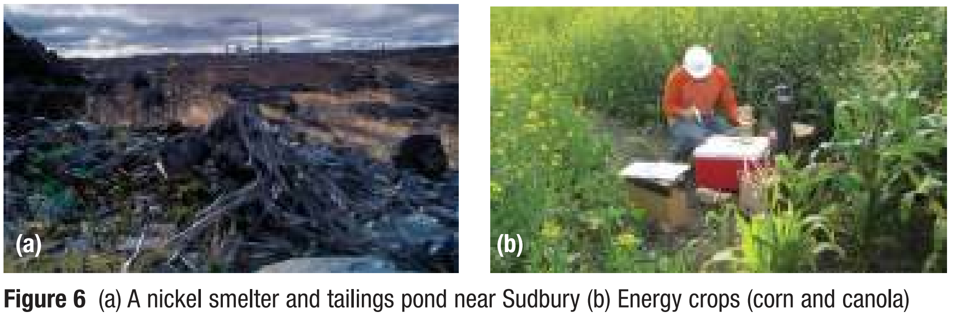


Figure 6 (a) A nickel smelter and tailings pond near Sudbury (b) Energy crops (corn and canola) grown on the tailings pond

Research This

Innovations in Bioremediation

Skills: Researching, Analyzing, Evaluating, Communicating

SKILLS HANDBOOK A5.1

In 2007, Natural Resources Canada initiated a project to remediate a tailings pond so that it could grow crops used to produce biofuels. A tailings pond is a mixture of water and waste solids from the processing of metal ores. Chemical reactions between oxygen and exposed minerals in the tailings often make the water very acidic. As a result, little can grow on or near a tailings pond.

The first phase of the project involved covering the pond with compostable waste from local paper mills. The natural basicity of this material helped to neutralize some of the acidity of the tailings. This material also provided nutrients for the growing plants. Covering the pond also slowed the rate at which oxygen could react with the minerals in the pond.

1. This example illustrates an innovative use of bioremediation and industrial waste to reclaim contaminated land. Research another example of an innovative use of bioremediation to solve an environmental problem.

A. Evaluate the effectiveness of the remediation technology. T/I

B. Is this technology a more appropriate way to rehabilitate the site than physical or chemical technologies? Summarize the pros and cons of each approach in a suitable graphic organizer. T/I C

GO TO NELSON SCIENCE

5.6 Summary

- Remediation is the process of removing contaminants from land or water so that there is no longer a threat to human health or the environment.
- Remediation technologies can be physical, chemical, or biological.
- Physical soil remediation includes flushing. Chemical remediation includes stabilization and solidification, oxidation, and electrolysis.
- Bioremediation uses living things to treat contaminated soil or water. Bioremediation can be less expensive, less disruptive, and sometimes more effective than physical or chemical remediation technologies.
- Phytoremediation is a specific type of bioremediation that uses plants to treat contaminated soil or water.

UNIT TASK BOOKMARK

You will use a form of chemical remediation in the Unit Task, described on page 242. You might find it useful to refer to the information in this section as you work on the Unit Task.

5.6 Questions

1. Use a particular example to show that the remediation and the total cleanup of a site are not necessarily the same. K/U
2. Which remediation strategy would you recommend if lead were discovered in the soil of an elementary school playground in August? Assume that relocating the school is not an option. Justify your choice. T/I A
3. Hydrogen peroxide is sometimes used to treat contaminants in the ground. As hydrogen peroxide reacts, it often releases a great deal of thermal energy. How do you think this helps decompose the contaminants? T/I A
4. What effect might covering a tailings pond with compostable material have on the quantity of acidic fluids draining from the pond? Why? K/U A
5. Some types of underground rock are fractured by many cracks. What effect do you think this characteristic has on the success of containing a chemical spill in the area? Why? K/U A
6. (a) In what ways is bioremediation more desirable than physical or chemical remediation technologies?
   (b) Suggest one limitation of bioremediation.
   (c) Compare the carbon footprint of bioremediation with that of physical or chemical remediation technologies. K/U T/I A
7. Thermal technologies are also sometimes used to treat contaminated soil. Research how this technology works and what type of contaminant it is best suited for. Summarize your research as a poster, an audiovisual presentation, or a web page. T/I C A

GO TO NELSON SCIENCE

5.7 Green Chemistry in Industry

Going "green" is often good for business. Many manufacturers have realized this as they look for new ways to become more profitable (**Figure 1**). Becoming "greener" involves changing a manufacturing process so that it is less harmful to the environment. Sometimes, this can be achieved simply by using less-toxic chemicals in the process. In other cases, major equipment upgrades are necessary. These changes usually make the process more efficient. A more efficient process can produce more finished goods using less energy and generating less waste. As a result of going green, companies become more profitable and competitive and the environment is safer and healthier. Of course, to really make a positive impact on the environment, this change must be supported by consumers. We will only see a benefit if people change their consumption habits and develop a less wasteful, more sustainable approach to consumer products.

Figure 1 The pulp and paper mill at Fort Frances, Ontario, has made great efforts to reduce its impact on the environment.

What Is Green Chemistry?

Manufacturing processes involving chemicals can be made greener by following the principles of green chemistry introduced earlier. Recall that the main idea behind green chemistry is: Why generate pollution if there is a greener (less polluting) alternative? A green manufacturing process is one in which each step is designed so that its environmental impact is minimal. The process is "benign by design."

Principles of Green Chemistry

In this section we will look at five of the major principles of green chemistry to see how each has been used to "green" a specific manufacturing process.

PRINCIPLE 1: USE RENEWABLE RESOURCES AS STARTING MATERIALS

Although many manufacturing processes are well established and accepted, it is always worth looking for a better starting material. The plastics industry is considering some new options.

Figure 2 Nylon—a synthetic fibre—is made industrially on a vast scale. Its manufacture uses chemicals derived from oil. Advances in the use of biocatalysts are helping to find renewable alternatives to oil-based raw materials.

The Raw Material for Making Plastic

Adipic acid is a key ingredient in the production of certain plastics, including nylon and polyurethane (**Figure 2**). Millions of tonnes of adipic acid are produced globally each year. The standard process for making this compound starts with an organic compound called benzene, C_6H_6. Benzene is anything but green! It is a known carcinogen (cancer-causing agent) and it is derived from oil, a non-renewable resource.

Recently, however, chemists have come up with a "sweet" substitute for benzene: glucose. Glucose is the sugar that is the body's main source of energy. Glucose can readily be made from renewable resources such as starch, produced by many food crops. The catalyst that makes possible this substitution of glucose for benzene is a genetically modified version of the *E. coli* bacteria. Using a food crop to make plastic is not necessary a "win-win" proposition. There is concern about the use of land for this purpose, and the possible increase in food costs. Some people are also concerned about the use of genetically modified organisms. Perhaps we should just use less plastic. "Going green" is not always simple: we must carefully consider any changes and try to predict any long-term effects.

PRINCIPLE 2: USE CHEMICALS THAT HAVE LESS ENVIRONMENTAL IMPACT

The manufacturing process may involve a step that is bad news for the environment. The polystyrene industry searched for a better option.

A Safer Expanding Agent

Polystyrene is used to manufacture foam packaging products such as fast-food containers and Styrofoam cups. These products are made by first injecting liquid polystyrene into a mould of the container. A gas is then injected which expands the polystyrene so that it takes the shape of the mould. In the twentieth century, chlorofluorocarbons, or CFCs, were used as the expanding gas. However, once CFCs escape into the atmosphere, they react with ozone, $O_3(g)$. Any loss of ozone in the upper atmosphere is a concern because ozone absorbs much of the harmful ultraviolet radiation travelling to Earth from the Sun. The wide-scale use—and release—of CFCs caused a significant thinning of the ozone layer.

In the 1990s, chemists at the Dow Chemical Company developed a process that uses carbon dioxide as the expanding agent, instead of CFCs. This compound is an ideal alternative because it is non-polluting and cheaper than CFCs. Because carbon dioxide is not flammable, it is also safer for workers involved in the process. Of course, carbon dioxide is a greenhouse gas, so the process is not entirely benign.

Mini **Investigation**

Shrinking Styrofoam (Teacher Demonstration)

Skills: Observing, Analyzing, Communicating — SKILLS HANDBOOK A1.1, A1.2

Styrofoam is an example of expanded polystyrene. A Styrofoam cup feels light because it contains millions of pockets of air. These pockets of air are surrounded by thin walls of polystyrene. Propanone (often known as acetone) breaks down the bonds within polystyrene, causing the air pockets—and the cup—to collapse. As it does, air trapped within the structure of the cup is released.

Equipment and Materials: chemical safety goggles; lab apron; 1 L beaker; metal spoon; propanone; Styrofoam cup

Propanone is highly flammable. It should not be used anywhere near open flames or other sources of ignition.

Propanone is toxic by inhalation. Use it only in a well-ventilated area. Avoid inhaling the fumes.

Propanone is an irritant and can cause eye damage. The investigator should wear chemically resistant gloves.

1. Put on your chemical safety goggles and lab apron.
2. Your teacher will pour about 40 mL of propanone into a 1 L beaker (to a depth of about 1 cm).
3. A Styrofoam cup will be placed in the propanone and gently pushed down.
4. The polystyrene will be collected with the spoon. The propanone can then be reused.

A. Where did the bubbles come from? T/I

B. Based on your observation, why do you think it is difficult to recycle Styrofoam economically? T/I

PRINCIPLE 3: USE CATALYSTS TO INCREASE REACTION EFFICIENCY

As you learned in Section 4.1, a catalyst makes a chemical reaction occur faster without being consumed in the reaction. The advantage of a catalyst in a manufacturing process is that it can be used over and over again. Sulfuric acid is an example of an important chemical made using a catalyst: vanadium pentoxide.

Speeding Up the Production of Sulfuric Acid

The process used to make sulfuric acid is called the contact process. Pure sulfur is converted into sulfuric acid in a three-step process (**Figure 3**). A key step in the process is the conversion of sulfur dioxide into sulfur trioxide:

$$2\ SO_2(g) + O_2(g) \rightarrow 2\ SO_3(g)$$

Normally, this reaction occurs very slowly. It occurs much faster if the sulfur dioxide and oxygen flow over a vanadium pentoxide, V_2O_5, catalyst.

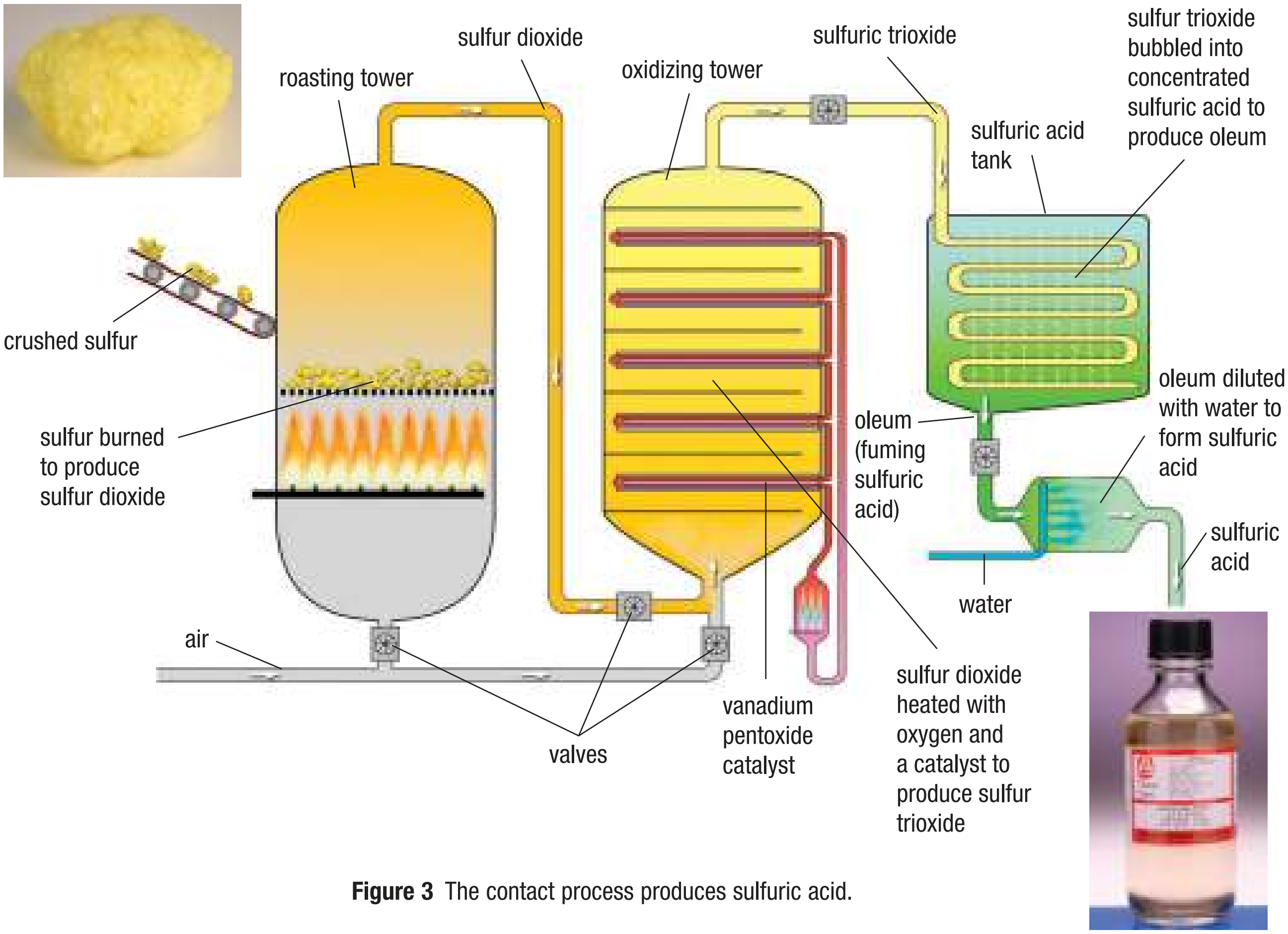


Figure 3 The contact process produces sulfuric acid.

PRINCIPLE 4: USE LESS ENERGY

Many manufacturing industries are huge energy users. As the cost of energy—both financial and environmental—increases, industries are trying to find ways to cut back.

Developing a More Efficient Blast Furnace for Steel

Manufacturing steel is a particularly energy-intensive industry. Keeping energy costs down is therefore very important. One way to reduce energy costs is to update aging equipment.

Switching to a more efficient blast furnace quickly helps to drive down energy costs (**Figure 4**). Furthermore, since smaller quantities of fossil fuels are burned to generate energy, the furnace upgrade can also result in a reduction of air pollutants (such as dust and acidic compounds) and greenhouse gases. A new blast furnace may initially cost tens of millions of dollars. However, this expense is soon recovered in increased productivity, decreased energy costs, and decreased costs associated with operating pollution control equipment.

A company can also reduce energy consumption by changing the process itself. For example, in Section 5.5 you learned that ore was traditionally smelted in a furnace fuelled by the combustion of fossil fuels. Much of the energy required for flash smelting, however, is released by the chemical reactions themselves. In effect, flash smelting fuels itself, resulting in significant energy savings and reduced emissions.

Figure 4 If a blast furnace can be made more efficient, it will cost less to run.

PRINCIPLE 5: WASTE NOT

In Section 5.4, you learned that the lead and sulfuric acid in car batteries are hazardous. A Canadian battery recycler has developed a recycling process that is totally enclosed. Old batteries go in; clean recycled materials come out. There are only a few places in the process—like the vats of molten lead—that are open to the atmosphere. These vats have exhaust hoods that collect dust and vapour and return them to the process (**Figure 5**). This process is described as being a "closed loop" because it prevents hazardous substances from escaping into the environment. The result is less waste, safer working conditions, and a cleaner environment. This leads to increased worker productivity and reduced waste cleanup costs.

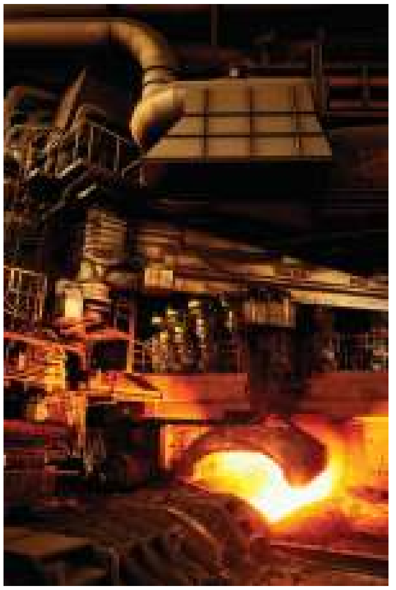


Figure 5 An exhaust hood (near the top of this photo) prevents any vapours or dust from leaving the process.

Applying the Five Principles to Pulp and Paper

Pulp and paper production is one of Canada's major industries, directly employing several thousand people in nine of our ten provinces. Canada is the world's fourth-largest manufacturer of pulp and paper and the largest supplier of newsprint.

Like the mining and mineral processing industry, the pulp and paper industry has a history of being a major source of both air and water pollution (**Figure 6**). We will now look at how paper is made and some of the steps that the industry is taking to become both environmentally responsible and more profitable.

Figure 6 A pulp mill at Port Edward, B.C.

The Paper-Making Process

Canada is a world leader in the pulp and paper industry. For decades, mills have been turning wood pulp into many different kinds of paper (**Figure 7**). The paper industry traditionally harvested trees from virgin forests. This practice destroyed vast areas of Canada's forest ecosystems. Many mills are now using specially planted, fast-growing trees to supply the needed fibre.

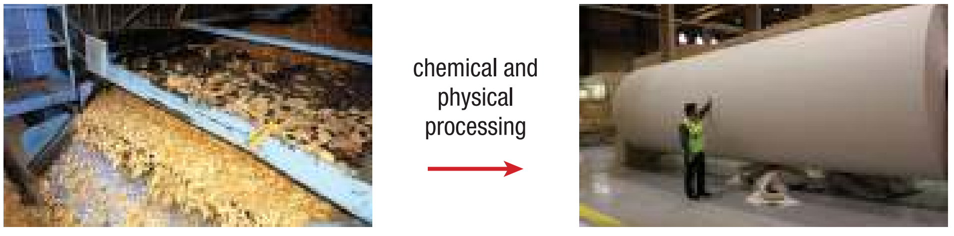


Figure 7 There are many steps in the the paper-making process.

About 70 % of dry wood is made up of very long molecules of a compound called cellulose. The remainder of wood is a substance called lignin. Lignin can be thought of as the glue that holds the cellulose strands together. Lignin is brownish in colour, so if it is not removed it gives paper an undesirable brown discoloration. To produce white paper, lignin must be removed before the pulp is bleached (**Figure 8**).

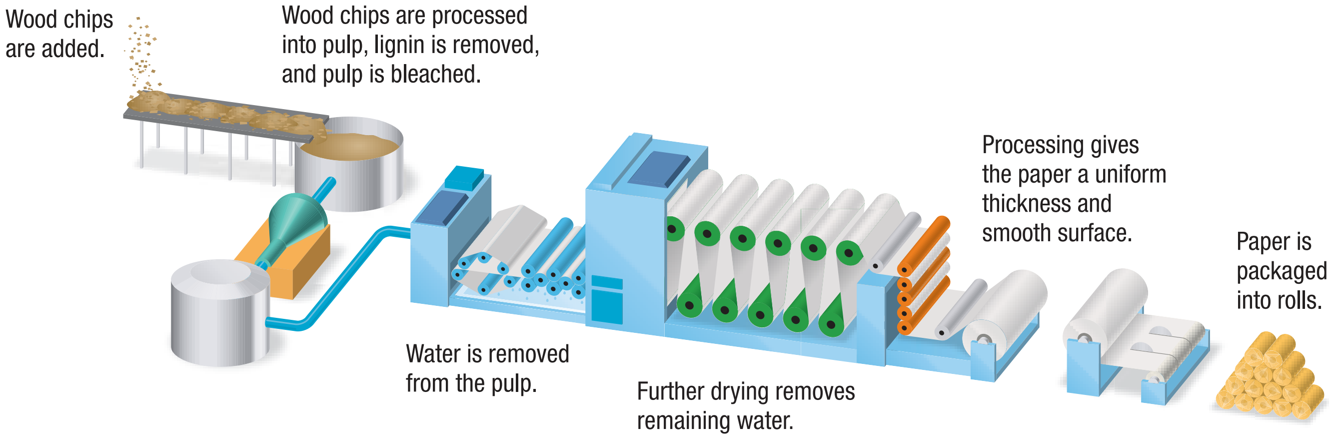


Figure 8 Making paper

The most common method of removing lignin is called the Kraft process. In the first step of the process, wood is chopped into small chips. The wood chips are then cooked in a vat containing sodium sulfide and sodium hydroxide. The chemical reactions that occur in the vat remove up to 90 % of the lignin from the wood chips. Modern pulp mills recover and reuse almost all unused chemicals from this stage of the process.

Figure 9 Chemical bleaches are used to whiten wood pulp. Would you be willing to use paper products—tissues, perhaps—if they were less white?

CAREER **LINK**

The pulp and paper industry employs teams of pulp and paper research chemists. If you are interested in finding out more about this field,

GO TO NELSON SCIENCE

BLEACHING ALTERNATIVES

To remove the remaining lignin, paper manufacturers use some kind of bleach (**Figure 9**). In the past, elemental chlorine, Cl_2, was used. However, this created serious environmental problems. The reaction of chlorine with lignin produces a group of highly toxic chlorinated molecular compounds called dioxins. Because of this problem, chlorine dioxide, ClO_2, has largely replaced chlorine. This change has reduced dioxin production by almost 90 %.

The only way to completely eliminate the production of dioxin would be to switch to non–chlorine-based bleaches. Hydrogen peroxide, which is the bleach used in hair-colouring products, has been tried with limited success. Hydrogen peroxide bleaching takes longer and requires higher reaction temperatures. Recently, however, research engineers have developed a catalyst called a TAML activator to overcome these limitations. The catalyst enables hydrogen peroxide to bleach pulp in a shorter period of time and at a lower temperature. Since it is a catalyst, TAML activator can be reused.

Despite these advances, chlorine dioxide remains the dominant bleach in the industry.

ENERGY USE

Making pulp and paper is an energy-intensive process. Consequently, paper manufacturers are continually looking for ways in which to save energy. Some of their energy savings result from upgrades to more efficient equipment. For example, more efficient pumps and drying equipment help reduce energy costs significantly. Energy savings are also achieved by using alternative energy sources. Burning wood waste (such as bark and solids collected from the pulping process) instead of fossil fuels is one way to reduce fuel costs.

CONTROL OF EMISSIONS AND WASTE

The complete elimination of waste emissions from pulp and paper mills is not yet possible. However, the mills have made significant improvements to emissions through the use of both internal and external control measures. Internal measures include using more closed-loop systems to prevent materials from escaping into the environment. External measures include capturing and treating gas emissions using scrubbing technology.

Scrubbers

Pulp mill emissions include a number of toxic gases, such as hydrogen sulfide, chlorine dioxide, and sulfur dioxide. A chemical scrubber traps and treats these emissions, converting them into safer substances before they leave the mill. Inside the scrubber, the pollutant gases are showered with a fine paste-like mixture of water and limestone. Sulfur dioxide reacts with calcium carbonate to form calcium sulfite. Calcium sulfite sinks to the bottom of the scrubber, where it can be removed. The waste calcium sulfite can then be converted into gypsum, which is used to make drywall.

5.7 Summary

- A manufacturing process that uses chemicals can be both profitable and environmentally responsible.
- The main theme of green chemistry is: Why generate pollution if there is a greener alternative?
- The principles of green chemistry are being used to design effective industrial processes that are environmentally responsible.

5.7 Questions

1. What can a chemical manufacturer do to improve the efficiency of a chemical process? K/U
2. Identify two ways in which the use of a catalyst can make a manufacturing process greener. K/U
3. Explain how an investment in pollution control equipment can indirectly help a manufacturer become more profitable. A
4. How does a reduction in energy consumption affect a manufacturer's greenhouse gas emissions? K/U A
5. (a) What is meant when a manufacturing process is described as being "benign by design"?
   (b) Describe how you could alter a process in your home to make it more "benign by design." K/U A
6. (a) Refer to Figure 3 to explain why the contact process is an example of a closed-loop manufacturing process.
   (b) Identify two advantages of designing a process in this way. K/U A
7. Many of the raw materials used to manufacture plastics come from crude oil. Recently, chemists have discovered how to use chemicals that occur naturally in corn to manufacture polylactic acid, a plastic used to make packaging materials. List two advantages and one disadvantage of using food crops like corn as a source of these raw materials. A
8. A great deal of water is required to manufacture pulp and paper. Why is it environmentally and economically important for pulp mills to recycle as much of their water as possible? A
9. Identify examples of how the pulp and paper industry applies the five principles of green chemistry introduced in this section. K/U A
10. The area around a pulp mill often has a very distinctive odour. Research the cause of this odour and what can be done about it. Summarize your findings in a paragraph. T/I C

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Investigation 5.1.1 CONTROLLED EXPERIMENT

SKILLS MENU

- Questioning
- Researching
- Hypothesizing
- Predicting
- Planning
- Controlling Variables
- Performing
- Observing
- Analyzing
- Evaluating
- Communicating

Combustion of Ethyne

A finely tuned car engine burns fuel efficiently and releases the least soot and unburned hydrocarbons. The air-to-fuel ratio is a critical factor in determining how efficiently a fuel burns. In this experiment, you will be burning three different air/fuel mixtures to see which gives the cleanest (least sooty) flame. The fuel you will be burning is ethyne, $C_2H_2(g)$, more commonly known as acetylene. This combustible gas can be produced through the reaction of calcium carbide, CaC_2, with water (**Figure 1**). The chemical equation for this reaction is

$$CaC_2(s) + 2\ H_2O(l) \rightarrow C_2H_2(g) + Ca(OH)_2(aq)$$

Figure 1 Calcium carbide reacts with water. Ethyne is less dense than air. This means that ethyne rises quickly when it is released in the air, just like a helium-filled balloon.

This investigation should be carried out in a fume hood. If a fume hood is not available, the investigation should be performed only by a teacher in a well-ventilated science laboratory.

Testable Question

Which air/ethyne mixture in **Table 1** results in the cleanest (least sooty) combustion?

Table 1 Three Air/Ethyne Mixtures

Mixture	Fraction of air in mixture	Fraction of ethyne in mixture
1	1/4	3/4
2	1/2	1/2
3	9/10	1/10

Hypothesis/Prediction

SKILLS HANDBOOK A2.2

Predict which air/ethyne mixture will result in the cleanest combustion. Give reasons for your hypothesis.

Variables

Identify all major variables that will be manipulated, measured, and/or controlled in this investigation. What is the responding variable?

Experimental Design

You will collect ethyne, from the reaction of calcium carbide and water, by displacement of water. You will prepare and burn three different mixtures of ethyne gas and air. You will compare the flames produced, and establish which mixture burns with the cleanest flame.

Equipment and Materials

SKILLS HANDBOOK A1.1, A1.2

- chemical safety goggles
- lab apron
- safety gloves
- 400 mL beaker
- 3 labelled test tubes with stoppers
- test-tube rack
- tongs (for handling calcium carbide)
- small dry dish (for carrying calcium carbide)
- test-tube tongs
- candle in a holder
- 1–2 pellets calcium carbide, CaC_2
- splint
- dropper bottle of phenolphthalein indicator
- 5 mL limewater, $Ca(OH)_2(aq)$

This experiment involves open flames. Tie back long hair and secure loose clothing.

Calcium carbide reacts vigorously with water to produce ethyne, a flammable gas. Keep calcium carbide away from water unless it is being used in a fume cupboard to produce ethyne. Avoid inhaling calcium carbide dust, or getting it in your eyes or on your skin. Only small quantities of calcium carbide should be in the room at any time. All calcium carbide must be reacted before the disposal of any liquids.

Phenolphthalein indicator solution is flammable. Keep it away from open flames. Work in a fume hood.
Wash your hands after completing this investigation.

Procedure

Part A: Collecting Ethyne

SKILLS HANDBOOK A1.2, A2.2

1. Put on your safety goggles, lab apron, and safety gloves.

2. Pour about 100 mL of tap water into a 400 mL beaker.
3. Fill three test tubes (labelled 1, 2, and 3) with water, then invert them into the beaker.
4. Using tongs, place a small piece of calcium carbide, CaC_2, into the beaker. Move test tube #1 so that it is over the calcium carbide (**Figure 2**). Collect ethyne in the test tubes as outlined in **Table 2**.

Figure 2 Collect the gas produced when calcium carbide reacts with water.

Table 2 Volume of Ethyne in Each Test Tube

Test tube	Volume of ethyne
1	three-quarters full
2	half full
3	about one-tenth full

5. Holding the test tubes so that the open ends always point down, remove each test tube from the beaker. Allow the water to drain out and be replaced by air. Stopper and shake each mixture to ensure that the gases are thoroughly mixed. Place the test tubes, stoppered ends down, in the test-tube rack.
6. Leave the beaker and its contents in the fume hood and allow any remaining calcium carbide to continue to bubble.

Part B: Burning Ethyne

7. You should now have three stoppered test tubes containing different mixtures of ethyne and air. Light a splint from your candle. Use the burning splint to test each mixture, one at a time. Keep the test tubes tilted slightly and inverted at all times (**Figure 3**). Record your observations for each.

Ethyne (acetylene) is flammable. Collect and ignite only small volumes. Work in an empty fume hood. Tie back long hair and secure loose clothing.

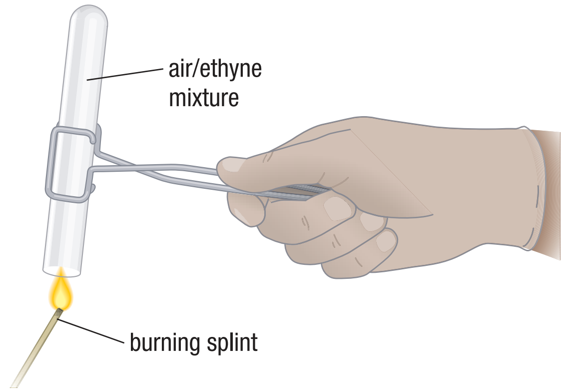


Figure 3 Test the gas mixtures with a burning splint.

Part C: Testing for Other Reaction Products

8. Plan a procedure to test for
   (i) hydroxide ions in the beaker after the reaction of calcium carbide with water
   (ii) the presence of carbon dioxide from the combustion of ethyne.
9. Proceed with the test once your teacher has approved your procedure.
10. Follow your teacher's instructions for the disposal of the remaining materials. Wash your hands.

Analyze and Evaluate

(a) Identify the major variables that you measured and/or controlled in this investigation. Was there a causal relationship between any of the variables? Explain. T/I

(b) If you can, answer the Testable Question. Support your answer with evidence. T/I

(c) Why was it necessary to keep the test tubes sealed or inverted prior to burning the gas mixtures? T/I

(d) What observation suggests that hydroxide ions are produced when calcium carbide reacts with water? T/I

(e) What observation suggests that carbon dioxide is a product of the combustion of ethyne? T/I

Apply and Extend

(f) A "rich" fuel mixture contains more fuel than the same volume of a "lean" fuel mixture. Which type of mixture gives the cleanest flames? Support your answer with experimental evidence. T/I A

(g) A gasoline engine becomes "flooded" and difficult to start if too much gasoline is initially pumped into it. Predict the appearance of the exhaust from this engine once it does start. A

Properties of Oxides

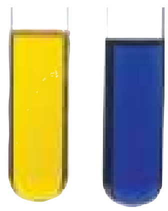
SKILLS MENU

- Questioning
- Researching
- Hypothesizing
- Predicting
- Planning
- Controlling Variables
- Performing
- Observing
- Analyzing
- Evaluating
- Communicating

Most of the elements of the periodic table react with oxygen to form an oxide. Some oxides change the pH of water as they dissolve. Oxides that make water acidic are called acidic oxides. Oxides that make water basic are called basic oxides. Acidic oxides are largely responsible for the environmental problems caused by acid precipitation. In this investigation, you will look for patterns in the properties of typical metallic and non-metallic oxides.

Bromothymol blue and phenolphthalein indicators will be used to determine whether the resulting mixtures are acidic or basic. Their colours are summarized in **Table 1**.

Table 1 Colours in Indicators in Acidic and Basic Solutions

Indicator	Colour in acid		Colour in base
phenolphthalein	colourless		magenta
bromothymol blue	yellow		blue

Some of the oxides will be provided; others will be made. Magnesium oxide is made by burning magnesium metal. Carbon dioxide is produced using the reaction of ethanoic acid, $HC_2H_3O_2(aq)$, and sodium hydrogen carbonate, $NaHCO_3(s)$. The combustion of the chemicals in the head of a match is the source of oxides of both phosphorus and sulfur. The chemical equation for this reaction, which also includes potassium chlorate, is

$$P_4(s) + 5\,O_2(g) + 3\,S(s) + 2\,KClO_3(s) \rightarrow P_4O_{10}(g) + 3\,SO_2(g) + 2\,KCl(s) + \text{energy}$$

Purpose

To compare the properties of metallic and non-metallic oxides

Equipment and Materials

SKILLS HANDBOOK A1.1, A1.2

- chemical safety goggles
- lab apron
- safety shield with UV protection
- Bunsen burner clamped to a retort stand
- spark lighter
- tongs
- ceramic dish
- scoopula
- well plate
- 25 mL graduated cylinder
- two 125 mL Erlenmeyer flasks
- strip of magnesium, Mg(s)
- vials of
  - calcium oxide, CaO(s)
  - magnesium oxide, MgO(s)
  - sodium hydrogen carbonate, $NaHCO_3(s)$
- dropper bottles of
  - phenolphthalein indicator solution
  - bromothymol blue indicator
- small beaker containing 20 mL dilute ethanoic acid, $HC_2H_3O_2(aq)$ (0.1 mol/L)
- cotton swab
- matches
- conductivity tester

This activity involves open flames. Tie back long hair and secure loose clothing.

Phenolphthalein indicator solution is flammable. Keep it away from open flames.

Procedure

SKILLS HANDBOOK A1.1, A1.2, A2.4

1. Create a table in which to record your observations. List the metallic oxides first, followed by the non-metallic oxides.

Part A: Synthesis of Magnesium Oxide (Teacher Demonstration)

This portion of the investigation should only be performed by a teacher. It should be done in a fume hood with a UV safety shield in place.

The flame produced by burning magnesium is very bright and produces UV light, which can permanently damage your eyes. Do not look directly at it.

Do not attempt to extinguish a magnesium fire with either a carbon dioxide fire extinguisher or a water fire extinguisher. Instead, smother the fire with sand.

2. Put on your chemical safety goggles and lab apron.

3. Working behind the safety shield, your teacher will use tongs to hold a 5 cm strip of magnesium.
4. The magnesium will be ignited in a Bunsen burner flame (**Figure 1**).

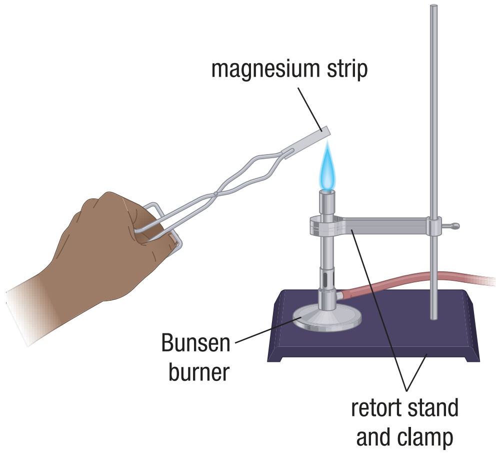


Figure 1 Your teacher will burn magnesium metal to make magnesium oxide.

5. Once the magnesium ignites, it will be removed from the flame and held over a ceramic dish. As it continues to burn, the reaction product will be allowed to fall into the dish (**Figure 2**).

Figure 2 Burning magnesium emits a light that is bright enough to damage your eyes. You should only look at the flame through a UV safety shield.

Part B: Properties of Metallic Oxides

6. Add a small quantity (as much as will fit on the end of a scoopula) of the provided metallic oxides (calcium and magnesium) to two separate wells in the well plate.
7. Add two drops of phenolphthalein indicator to each of the metallic oxides. Record your observations.
8. Repeat Steps 6 and 7 using fresh samples of the metallic oxides and bromothymol blue indicator.

Part C: Properties of Non-Metallic Oxides

Preparation and testing of carbon dioxide

9. Add about 20 mL of distilled water to an Erlenmeyer flask.
10. Add 4 drops of bromothymol blue indicator to the flask.
11. Plan a procedure, using the materials provided, that produces carbon dioxide gas. Proceed once your teacher has approved your plan.
12. Test the gas by pouring it into the flask containing the indicator solution. Record your observations.

Preparation and testing of sulfur dioxide (Teacher Demonstration)

This portion of the investigation should be performed by the teacher in a fume hood or well-ventilated lab.

13. Your teacher will drip bromothymol blue indicator onto one end of the cotton swab until the end is saturated.
14. Your teacher will light a match and hold the saturated end of the cotton swab above the match for about 5 s so that the combustion gases from the flame can reach the cotton swab. The cotton swab will not be allowed to ignite.
15. Examine the portion of the cotton swab that was directly above the flame for colour changes. Record your observations.

Analyze and Evaluate

(a) Referring to your observations table, summarize the properties of the metallic and non-metallic oxides tested. C

(b) Classify the oxides as either acidic oxides or basic oxides, according to your evidence. K/U T/I

(c) In Step 12, you poured carbon dioxide from one flask into another. What does this suggest about the density of carbon dioxide compared to the density of air? T/I

(d) Predict the colour change that would be observed if carbon dioxide gas were bubbled into a mixture of calcium oxide, bromothymol blue, and water. Justify your prediction. K/U T/I

Apply and Extend

(e) Write chemical equations that explain how magnesium oxide and carbon dioxide, when they are added to water, form acids or bases. T/I C

(f) Classify the following compounds as either acidic or basic oxides. Justify your prediction. K/U T/I

(i) Na_2O (ii) P_2O_5 (iii) As_2O_3

(g) A burning match produces $P_4O_{10}(g)$ and $SO_2(g)$ as well as combustion gases from the burning of wood. Design an investigation that you could do to rule out that burning wood was responsible for most of the colour change observed in Step 15. T/I

Neutralizing Pain Relievers

Aspirin is one of the world's most widely used drugs. Originally developed as a pain reliever, Aspirin now has a number of important medical applications. For example, Aspirin has been shown to help prevent the formation of blood clots. It can be given to someone having a heart attack, or to prevent thromboses during long plane trips.

The active ingredient in Aspirin is acetylsalicylic acid, $HC_9H_7O_4$. This compound produces small quantities of hydrogen ions in solution, making the solution mildly acidic. Due to this acidity, frequent ingestion of Aspirin can irritate the stomach lining (**Figure 1**). To remedy this problem, some types of Aspirin contain compounds that neutralize some of the acidity of acetylsalicylic acid. These products are advertised as being acid-reduced or buffered.

Figure 1 The white area in this image is an ulcer. The photograph was taken using a tiny camera, on a device called an endoscope, that is inserted into the stomach through the mouth.

In this investigation, you will compare regular Aspirin to buffered Aspirin to see if the buffered product actually is acid-reduced, as it claims to be. To make your comparison, you will neutralize samples of regular and buffered Aspirin using a sodium hydroxide solution. The chemical equation for this reaction is

$$HC_9H_7O_4(aq) + NaOH(aq) \rightarrow H_2O(l) + NaC_9H_7O_4(aq)$$

The colour change of phenolphthalein will be used to indicate when neutralization is complete. An acidic solution has been neutralized when the next drop of sodium hydroxide permanently changes the indicator from colourless to pink.

SKILLS MENU

• Questioning	• Planning	• Observing
• Researching	• Controlling Variables	• Analyzing
• Hypothesizing	• Performing	• Evaluating
• Predicting		• Communicating

Purpose

To determine whether acid-reduced Aspirin contains less acid than regular Aspirin

Equipment and Materials

SKILLS HANDBOOK A1.1, A1.2

- chemical safety goggles
- lab apron
- 25 mL graduated cylinder
- two 125 mL Erlenmeyer flasks
- mortar and pestle
- dropper bottle of
  - phenolphthalein indicator solution
  - dilute sodium hydroxide solution, NaOH(aq) (0.1 mol/L)
- a regular Aspirin tablet
- a buffered Aspirin tablet

Phenolphthalein indicator solution is flammable. Keep it away from open flames.

Sodium hydroxide is corrosive. It can cause blindness if it comes in contact with the eyes. Wash any spills on skin or clothing immediately with plenty of cold water. Report any spills to your teacher.

Procedure

SKILLS HANDBOOK A1.1, A1.2, A2.4

1. Put on your chemical safety goggles and lab apron.
2. Use the graduated cylinder to add 20 mL of warm tap water to an Erlenmeyer flask.
3. Add five drops of phenolphthalein indicator to the flask.
4. Use a mortar and pestle to crush a regular Aspirin tablet. Transfer the crushed tablet to the flask.
5. Repeat Steps 2 to 4 with the second Erlenmeyer flask and the buffered Aspirin tablet.
6. Add one drop of sodium hydroxide solution to each flask and swirl. Note any change in colour.
7. Continue adding sodium hydroxide to each flask, one drop at a time, until the contents of one flask change colour permanently. Record your observations.
8. Follow your teacher's directions for disposal and cleanup.

Analyze and Evaluate

(a) Suggest an explanation for your observations. T/I

(b) Based on your observations, which type of Aspirin might be a better choice for someone who must take Aspirin on a regular basis? Why? T/I A

(c) Explain the role of phenolphthalein in this investigation. K/U T/I

(d) Why do you think it was necessary to crush the tablets and use warm water in this investigation? T/I

Apply and Extend

(e) Modify the procedure of this investigation so that it can be used to verify that a sample of extra-strength Aspirin contains more pain reliever than regular-strength Aspirin. T/I

(f) Ulcers caused by ASA (Aspirin) are often called peptic ulcers. Research peptic ulcers to find out how common they are, who is most susceptible, and how they are treated. Summarize your findings in a pamphlet, poster, or informative web page. T/I C A

GO TO NELSON SCIENCE

Summary Questions

1. Create a study guide based on the points listed in the margin on page 190. For each point, create three or four sub-points that provide relevant examples, explanatory diagrams, or general equations.
2. Look back at the Starting Points questions on page 190. Answer these questions using what you have learned in this chapter.
3. Use two specific examples to show how adopting the principles of green chemistry creates a "win-win" situation for industry and the environment. Why do you think more manufacturers have not adopted these principles?
4. Select a particular chemical process mentioned in this chapter. Assess the effectiveness of your chosen process in addressing an important health or environmental need.

Vocabulary

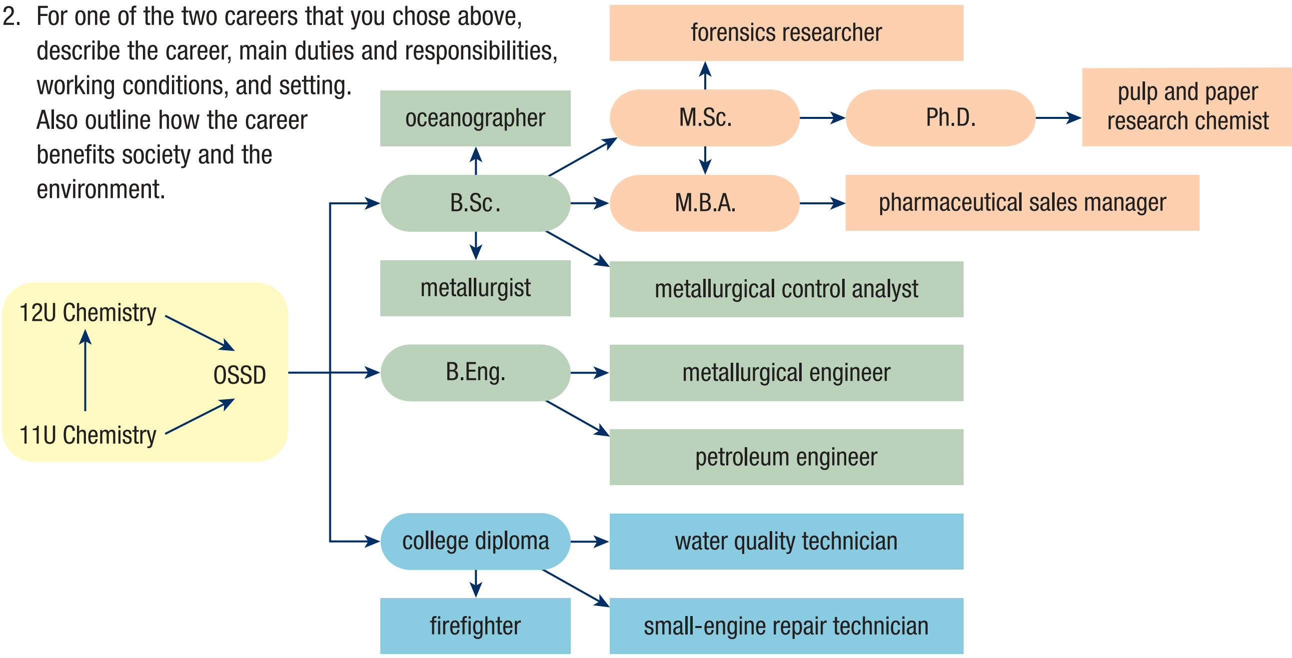
combustion (p. 192)
greenhouse gas (p. 192)
organic compound (p. 192)
air pollution (p. 192)
complete combustion of a hydrocarbon (p. 193)
incomplete combustion of a hydrocarbon (p. 193)
oxide (p. 200)
acid (p. 200)
base (p. 200)
acidic oxide (p. 200)
basic oxide (p. 200)
mineral (p. 212)
ore (p. 212)
metallurgy (p. 212)
smelting (p. 214)
flash smelting (p. 214)
remediation (p. 218)
bioremediation (p. 219)
phytoremediation (p. 220)

CAREER PATHWAYS

SKILLS HANDBOOK A7

Grade 11 Chemistry can lead to a wide range of careers. Some require a college diploma or a B.Sc. degree. Others require specialized or post-graduate degrees. Here are just a few pathways to careers mentioned in this chapter.

1. Select two careers related to chemical processes that you find interesting. Research the educational pathways that you would need to follow to pursue these careers. What is involved in the required educational programs? Prepare a brief report on your findings.
2. For one of the two careers that you chose above, describe the career, main duties and responsibilities, working conditions, and setting. Also outline how the career benefits society and the environment.



GO TO NELSON SCIENCE

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

For each question, select the best answer from the four alternatives.

1. Which is the formula of the hydrocarbon shown in the chemical equation below? (5.1) K/U
   $2\ C_xH_y(g) + 9\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(g)$
   (a) $C_2H_4(g)$
   (b) $C_2H_6(g)$
   (c) $C_3H_4(g)$
   (d) $C_3H_6(g)$
2. If a car's exhaust system has a hole that allows exhaust gases to leak into the passenger compartment, the biggest danger when the car is running would be
   (a) unburned hydrocarbons
   (b) carbon monoxide
   (c) carbon dioxide
   (d) soot (5.1) K/U
3. Which would be present in carbonated pop? (5.3) K/U
   (a) $CO(aq)$
   (b) $CO_3(aq)$
   (c) $H^+(aq)$ and $HCO_3^-(aq)$
   (d) $C^{2-}(aq)$ and $OH^-(aq)$
4. Which shows the general equation for a neutralization reaction that involves a hydroxide compound? (5.4) K/U
   (a) acid + ionic compound → water + base
   (b) water + ionic compound → acid + base
   (c) acid + base → water + ionic compound
   (d) water + base → acid + ionic compound
5. During the refining of an ore, a detergent-like ionic compound is used to concentrate the part of the ore containing the metal. What is this process called? (5.5) K/U
   (a) mining
   (b) flotation
   (c) smelting
   (d) flash smelting
6. Which best reduces the impact of mining on the environment? (5.5) K/U
   (a) dams and government regulation
   (b) improved technology and native species
   (c) air pollution and water pollution
   (d) government regulation and improved technology
7. Which of the following terms refers to using living things to remove pollutants from a contaminated site? (5.6) K/U
   (a) bioremediation
   (b) electrolysis
   (c) solidification
   (d) oxidation
8. Which toxic hydrocarbon can be replaced by glucose to make the synthesis of adipic acid a greener process? (5.7) K/U
   (a) chlorine dioxide
   (b) benzene
   (c) sulfuric acid
   (d) sodium sulfide

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

9. If a cloud of black smoke pours from the tailpipe of a running car, it is likely that the air/fuel mixture is too lean. (5.1) K/U
10. Adding sodium oxide to water produces sodium hydroxide, an acid. (5.3) K/U
11. If an oxide is introduced into water and the pH changes from 7 to 9, the oxide most likely contains a metallic element. (5.3) T/I
12. The catalytic converters in newer-model cars are designed to convert nitrogen monoxide to nitric acid and oxygen. (5.3) K/U
13. The difference between a neutralization reaction involving a hydroxide compound and a neutralization reaction involving a carbonate compound is the formation of water as a product. (5.4) K/U
14. Sulfur dioxide is a common air pollutant produced during the refining of metals. (5.5) K/U
15. Electrolysis is a type of physical remediation used to remove toxic metal ions that are dissolved in groundwater. (5.6) K/U
16. Remediation involves the complete removal of hazardous substances from the environment. (5.6) K/U
17. Phytoremediation involves cleaning up contamination by using plants. (5.6) K/U
18. About 70 % of dry wood is made up of cellulose, which is composed of very long molecules. (5.7) K/U

To do an online self-quiz,
GO TO NELSON SCIENCE

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

Knowledge

For each question, select the best answer from the four alternatives.

1. Which chemical equation shows a reaction that would produce a yellow flame? (5.1) K/U
   (a) $C_2H_4(g) + 2\ O_2(g) \rightarrow 2\ CO(g) + 2\ H_2O(g)$
   (b) $2\ C_2H_4(g) + 7\ O_2(g) \rightarrow 2\ CO_2(g) + 2\ CO(g) + 8\ OH(g)$
   (c) $3\ C_2H_4(g) + 6\ O_2(g) \rightarrow 2\ C(s) + 2\ CO_2(g) + 2\ CO(g) + 6\ H_2O(g)$
   (d) $C_2H_4(g) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 2\ H_2O(g)$
2. The balanced equation for a complete combustion reaction is shown below. Which is the formula of the hydrocarbon? (5.1) K/U
   $C_xH_y + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$
   (a) C_2H_2 (c) C_6H_4
   (b) C_3H_8 (d) C_8H_3
3. Carbon monoxide gas acts as a poison by
   (a) binding to substances in red blood cells
   (b) rupturing lung cells
   (c) paralyzing muscles
   (d) creating acid in the bloodstream (5.1) K/U
4. Which is a chemical formula for an oxide? (5.3) K/U
   (a) $NaOH(s)$ (c) $CH_3OH(l)$
   (b) $MgO(s)$ (d) $H_2SO_4(l)$
5. A baker added baking soda to a cookie recipe. What would he also need to add for the baking soda to act as a leavening agent? (5.4) K/U
   (a) water
   (b) carbon dioxide
   (c) a slightly acidic substance
   (d) a slightly basic substance
6. When choosing a neutralizing agent, what factors must be considered? (5.4) K/U
   (a) safety, cost, chemical properties
   (b) chemical properties, physical properties
   (c) physical properties, safety
   (d) safety, state of matter, cost
7. What is the name applied only to a highly desired metal? (5.5) K/U
   (a) ore (c) mineral
   (b) base metal (d) precious metal
8. What process makes contaminated land safe to use again? (5.6) K/U
   (a) neutralization (c) metallurgy
   (b) remediation (d) cleanup
9. Nylon is a type of
   (a) biocatalyst
   (b) plastic foam
   (c) synthetic compound
   (d) pulping agent (5.7) K/U

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

10. The chemical formula for the hydrocarbon in the chemical equation below is C_6H_{12}. (5.1) K/U
    $2\ C_xH_y + 13\ O_2(g) \rightarrow 8\ CO_2(g) + 4\ CO(g) + 6\ H_2O(g)$
11. When incomplete combustion occurs, more than one chemical reaction may take place. (5.1) K/U
12. A solution with a pH of 10 is acidic. (5.3) K/U
13. Nitrogen dioxide, $NO_2(g)$, is a basic oxide. (5.3) K/U
14. Sodium hydroxide is used to neutralize the acid in a lake. (5.4) K/U
15. Surface mining is used to extract ores found deep within Earth's crust. (5.5) K/U
16. Before chemically removing a metal from its ore, the metal-containing ore must first be concentrated. (5.5) K/U
17. Bioremediation usually is quicker than other methods of remediation. (5.6) K/U
18. Soil flushing is a type of chemical remediation. (5.6) K/U
19. Nitric acid is produced industrially using the contact process. (5.7) K/U

Match each term on the left with the most appropriate description on the right.

20. (a) Super Glue (i) vapours from ashes of burned seaweed
    (b) iodine (ii) synthetic latex spilled on shoes
    (c) Teflon (iii) a plastic that did not turn out as planned
    (d) Scotchgard (iv) a gas cylinder that seemed empty, but was not (5.2) K/U
21. (a) battery acid (i) pH = 0
    (b) aqueous ammonia (ii) pH = 6
    (c) tomatoes (iii) pH = 14
    (d) milk (iv) pH = 2
    (e) concentrated NaOH(aq) (v) pH = 4
    (f) lemon juice (vi) pH = 11 (5.3) K/U

22.

(a) crushing	(i) Ore and air are heated to high temperatures.
(b) flash smelting	(ii) Metal-rich pieces of ore are pulled to the top by bubbles.
(c) flotation	(iii) Grinding machines break ore into small pieces.
(d) electrolysis	(iv) Ore is retrieved from beneath Earth's surface.
(e) mining	(v) Ore and oxygen are heated to high temperatures.
(f) smelting	(vi) An electric current is used to increase the purity of the metal. (5.5) K/U

Write a short answer to each question.

23. What gas is often referred to as the "silent killer"? (5.1) K/U
24. How could a traffic jam "chemically" give you a headache? (5.1) K/U
25. The roof of the BC Place Stadium in Vancouver and the coating in a non-stick cooking pan have a chemical substance in common. What could it be? (5.2) K/U
26. What common product comes from a family of chemicals called cyanoacrylates? (5.2) K/U
27. What general type of element conducts thermal energy and electricity well, is shiny when polished, and can be bent and pounded into new forms without breaking? (5.3) K/U
28. When magnesium burns in air, it leaves behind a white ash. Give the chemical name and formula for this product. (5.3) K/U
29. (a) What is a mineral?
    (b) What three types of materials can make up a mineral? (5.5) K/U
30. Why might the refining of gold be especially hazardous to the environment? (5.5) K/U
31. What is acid mine drainage? (5.5)
32. What green chemistry principle is illustrated by the use of vanadium pentoxide in the production of sulfuric acid? (5.7) K/U
33. How did chemists at the Dow Chemical Company use carbon dioxide to help protect the ozone layer in Earth's atmosphere? (5.7) K/U
34. (a) Why has Canada's pulp and paper industry greatly reduced the use of elemental chlorine as a bleach in recent years?
    (b) What chlorine-containing substance has come into use in Canada's pulp and paper industry as a substitute for elemental chlorine? Has this substitute completely solved the problem posed by chlorine? Why or why not? (5.7) K/U

Understanding

35. Write the balanced chemical equation for the complete combustion of pentane, $C_5H_{12}(l)$. (5.1) K/U C
36. Describe the function of a choke on a lawn mower with a gasoline engine. (5.1) K/U
37. How can keeping a home's natural gas furnace well maintained save money? (5.1) K/U
38. Under what conditions does a car's engine burn gasoline most completely? Why? (5.1) K/U
39. Copy and complete the following pairs of chemical equations in your notebook. For each pair, the product of the first reaction is a reactant in the second reaction. (5.3) K/U T/I C
    (a) $2\ Sr(s) + O_2(g) \rightarrow$ ________
    ________ $+ H_2O(l) \rightarrow$ ________
    (b) $P_4(s) + 3\ O_2(g) \rightarrow$ ________
    ________ $+ 6\ H_2O(l) \rightarrow$ ________
40. Which of these chemical formulas represents hydrochloric acid: $HCl(g)$ or $HCl(aq)$? Explain your answer. (5.3) T/I
41. A neutralization reaction occurs when nitric acid, $HNO_3(aq)$, and aluminum hydroxide, $Al(OH)_3(aq)$, are mixed. (5.4) K/U T/I C
    (a) Write the balanced chemical equation for the neutralization reaction.
    (b) Give the name of the ionic compound produced by the reaction.
42. A friend wants to bake a cake. (5.4) K/U A
    (a) Which should your friend use as a leavening agent: baking soda or baking powder? Why?
    (b) How will the leavening agent that you recommend cause the dough to rise?
43. Shells of organisms are mostly calcium carbonate. What might happen to shells if the water they are in becomes acidic? (5.4) K/U
44. Ore must be processed to obtain the valuable metals it contains. (5.5) K/U T/I A
    (a) Why must the ore undergo physical processes in its refining?
    (b) Why must it undergo chemical processes?
45. What is the difference between phytoremediation and bioremediation? (5.6) K/U T/I
46. Stabilization and solidification are not the same thing. Explain why these two processes are classified together as one means of remediation. (5.6) K/U T/I
47. Suppose the land beside a road has become so contaminated with road salt that plants will not grow there. Then spring rains come, and plants start to grow again. What type of remediation occurred? (5.6) K/U

48. Compounds extracted from corn are sometimes used to make adipic acid. Outline two concerns with this process that highlight why "going green" is not always simple. (5.7) K/U
49. Flash smelting creates even more of the pollutant sulfur dioxide, $SO_2(g)$, than traditional smelting. Yet flash smelting is a greener process than traditional smelting. How is this possible? (5.7) K/U
50. How does a catalyst make an industrial chemical process more sustainable? (5.7) K/U

Analysis and Application

51. Turbochargers are air compressors mounted on internal combustion engines such as those found in race cars. Turbochargers feed compressed (high-pressure) air to the engine's combustion cylinders. How could this make cars go faster? (5.1) T/I A
52. Ontario gas stations sell gasoline containing about 5 % ethanol, $C_2H_5OH(l)$. How might the presence of ethanol in a gasoline blend reduce air pollution from cars? (5.1) T/I A
53. Kerosene is a liquid mixture of hydrocarbons with about 10 to 16 carbon atoms per molecule. Hexane is a liquid hydrocarbon with 6 carbon atoms per molecule. Which do you believe would produce a cleaner-burning flame in a lamp? Explain your thinking. (5.1) T/I A
54. A motorcyclist has travelled from a low elevation to a high mountain road and her motorcycle has stalled. When she takes out a spark plug from one of the engine's combustion cylinders, she sees that it is fouled with black material. Explain what has likely happened. (5.1) K/U T/I A
55. Forest managers sometimes conduct controlled burns of forested areas. What reasons would there be for requiring the forest managers to notify local citizens of the date and time of a planned burn? (5.1) T/I A
56. Scientists are trained never to erase or delete experimental observations from their records even when observations appear to be in error and, therefore, of no value. Why is it important to keep a record of all observations regardless of whether or not they seem to be meaningful? (5.2) T/I
57. Freshly distilled (purified) water has a pH of 7. Distilled water that has been left in an open container for several days has a pH that is slightly below 7. How do you account for the difference? Include two chemical equations in your answer. (5.3) K/U C
58. Many substances are added to gasoline for various purposes. However, some additives contain metals that can coat the active surfaces of a car's catalytic converter and prevent the catalyst from coming into contact with the exhaust gases, thus "poisoning" the catalytic converter. (5.1, 5.3) T/I C A
    (a) Name a chemical substance that could be detected in a car's exhaust if its catalytic converter becomes "poisoned." Explain, using a chemical equation for a synthesis reaction, how this substance is introduced into exhaust.
    (b) Newer-model cars usually have a message visibly displayed next to the gas cap that says something like "Use unleaded gasoline only." Why would such a message be displayed?
    (c) In a car whose engine has an air/fuel mixture that is much too rich, large quantities of unburned hydrocarbons can reach the catalytic converter. In the catalytic converter, the hydrocarbons are converted to carbon dioxide and water. How might this reaction damage the catalytic converter?
59. A sample of lithium oxide is placed into water and allowed to dissolve. (5.3) K/U T/I C
    (a) How will the pH of the resulting solution compare to the pH of water? Why?
    (b) Write a balanced chemical equation to show the reaction between lithium oxide and water.
    (c) How will the temperature of the water be affected by the dissolving of the lithium oxide? Why?
60. A sample of burning sulfur, $S_8(s)$, is held in a glass bottle using a special spoon as shown in **Figure 1**. When the sulfur has burned for a minute or so, it is removed and distilled water is poured into the bottle and immediately swirled around. (5.3) K/U T/I C

Figure 1

    (a) Write a balanced chemical equation for the combustion of sulfur.
    (b) When the water in the bottle is tested with litmus paper, the paper turns red. Explain this observation and write a balanced chemical equation for the reaction.

61. Hydrogen carbonate compounds react similarly to the way carbonate compounds react. An example of a hydrogen carbonate compound is baking soda, sodium hydrogen carbonate, $NaHCO_3(s)$. (5.4) K/U T/I A
   (a) Complete the following general chemical equation for the reaction between an acid and a hydrogen carbonate compound:
   acid + hydrogen carbonate →
   (b) What happens when baking soda is used to neutralize a spill of vinegar in a kitchen? Vinegar is dilute ethanoic acid, $HC_2H_3O_2(aq)$.
62. Consider a neutralization reaction involving an antacid and stomach acid. (5.4) K/U T/I A
   (a) Describe one problem that might result from using a carbonate as an antacid.
   (b) Stomach acid is very acidic. Why should you avoid using a very basic substance to neutralize it?
63. The three requirements for a fire are fuel, oxygen, and a high temperature. Some fire extinguishers contain an acid and a carbonate compound. When the extinguisher is used, the acid and the carbonate mix and react. The products of this reaction help extinguish a fire without harming the environment. (5.4) K/U T/I A
   (a) Why might the liquid product of the reaction help put out a fire?
   (b) How is the gas produced by the reaction helpful in extinguishing a fire?
64. Look at where sodium, calcium, and magnesium are on the periodic table. (5.4) K/U T/I A
   (a) Use what you know about how basic sodium hydroxide, calcium hydroxide, and magnesium hydroxide are to predict how basic the compound potassium hydroxide is.
   (b) Would you likely use potassium hydroxide to neutralize an acidic lake? Explain.
65. The Diavik diamond mine is in the Northwest Territories. Miners there extract diamonds from near the surface and also from deep underground. Name and describe the types of mining used in the Diavik mine. (5.5) K/U A
66. Explain why, in most metallurgical processes, physical processing takes place before chemical processing occurs. (5.5) K/U T/I
67. Why must plans to close a mine be in place before the mine is opened? In your answer, suggest potential problems that such a policy might prevent. (5.5) K/U T/I
68. An environmental technician found that a well located near a newspaper print shop contained contaminants from the printing business. The well was treated with hydrogen peroxide, $H_2O_2(aq)$, and almost all of the contaminants were changed to less harmful substances. (5.6) K/U T/I A
   (a) Was this remediation physical or chemical?
   (b) What type of remediation was used at this site?
69. Sometimes more than one method of remediation is used at the same time. In the 1990s, researchers found contaminated soil at Canadian Forces Base Borden in Ontario. A solution of a powerful oxidizing agent, potassium permanganate, $KMnO_4(aq)$, was added to the soil. Contaminants in the soil were changed to less harmful substances, which were then washed out of the soil. (5.6) K/U T/I A
   (a) What two means of remediation were used?
   (b) Classify each of these means of remediation as either physical or chemical.
70. The land around a manufacturing plant was contaminated with compounds that contained metal ions. The company decided to reclaim the metal from the compounds by using electrolysis. They put electrodes into the soil and ran a current from one electrode to the other through a wire above ground. No current flowed through the ground. What change needs to be made so that electrolysis can be used? (5.6) K/U T/I A
71. Some of the methods used to detoxify land can also improve the quality of water and air. For example, the disappearance of rainforests worldwide has a negative effect on the atmosphere. Trees naturally help remove one particular substance from the air. Explain how we could use trees as a means of phytoremediation for the air. (5.6) K/U T/I A
72. An entrepreneur plans to build a factory in which precious metals will be recovered from scrap electronic equipment. The first step is to be grinding. The grindings, along with scrap parts that are too difficult to grind, will then be placed into vats of hot acid to dissolve the metals so they can later be recovered as precipitates. (5.7) K/U T/I A
   (a) How should the entrepreneur construct the machinery in the planned factory to make the processing “closed loop?”
   (b) What are three advantages that closed-loop processing would offer?
73. An effective catalyst can make an industrial chemical process greener. How might a catalyst also make an industrial chemical process more profitable? (5.7) K/U T/I

74. Allyl alcohol, $C_3H_5OH(l)$, is used to make glycerol, an important chemical in the food industry. Whereas allyl alcohol is toxic, glycerol can be used as a sugar substitute. The chemical equations below show two possible ways to synthesize allyl alcohol. Examine the reactions and then, using only the provided information, state which you believe to be the greener synthesis. Explain your reasoning. (5.7) K/U T/I A

Synthesis 1:

$$\underset{\text{propene}}{2\,C_3H_6(g)} + \underset{\text{ethanoic acid}}{2\,HC_2H_3O_2(aq)} + O_2(g) \rightarrow \underset{\text{allyl acetate}}{2\,C_3H_5OCOCH_3(l)} + 2\,H_2O(l)$$

$$\underset{\text{allyl acetate}}{C_3H_5OCOCH_3(l)} + H_2O(l) \rightarrow \underset{\text{allyl alcohol}}{C_3H_5OH(l)} + \underset{\text{ethanoic acid}}{CH_3COOH(l)}$$

Synthesis 2:

$$\underset{\text{allyl chloride}}{C_3H_5Cl(l)} + H_2O(l) \rightarrow \underset{\text{allyl alcohol}}{C_3H_5OH(l)} + HCl(aq)$$

75. Recycled paper must be pulped, de-inked, and bleached before it can be used to make new paper. An inventor proposes a process that uses soybean peroxidase, a biological catalyst, to bleach the de-inked paper fibres. Give at least two reasons why this could be a green strategy. (5.7) K/U T/I A

Evaluation

76. In renovating his basement, a homeowner needs to replace the duct pipes that bring combustion air from the outdoors to the furnace. The homeowner wants to buy cheaper replacement pipes, but that means using pipes with a significantly narrower diameter than the existing pipes. Is using the narrower pipes an acceptable approach? Explain your answer. (5.1) T/I A

77. A friend is considering modifying his car to burn pure hydrogen instead of gasoline. He believes that a hydrogen-powered car will be less polluting and help conserve fossil fuels. (5.1, 5.3) T/I A
    (a) Why might a hydrogen-powered car be less polluting?
    (b) What information would your friend have to find, through research, to determine if hydrogen-powered cars can help to preserve fossil fuels?
    (c) Cite and explain one drawback or problem that you can foresee for hydrogen-powered cars.

78. A number of scientists and engineers believe that dumping great quantities of calcium oxide (lime) into the oceans could help seawater absorb more atmospheric carbon dioxide and thus limit climate change. (5.3, 5.4) K/U C A
    (a) Explain how calcium oxide could help seawater absorb more carbon dioxide out of the air. Include balanced chemical equations as a part of your answer.
    (b) Thermal energy from fossil fuels is used to produce calcium oxide from limestone. Knowing this, what would you have to find out, through research, to decide whether the proposed plan could actually reduce the concentration of carbon dioxide in the atmosphere?

79. Nelson Lake is located just a few kilometres from smelters that are near Sudbury, Ontario. Smelters release pollutants that can make water more acidic. Over the years, the smelters have been modified to reduce the quantity of pollutants that they release into the atmosphere. In the early 1970s, the low pH of the water in Nelson Lake limited the types of fish living there. In the mid-1970s, the lake was limed. Several years later, more types of fish were identified in the lake, indicating that the lake was less polluted. Describe two possible causes of the improved health of the lake. (5.4) K/U T/I A

80. Working individually, compare and contrast the following processes. Then discuss your answers with a partner. (5.5) K/U T/I C
    (a) flotation and electrolysis
    (b) smelting and flash smelting

81. Phenol is a common waste product found in the wastewater from steelmaking. It is toxic and very harmful to the environment. Two proposed processes for remediating the wastewater are given below. From a green chemistry perspective, compare the potential environmental friendliness and usefulness of these processes.

    Process 1: The wastewater containing phenol is passed over activated charcoal. The phenol collects on the surface of the charcoal. The charcoal is then burned in air to convert the phenol to carbon dioxide and water.

    Process 2: The wastewater containing phenol is trickled through reactors packed with ceramic beads. The beads are coated with bacteria obtained from cow dung. The bacteria consume the phenol and convert it to carbon dioxide and water. Periodically, some of the bacterial mass must be removed and discarded to prevent the reactor from clogging. (5.6, 5.7) K/U T/I A

82. In sandblasting, compressed air blows grains of sand at high speed against a surface. The sand can strip off rust, grime, and paint. However, dust from the sand is an inhalation hazard. Could recycled glass crushed into tiny bits serve as a greener alternative to sand? Describe what you would have to learn, through research, to answer this question. (5.7) T/I A

83. Many politicians and environmentalists recommend laws that would require manufacturers to upgrade their processes to be greener. One of your relatives is very opposed to these laws. He points out that upgrading manufacturing processes is very expensive and, if required by law, could cause companies to go out of business. This would hurt the economy. What argument could you offer to change your relative's opinion? (5.7) T/I A

Reflect on Your Learning

84. You have learned about flames and combustion in this chapter. What one thing, related to combustion, do you find most relevant to your life? Explain its relevance. K/U A
85. Aluminum is a reactive metal that would rapidly undergo chemical changes to form compounds if it did not form tightly adhering, protective coatings of oxide. How would your life be affected if aluminum did not form a self-protecting oxide coating? K/U A
86. In July 2010, a hazardous material alert was issued as the result of a spill of sulfuric acid in Sudbury, Ontario. Fortunately, no damage resulted from the spill. If there had been need to clean up the spill, what two types of materials could have been used to neutralize the acid and reduce the effects of the spill? Give an example of each of these types of materials. T/I A
87. Suppose that your province is considering mining a recently discovered valuable ore deposit. T/I A
    (a) What advantages do you think there will be for your community if the mine becomes operational?
    (b) What cautions and preventive measures must the operators of the mine establish to prevent damage to the environment?
88. Write a paragraph that explains how the following terms relate to chemistry, the environment, and each other: neutralization; detoxification; mining; reclamation; contamination. K/U T/I C
89. Consider your own community. K/U T/I A
    (a) What are possible sources of chemical contamination in your community?
    (b) For each of these sources of contamination, describe what methods of remediation might be used.
90. Considering what you have learned in this chapter, would you encourage your family to buy products whose manufacturing and uses are green? Why or why not? T/I A
91. To avoid converting to greener processes and to access cheaper labour, some manufacturers have moved their factories and the resulting pollution to countries that have lax environmental regulation. Suppose that a Canadian government official proposed a law to put a large import tax, as a penalty, on products from companies that conduct highly polluting operations in foreign countries. What are your thoughts about such a law? T/I A

Research

GO TO NELSON SCIENCE

92. Would burning garbage to generate electricity produce combustion gases full of harmful pollutants? Research what is happening in the field of obtaining clean energy from garbage. Based on your findings, write a letter that could be sent to a government official in which you argue whether or not plants that burn garbage should be supported by the government. Support your argument with specific information. T/I C A
93. Research the history of the use of Super Glue in medical applications. Also find out what kinds of related adhesives are employed in medicine currently, including examples of how they are used. Give an oral presentation of your findings to your class that includes information about whether or not it is wise to use Super Glue in the home to treat cuts. K/U C A
94. Research the sequence of chemical reactions responsible for the formation of the acids found in acid rain. Also find out how acid rain harms plants and aquatic animals. Write a report that includes balanced chemical equations and diagrams. K/U C A
95. In Canada, recycling centres and many local car service centres collect car batteries for recycling. T/I A
    (a) Use the Internet to find the nearest location where you can hand in old car batteries for recycling.
    (b) What does the Canadian Centre for Occupational Health and Safety list as some health hazards associated with sulfuric acid?
96. The Kemess South mine in British Columbia is a good example of a mine that strives to prevent harm to the environment. K/U T/I
    (a) Does the Kemess South mine use surface methods or underground methods to mine the ore?
    (b) What two metals are mined at Kemess South?
    (c) The mine can produce up to 56 000 t of ore a day. This ore is then reduced to approximately 800 t of concentrate by flotation. Find out how the water used in this process is then purified.
    (d) What is done at this mine to minimize possible damage to the environment?
97. The TAML activator, a type of catalyst, has helped to make Canada's pulp and paper industry greener. Research the TAML activator in detail. Find out its chemical structure and how it works. Also, research other applications of TAML catalysts in remediating pollutants. Report your findings, including a diagram of the structure of a TAML molecule. T/I C A

Detoxing Contaminated Soil

Canada has its share of sites contaminated with waste metals. Environmental chemists have to plan aggressive rehabilitation measures to make a contaminated site suitable for new home construction (**Figure 1**). There are several options for cleaning up a site. For example, contaminated soil can be stripped from the site, treated to remove the contaminants, and then returned. One method of treating soil contaminated with metals or metal compounds is called acid leaching. This process involves passing an acid such as sulfuric acid through the excavated soil. As the acid percolates through the soil it reacts with the metal contaminants, forming dissolved metal ions. These ions drain out of the soil along with the acid. The metal ions can then be precipitated from the solution or converted to solid elements using a single displacement reaction.

Figure 1 A soil researcher collects a sample of groundwater for testing.

In this Unit Task, you will take the role of an environmental chemist trying to determine the best way to recover copper metal from soil contaminated with copper(II) oxide. Your task is divided into four parts.

Part A provides a detailed procedure for using sulfuric acid to leach copper(II) ions from a soil sample contaminated with copper(II) oxide.

In Parts B and C, you will design two different procedures to recover copper metal, Cu(s). At least one of the procedures should involve a single displacement reaction.

In Part D, you will suggest an appropriate way to neutralize the remaining solution so that it can be safely disposed of.

Purpose

To evaluate two methods of recovering copper metal from a soil sample contaminated with copper(II) oxide.

Equipment and Materials

Part A: Collecting $Cu^{2+}(aq)$

- chemical safety goggles
- lab apron
- balance
- 50 mL graduated cylinder
- two 125 mL Erlenmeyer flasks
- hot plate with wire gauze
- thermometer
- retort stand with two clamps
- scoopula
- flask tongs
- funnel
- glass rod
- filter paper
- mixture of sand and copper(II) oxide, CuO(s)
- massing paper
- dilute sulfuric acid, $H_2SO_4(aq)$ (0.1 mol/L)
- dropper bottle of distilled water

Sulfuric acid is corrosive. Avoid skin and eye contact. Wash any spills on the skin, in the eyes, or on clothing immediately with cold water. Report any spills to your teacher.

Copper(II) oxide and the solutions it forms are toxic and irritating. Avoid skin and eye contact. Wash any spills on the skin immediately with cold water.

Procedure

Part A: Collecting $Cu^{2+}(aq)$

1. Put on your chemical safety goggles and lab apron.
2. Measure a 2.0 g sample of contaminated soil (sand and copper(II) oxide) into a piece of paper.
3. Measure 40 mL of sulfuric acid and pour it into an Erlenmeyer flask.
4. Place the flask on the gauze on the hot plate. Clamp the flask to the retort stand.
5. Warm the acid on the hot plate to about 50 °C. Check the temperature periodically with the thermometer. Use the thermometer to very gently stir the acid. When the acid reaches 50 °C, turn down the hot plate.

Do not allow the sulfuric acid to boil. Do not allow the thermometer to touch the bottom of the flask. Remove the thermometer when you are not taking a reading.

6. Use a scoopula to add small portions of the soil sample to the acid in the flask. Stir the solution with a glass rod after each addition. Record your observations.
7. If you do not observe the blue colour of $Cu^{2+}(aq)$ ions, the copper(II) oxide in the mixture did not completely react. Continue to warm the flask and its contents on the hot plate, but do not let the temperature exceed 50 °C.
8. Continue adding the soil to the flask until the entire sample has been added, mixed, and reacted.
9. Turn off the hot plate and allow the mixture in the flask to cool to room temperature.

To unplug the hot plate, pull on the plug not the cord.

10. Filter the $Cu^{2+}(aq)$ solution into the other Erlenmeyer flask. Use the glass rod to guide the solution onto the filter paper. Record your observations.
11. Wash the sand residue with small volumes of water to ensure that all the $Cu^{2+}(aq)$ ions have been transferred to the flask. Rinse out the flask.
12. Divide the solution into two equal parts in the two flasks.

Part B: Recovering Cu(s) (Method 1)

13. Design a procedure to convert $Cu^{2+}(aq)$ ions into copper metal, Cu(s), using a single displacement reaction. Your procedure should include collecting the solid. Be sure to list safety precautions and equipment.
14. Proceed once your plan has been approved by your teacher. Record your observations.

Part C: Recovering Cu(s) (Method 2)

15. Plan another procedure to convert $Cu^{2+}(aq)$ ions into copper metal, Cu(s). Use a different reaction or process than that used in Part B. Again, include safety precautions and equipment.
16. Proceed with your teacher's approval.

Part D: Neutralizing Lab Waste

17. Write a procedure to safely neutralize any remaining sulfuric acid in the flask and to dispose of the solution.
18. Proceed with your teacher's approval.

Analyze and Evaluate

(a) Write a chemical equation for the reaction that occurred in Part A. Classify this reaction as synthesis, decomposition, single displacement, or double displacement. Justify your choice. K/U T/I C

(b) Why was it necessary to add the soil sample in small portions? T/I

(c) Justify your choice of the solid reactant in Part B. T/I

(d) Write the chemical equation for the reaction that occurred in Part B. K/U C

(e) Explain why you chose the reaction in Part C. T/I

(f) Justify your choice of the chemical used in Part D to neutralize remaining sulfuric acid. T/I

(g) Write a chemical equation for the neutralization reaction in Part D. K/U C

(h) Evaluate your two procedures to recover copper metal from the soil sample, using your own evaluation criteria. Which worked better? Why? What improvements would you recommend? T/I

Apply and Extend

(i) Excavating soil and then treating it off-site is very costly and disruptive. Acid leaching of soil can be done without excavation. Research how this form of remediation is done on-site. Based on your research, which of the types of acid leaching would you recommend to a town wishing to remediate a site for use as a sports field? T/I A

ASSESSMENT **CHECKLIST**

Your completed Unit Task will be assessed according to the following criteria:

Knowledge/Understanding

- ✓ Write balanced chemical equations to represent the reactions observed.
- ✓ Demonstrate an understanding of activity series in selecting a metal to displace copper from the solution.
- ✓ Demonstrate an understanding of neutralization reactions when selecting the most appropriate way to neutralize the waste solution.

Thinking/Investigation

- ✓ Plan safe and effective procedures for extracting the metal.
- ✓ Conduct the procedures safely and effectively.
- ✓ Record all observations carefully and in an organized manner.
- ✓ Analyze the results.
- ✓ Evaluate the two procedures.

Communication

- ✓ Prepare a suitable lab report that includes the complete procedures, summary of the observations, any necessary analysis, and evaluation.

Application

- ✓ Evaluate two different chemical reactions to see which is best for a given application, using your own criteria.

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

For each question, select the best answer from the four alternatives.

1. The law of conservation of mass describes which of the following statements? (4.1) K/U
   (a) The atoms of the reactants are the same as the atoms of the products.
   (b) A catalyst exists before and after a reaction without being used up.
   (c) Natural resources used in laboratories must be conserved.
   (d) A precipitate is formed in all changes that are classified as chemical reactions.
2. Which statement about the reaction described by this chemical equation is true? (4.1) K/U
   $Na_2O(s) + H_2O(l) \rightarrow 2\ NaOH(aq)$
   (a) The reaction yields solid elemental sodium.
   (b) Sodium oxide is produced during the reaction.
   (c) Thermal energy is required for the reaction to proceed.
   (d) Liquid water is one of the reactants.
3. Which of the following chemical equations is correctly balanced? (4.1) K/U
   (a) $2\ Al(s) + 2\ O_2(g) \rightarrow 2\ Al_2O_3(s)$
   (b) $C_6H_6(l) + 13\ O_2(g) \rightarrow 6\ CO_2(g) + H_2O(g)$
   (c) $FeCl_3(s) + 3\ NaOH(aq) \rightarrow Fe(OH)_3(s) + 3\ NaCl(aq)$
   (d) $(NH_4)_2CO_3(s) \rightarrow 2\ NH_3(g) + 2\ H_2O(l) + 2\ CO_2(g)$
4. Beryllium and oxygen combine in a synthesis reaction. Which of the following chemical formulas represents the product? (4.2) K/U
   (a) Be_2O_2
   (b) Be_2O
   (c) BeO_2
   (d) BeO
5. Which of the following industrial chemicals might be produced from components in syngas? (4.3) K/U
   (a) ethanol, C_2H_5OH
   (b) ammonia, NH_3
   (c) sodium chloride, $NaCl$
   (d) silver nitrate, $AgNO_3$
6. Which of the following metals may displace sodium in a displacement reaction? (4.4) K/U
   (a) magnesium
   (b) hydrogen
   (c) lithium
   (d) silver
7. What is the largest source of atmospheric mercury? (4.5) K/U
   (a) drilling for oil
   (b) combustion of coal
   (c) storage of nuclear waste
   (d) acidification of rain
8. Which type of reaction occurs when sulfuric acid is heated to produce water, oxygen, and sulfur dioxide? (4.2, 4.4, 4.6) K/U
   (a) synthesis
   (b) decomposition
   (c) single displacement
   (d) double displacement
9. Which combination of reactants will lead to the formation of a gas? (4.6) K/U
   (a) $Na_2SO_3(aq) + 2\ HCl(aq)$
   (b) $Mg(OH)_2(s) + 2\ HCl$
   (c) $KCl(aq) + AgNO_3(aq)$
   (d) $CuSO_4(aq) + 2\ NaOH(aq)$
10. What is usually necessary for combustion to occur? (5.1) K/U
    (a) a hydrocarbon
    (b) pure oxygen
    (c) a fuel and oxygen
    (d) a fuel and carbon dioxide
11. Which of the following best applies to an accidental scientific discovery? (5.2) K/U
    (a) Its use might not be determined until years after the material is discovered.
    (b) It is the desired end product of scientific research.
    (c) It is a waste product with no uses.
    (d) Its uses are all for scientific purposes; there are no practical uses.
12. What does it usually mean when unexpected observations are made in an investigation? (5.2) K/U
    (a) Useful products are produced accidentally.
    (b) The products of the investigation are not useful.
    (c) Something went wrong with the investigation.
    (d) Observations were made incorrectly.
13. Which of these compounds forms an acidic solution when it is dissolved in water? (5.3) K/U
    (a) sodium oxide
    (b) nitric oxide
    (c) calcium oxide
    (d) potassium oxide

14. Which of the following pH values indicates a basic solution? (5.3) K/U
    (a) 3
    (b) 5
    (c) 7
    (d) 9
15. Which statement describes what happens when an acid is neutralized by a hydroxide compound? (5.4) K/U
    (a) An acid and a base react to form water and an ionic compound.
    (b) An acid and water react to form a base and an ionic compound.
    (c) An acid and a base react to form water, carbon dioxide, and an ionic compound.
    (d) Water and an ionic compound react to form an acid and a base.
16. During the neutralization of hydrochloric acid, HCl(aq), by sodium carbonate, Na_2CO_3(s), what is the source of the carbon dioxide produced? (5.4) K/U
    (a) the decomposition of hydrochloric acid
    (b) the decomposition of carbonic acid
    (c) the decomposition of sodium carbonate
    (d) the formation of water as a product
17. Some rocks contain enough of a valuable mineral that it is profitable to mine the rock. What is the name given to such a rock? (5.5) K/U
    (a) ore
    (b) mineral
    (c) metal
    (d) tailings
18. Which of the following is an example of physical remediation? (5.6) K/U
    (a) electrolysis
    (b) bioremediation
    (c) soil flushing
    (d) oxidation
19. Developing a furnace that uses less fuel to heat a building is an example of what principle of green chemistry? (5.7) K/U
    (a) Waste not.
    (b) Use less energy.
    (c) Use chemicals that have less environmental impact.
    (d) Use renewable resources.

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

20. The symbol (l) describes a substance dissolved in water. (4.1) K/U
21. Electricity causes molten potassium chloride to break up into its elements during a synthesis reaction. (4.2) K/U
22. The process of gasification is used to convert toxic waste into safer substances. (4.3) K/U
23. When solid zinc is combined with aqueous sulfuric acid, zinc displaces sulfur. (4.4) K/U
24. Photochemical reactions in the Arctic ecosystem might cause mercury ions to lose two electrons. (4.5) K/U
25. The quantity of solute that will dissolve in a given volume of solvent is the solute's solubility. (4.6) K/U
26. Carbon monoxide is a product of complete combustion. (5.1) K/U
27. More energy is released during complete combustion than is released during incomplete combustion for the same quantity of fuel. (5.1) K/U
28. Not all important discoveries are the result of carefully designed experiments. (5.2) K/U
29. Sodium oxide is an example of an acidic oxide. (5.3) K/U
30. Antacids might contain either a base or a carbonate. (5.4) K/U
31. Metallurgy involves removing metals from ores and making them into useful products. (5.5) K/U
32. The toxic metals removed from soil by electrolysis often can be recycled into useful products. (5.6) K/U
33. Phytoremediation is the use of micro-organisms to remove contaminants from soil and water. (5.6) K/U
34. Scrubbers remove emissions that could be harmful to the environment. (5.7) K/U
35. Catalysts are materials that decrease the efficiency of a reaction. (5.7) K/U

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

Knowledge

For each question, select the best answer from the four alternatives.

1. Which of the following equations describes a decomposition reaction? (4.2) K/U
   (a) $2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$
   (b) $H_2CO_3 \rightarrow H_2O + CO_2$
   (c) $BaO + H_2O \rightarrow Ba(OH)_2$
   (d) $2\ KI + Br_2 \rightarrow 2\ KBr + I_2$
2. What is one potential use of plasma gasification in Canada? (4.3) K/U
   (a) to convert garbage in landfills into usable substances
   (b) to remove greenhouse gases produced during the combustion of coal
   (c) to make natural gas burn more cleanly and efficiently
   (d) to clean contaminants from underground water resources
3. Which of the following will not result in a chemical reaction? (4.4) K/U
   (a) Aluminum is dipped into a zinc nitrate solution.
   (b) Zinc is added to a solution of copper(II) sulfate.
   (c) Lead is placed into an iron(III) solution.
   (d) Sodium is placed into cold water.
4. Which metal is least likely to displace hydrogen from acids? (4.4) K/U
   (a) potassium (c) gold
   (b) magnesium (d) aluminum
5. Which of the following equations represents a single displacement reaction? (4.4) K/U
   (a) $HCl(aq) + NaOH(aq) \rightarrow HOH(l) + NaCl(aq)$
   (b) $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$
   (c) $Ca(ClO_3)_2(s) \rightarrow CaCl_2(s) + 3\ O_2(g)$
   (d) $Cl_2(g) + 2\ NaBr(aq) \rightarrow 2\ NaCl(aq) + Br_2(l)$
6. What property of mercury makes its entry into Arctic food chains a cause for concern? (4.5) K/U
   (a) It is a liquid at room temperature.
   (b) It is a potent neurotoxin.
   (c) It is the main component in air.
   (d) It destroys ozone near Earth's surface.
7. Which of the following equations shows the general pattern for a double displacement reaction? (4.6) K/U
   (a) $AB + CD \rightarrow AD + CB$
   (b) $AB \rightarrow A + B$
   (c) $AB + C \rightarrow A + CB$
   (d) $A + B \rightarrow AB$
8. When ethane, C_2H_6, burns, it combines with oxygen to form carbon dioxide and water. What coefficient is used with O_2 in the balanced chemical equation for the reaction? (4.1, 5.1) K/U
   (a) 1
   (b) 2
   (c) 4
   (d) 7
9. What are the products of the complete combustion of a hydrocarbon? (5.1) K/U
   (a) water and carbon monoxide
   (b) water and carbon dioxide
   (c) carbon monoxide and carbon dioxide
   (d) other hydrocarbons
10. Which of these compounds is an example of a basic oxide? (5.3) K/U
    (a) sulfur trioxide
    (b) carbon dioxide
    (c) dinitrogen pentoxide
    (d) magnesium oxide
11. Which type of reaction involves an acid and a base producing water and an ionic compound? (5.4) K/U
    (a) combustion
    (b) acidification
    (c) neutralization
    (d) decomposition
12. In metallurgy, what is the main purpose of flotation? (5.5) K/U
    (a) to remove the metal from the rock
    (b) to remove the ore from Earth's crust
    (c) to make useful products from the metal
    (d) to concentrate the metal-containing ore
13. What is the name applied to a plant that has a high ability to take in contamination and trap it? (5.6) K/U
    (a) micro-organism
    (b) hyperaccumulator
    (c) phytoremediator
    (d) bioremediator
14. Which of the following is an example of using renewable resources as a starting material? (5.7) K/U
    (a) using plants instead of petroleum to make certain plastic products
    (b) using safe materials instead of CFCs as refrigerants
    (c) recycling steel cans into new steel cans instead of obtaining iron from ore
    (d) using plastic milk jugs to make plastic lumber

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

15. In the following chemical equation the 3 in front of O_2 is called a subscript:
    $2\ ZnS + 3\ O_2 \rightarrow 2\ ZnO + 2\ SO_2$ (4.1) K/U
16. Potassium chlorate is heated to produce potassium chloride and oxygen in a synthesis reaction. (4.2) K/U
17. Syngas is a mixture that can be used to generate electricity because it burns like a fossil fuel. (4.3) K/U
18. The reactivity of lead is lower than the reactivity of zinc. (4.4) K/U
19. Every winter, mercury disappears from the Arctic atmosphere. (4.5) K/U
20. When potassium chloride and silver nitrate undergo a single displacement reaction, silver chloride precipitates out of the solution. (4.6) K/U
21. A hydrocarbon is one type of organic compound. (5.1) K/U
22. Sometimes, discoveries made accidentally during scientific research have practical uses. (5.2) K/U
23. Oxides of non-metallic elements are often sources of air pollution. (5.3) K/U
24. When acting as a leavening agent, the acid and the carbonate compound in baking soda react with each other when water is added. (5.4) K/U
25. Flash smelting is a type of physical processing. (5.5) K/U
26. During stabilization and solidification, contaminants are changed to a safer form. (5.6) K/U

Match each term on the left with the most appropriate description on the right.

27. (a) subscript
    (b) decomposition
    (c) coefficient
    (d) single displacement

    (i) a number that appears as a factor in front of a formula in a chemical equation
    (ii) type of reaction useful for recovering a dissolved metal from solution
    (iii) a small number that appears after a symbol in a chemical formula
    (iv) type of chemical reaction in which a complex compound breaks down into simpler products (4.1, 4.2, 4.4) K/U

28. (a) HCl(aq)
    (b) MgO(s)
    (c) KOH(aq)
    (d) SO_3(g)

    (i) acid
    (ii) base
    (iii) acidic oxide
    (iv) basic oxide (5.3) K/U

Write a short answer to each question.

29. How do chemical equations compare to mathematical equations? (4.1) K/U
30. What can you do to confirm that a chemical equation is balanced? (4.1) K/U
31. How do you know that melting ice is a physical change and not a chemical change? (4.1) K/U
32. How might synthesis and decomposition reactions be reversible? Give an example. (4.2) K/U
33. Why can the products of some synthesis reactions, such as between carbon and oxygen, be determined only through chemical tests? (4.2) K/U
34. What happens to atoms as a result of the intense heat produced during plasma gasification? (4.3) K/U A
35. Why is hydrogen included in the activity series of metals even though it is not a metal? (4.4) K/U
36. How can you use the activity series of metals to decide if a single displacement reaction will occur? (4.4) K/U
37. How is the single displacement pattern for halogens different from the pattern used for metals in terms of the ion being displaced? (4.4) K/U
38. Why are bromine atoms more reactive than bromide ions? (4.5) K/U
39. Describe three factors that affect the solubility of a solute. (4.6) K/U
40. Which gas is produced when sulfides, such as Na_2S, react with acids? (4.6) K/U
41. Explain how human health might be affected by the combustion of trees during a forest fire. (5.1) K/U
42. Briefly explain the relationships among the following terms: organic compound, hydrocarbon, and fuel. (5.1) K/U
43. How might the discovery of iodine have been changed if Bernard Courtois had had a large supply of firewood? (5.2) K/U T/I
44. From what you know about basic oxides, explain how a base might be formed from a sample of magnesium. (5.3) K/U
45. Briefly explain why diphosphorus pentoxide is an acidic oxide. (5.3) K/U
46. When recycling a battery, the sulfuric acid from the battery might be neutralized. (5.4) K/U T/I
    (a) If the products are water and sodium sulfate, what base was used to react with sulfuric acid?
    (b) Justify your answer to (a).
47. What are two main advantages of flash smelting over regular smelting? (5.5) K/U

48. Mining has a notoriously bad record of pollution. (5.5) K/U
   (a) What is the single most important air pollutant released by the mining industry?
   (b) What is being done to reduce the amount of this pollutant released into the air?
49. The use of micro-organisms for bioremediation is often an environmentally benign choice. However, the result may not be as "green" as we would like. What is the downside to using micro-organisms to clean up spilled hydrocarbons? (5.6) K/U
50. (a) Name one compound that is used for the chemical oxidation of contaminants.
   (b) What happens during the oxidation process? (5.6) K/U
51. How might using recycled paper instead of wood pulp change the bleaching process used while making paper? (5.7) K/U
52. List at least three ways in which the pulp and paper industry is becoming greener. (5.7) K/U

Understanding

53. Silicon reacts with carbon dioxide to form silicon carbide, SiC, and silicon dioxide, SiO_2. (4.1) K/U C
   (a) What are the reactants in this reaction?
   (b) What are the products in this reaction?
   (c) Write the balanced chemical equation describing the reaction.
54. Balance the following chemical equations: (4.1) K/U
   (a) $Al_2O_3 + C \rightarrow Al + CO_2$
   (b) $Fe(OH)_2 + H_2O_2 \rightarrow Fe(OH)_3$
   (c) $Ag_2O \rightarrow Ag + O_2$
55. Photosynthesis is an essential chemical reaction. During photosynthesis, carbon dioxide, CO_2, and water react to form glucose, $C_6H_{12}O_6$, and oxygen, O_2. This reaction takes place only in sunlight, with chlorophyll as a catalyst. (4.1) K/U T/I C
   (a) Write a word equation for the chemical reaction that occurs during photosynthesis. Be sure to indicate the roles of sunlight and chlorophyll.
   (b) Write a balanced chemical equation for the reaction.
56. Predict the product and write a balanced equation for each of the following examples: (4.1, 4.2) K/U C
   (a) $H_2O + K_2O \rightarrow$
   (b) $Ca + O_2 \rightarrow$
   (c) $CaCO_3 \rightarrow$
57. Examine the chemical equations below. Compare and contrast the types of reaction and the products formed. (4.1, 4.2) K/U

   $2\,C + O_2 \rightarrow 2\,CO$
   $CaO(s) + H_2O \rightarrow Ca(OH)_2$
58. Mercury is heated with chlorine gas to yield solid mercury(II) chloride. (4.1, 4.2) K/U C
   (a) Classify the reaction.
   (b) Write the balanced chemical equation for the reaction.
59. Silver nitrate is heated to form silver nitrite and oxygen. (4.1, 4.2) K/U C
   (a) Classify the reaction.
   (b) Write the balanced chemical equation for the reaction.
60. During plasma gasification, garbage is crushed into pellets and fed into a sealed chamber. (4.3) K/U A
   (a) What is the role of the plasma torch?
   (b) Name two products formed and how each can be used.
61. If metallic potassium is dropped into water, it reacts violently as potassium atoms replace hydrogen atoms in water molecules. So much energy is released that the resulting hydrogen gas burns. (4.1, 4.4) K/U T/I C
   (a) What type of chemical reaction occurs between potassium and water?
   (b) Write a word equation for the reaction.
   (c) Write a balanced chemical equation for the reaction.
62. A chemistry student attempts to complete the two reactions below. Predict what will be observed. Explain your predictions, including any chemical equations for the reactions that occur. (4.4) K/U C
   (a) The student places a copper wire into an aluminum nitrate solution.
   (b) The student places an aluminum wire into a copper(II) nitrate solution.
63. Small pieces of several metals are added to separate samples of water. Predict which metal will result in a chemical reaction. Write a balanced chemical equation for each reaction that occurs. (4.4) K/U C
   (a) copper
   (b) calcium
   (c) lithium
   (d) lead
64. A sample of silver nitrate is mixed with hydrochloric acid. (4.1, 4.2, 4.4, 4.6) K/U C
   (a) Which type of reaction occurs?
   (b) Predict the products of the reaction.
   (c) Write a balanced equation describing the reaction.
65. A sample of solid lead(II) nitrate is mixed with sodium carbonate solution. (4.1, 4.2, 4.4, 4.6) K/U C
   (a) Predict the precipitate formed during this reaction.
   (b) Write a balanced equation describing the reaction.
66. Explain the difference between complete combustion and incomplete combustion. (5.1) T/I

67. Compare the products of complete combustion to the products of incomplete combustion. (5.1) K/U T/I
68. Use the accidental discoveries mentioned in Section 5.2 to explain the importance of keeping accurate records during scientific experiments. (5.2) K/U T/I
69. How does the accidental discovery of useful products support the importance of communication among scientists? (5.2) T/I C
70. How can you use the periodic table to determine whether an element forms an acidic oxide or a basic oxide? (5.3) K/U T/I
71. For each of the following elements, state whether it forms an acidic oxide or a basic oxide: (5.3) K/U
    (a) sodium
    (b) sulfur
    (c) nitrogen
    (d) magnesium
    (e) carbon
    (f) aluminum
72. (a) What two requirements are necessary for an effective antacid?
    (b) Can all bases be used as antacids? Explain your answer. (5.4) K/U T/I A
73. (a) Explain how acid mine drainage can damage surrounding water supplies.
    (b) What can be done to minimize the environmental damage caused by mine drainage? (5.5) K/U T/I
74. The chemical process of flash smelting can produce a metal that is relatively pure. That relatively pure metal is either used in various products or purified further. (5.5) K/U T/I
    (a) What process is used to further purify a metal after smelting?
    (b) Explain this purifying process.
75. What determines whether land is contaminated enough to need remediation? (5.6) T/I
76. In your own words, describe four remediation strategies that should be considered when attempting to make contaminated land safe for use. (5.6) K/U
77. Canada is the world's largest producer of zinc. There is a zinc mine in nearly every province. Tailings from these zinc mines contain zinc compounds that were not removed by the initial processing. Briefly describe how electrolysis can be used to recover zinc metal from the mine tailings. (5.6) K/U A
78. Both a catalyst and a reactant are part of a chemical reaction. Contrast a catalyst and a reactant in terms of what happens to them during a chemical reaction. (5.7) K/U T/I

Analysis and Application

79. Section 4.1 lists four different types of evidence of chemical reaction. Using examples, explain why this evidence does not absolutely determine whether a change is chemical or physical. (4.1) K/U T/I C
80. A student is asked to write a balanced chemical equation to describe the reaction between hydrogen gas and oxygen gas to form water. The student first wrote the following unbalanced equation: (4.1) T/I C

    $H_2 + O_2 \rightarrow H_2O$

    Then the student changed the subscript of oxygen in the product to balance the equation as follows:

    $H_2 + O_2 \rightarrow H_2O_2$

    (a) What error did the student make in balancing the equation? Explain.
    (b) Write a correctly balanced equation for this reaction.
81. Lithium compounds are used to purify air in space vehicles. For example, lithium hydroxide, LiOH, reacts with carbon dioxide to form lithium carbonate, Li_2CO_3, and water. (4.1) K/U T/I C
    (a) Write a balanced chemical equation describing this reaction.
    (b) Explain why lithium compounds are used rather than more common elements, such as sodium.
82. Several oxides of sulfur react with water in the air to form acids. (4.1, 4.2) K/U A
    (a) Write a balanced equation to represent the reaction of sulfur trioxide with water to form sulfuric acid.
    (b) Classify the reaction that occurs when acid precipitation is formed in this way.
    (c) Describe how the formation of acid rain is related to human activities.
83. An industrial method for making nitric acid is known as the Ostwald process. In the first step, ammonia and oxygen° gases are heated to produce nitrogen monoxide and steam. In the next step, nitrogen monoxide and oxygen gases combine to produce nitrogen dioxide. In the last step, nitrogen dioxide gas is passed through water to produce nitric acid and nitrogen monoxide gas. Write a balanced equation for each step of the Ostwald process. (4.1, 4.2) K/U T/I C

84. Airbags are used in automobiles to protect drivers and passengers in the case of an accident. They operate using the decomposition of sodium azide, NaN_3. (4.2) K/U
   (a) Write a balanced equation to represent the reaction. Be sure to include symbols to indicate the state of each substance.
   (b) What do you think might be required to start the reaction? Explain.
85. Many kinds of jewellery are made from platinum because it does not corrode. Predict where platinum would be placed on the activity series of metals. (4.4) K/U
86. A laboratory assistant accidentally drops a small piece of potassium into a container of water. (4.4) K/U C
   (a) Predict what type of reaction occurs. Explain your prediction.
   (b) Bubbles appear in the water. What gas are the bubbles formed from?
   (c) Write a balanced chemical equation for the reaction.
87. Baking powder contains both sodium hydrogen carbonate, $NaHCO_3(s)$, and a safe acid such as tartaric acid, $H_2C_4H_4O_6(s)$. When these two compounds are mixed together in a solution, sodium tartrate, $Na_2C_4H_4O_6$, and carbonic acid, H_2CO_3, form. The carbonic acid then breaks down into water and the bubbles of carbon dioxide that make your cake rise. (4.1, 4.2, 4.6) K/U C A
   (a) What two of the four general types of chemical reactions occur during this process?
   (b) Write a balanced chemical equation for each of the two reactions. Underneath each equation, identify what type of reaction it is.
88. Buffalo Mine is an abandoned silver mine in Cobalt, Ontario. The tailings of such mines often contain materials that contaminate groundwater and surface water in the area. The tailings of the Buffalo Mine contain nickel and cobalt, as well as arsenic. (4.6) K/U T/I A
   (a) Use Table 1 in Section 4.6 to identify anions that would likely react with the nickel and cobalt cations in the runoff from the tailings. The anions you choose should form only slightly soluble compounds with most cations.
   (b) Propose a method that could be used to remove nickel and cobalt cations from water so they do not become contaminants.
89. Vinegar, which is dilute acetic acid, is effective at removing hard water deposits from sinks and shower heads. These deposits react with vinegar, producing bubbles of gas. What type of compound do these deposits most likely contain? Explain your answer. (4.6) T/I A
90. Baking soda, $NaHCO_3$, reacts with an acid, $H_2C_4H_4O_6$, in cake batter. One of the products is carbonic acid, HCO_3:

   $$2\ NaHCO_3 + H_2C_4H_4O_6 \rightarrow Na_2C_4H_4O_6 + 2\ H_2CO_3$$

   (4.1, 4.2, 4.4, 4.6) T/I C A
   (a) What type of chemical reaction is shown?
   (b) The carbonic acid then decomposes. Write a balanced equation to represent this reaction.
91. A chemist adds an aqueous solution of manganese(II) sulfate to an aqueous solution of ammonium hydroxide. (4.1, 4.6) K/U T/I C
   (a) What type of chemical reaction occurs?
   (b) Use solubility rules to write a balanced equation to describe the reaction.
92. Canada has deposits of approximately 10 billion tonnes of coal, which is used as a fuel. The combustion of coal provides most of the energy used in North America. (5.1) T/I A
   (a) What are some advantages of using coal as a fuel?
   (b) What are some disadvantages of burning coal?
93. Suppose a home furnace burns propane, $C_3H_8(g)$. The air intake of this furnace contains a lot of dust, and the air flow is reduced. (5.1) T/I A
   (a) Predict what type of combustion is likely to occur in the furnace. Explain your answer.
   (b) What are the most likely products from the furnace?
94. The properties of Teflon make it useful for communications, cookware, and other purposes. How might Teflon be useful in sports? (5.2) T/I A
95. The effect of acid rain on Canadian lakes was first discovered at Lumsden Lake in Ontario. What is a likely pH for a lake affected by acid rain? (5.3) K/U T/I A
96. On October 4, 2010, the Ajka waste sludge reservoir in Hungary broke. One of the substances in the sludge was sodium oxide, Na_2O. Cleanup crews treated the area with ethanoic (acetic) acid. Why was this an effective way to reduce the spill's harmful effects? (5.3, 5.4) K/U T/I A
97. When someone is stung by a bee, part of the resulting pain and swelling is caused by the venom that the bee has injected into the skin. Bee venom is an acidic solution. (5.4) K/U T/I A
   (a) Why should a base be used to treat a bee sting?
   (b) In contrast, wasp stings are basic. What might you use to treat a wasp sting?
   (c) Why might these methods of treating stings not be effective?

98. The northern pike is a fish commonly found in Ontario lakes with a pH ranging from 6.1 to 8.6. (5.4) K/U T/I A
    (a) How could the pH of a lake be adjusted if it is below this range?
    (b) How could the pH of a lake be adjusted if it is above 8.6?
99. Gold is found in ore as metallic gold, not as gold compounds. How might this fact affect the processing of gold ore into gold products? (5.5) K/U T/I A
100. Most metal refineries in Canada are located on the Great Lakes or near a river or an ocean. Why do you think the refineries are located at these places? (5.5) T/I A
101. In 1989, the tanker *Exxon Valdez* struck a reef in Prince William Sound. It spilled hundreds of thousands of barrels of crude oil into the waters of the region. The rugged coastline provided many small inlets in which the oil collected. Cleanup efforts included plans to remove the oil from the shore with high-pressure hot water. However, it was soon discovered that removing the oil also removed natural micro-organisms from the shore. Knowing this, why do you think it was decided to leave the oil on the shore where it was? (5.6) K/U T/I A
102. Making beverage cans from recycled aluminum cans uses about 20 % of the energy it takes to make the cans from aluminum ore. What two principles of green chemistry are used in recycling cans rather than making new ones from ore? (5.7) K/U T/I A

Evaluation

103. Four atoms of phosphorus can produce only 2 molecules of diphosphorus trioxide, even when plenty of oxygen atoms are present. (4.1) T/I C A
    (a) Write the chemical equation describing the reaction.
    (b) Explain why no more than 2 molecules of diphosphorus trioxide are produced for every 4 atoms of phosphorus.
104. A student added sugar to hot water until no more sugar dissolved. As the water cooled, solid sugar appeared. The student stated that, since a solid had formed, a chemical reaction must have occurred between the sugar and the water. (4.1) K/U T/I A
    (a) Was the student correct? Explain your answer.
    (b) What is the only true proof of chemical reaction?
105. A new company is proposing a plasma gasification plant in a rural location in Eastern Canada. The company's proposal suggests that the plant will be completely environmentally friendly because it will not release pollution into the atmosphere and will not depend on fossil fuels. Based on what you know about the process of gasification, evaluate the claims in the proposal. State what additional information you need. (4.3) T/I A
106. A shipyard produces a ship with a steel hull. The manufacturer wants to add a layer of metal to the outside of the hull to protect it from rusting. The metal will be plated onto the hull. The manufacturer has chosen lead. Evaluate the selection and justify your reasoning. (4.4) T/I C A
107. Based on what you know about the complete and incomplete combustion of hydrocarbons, describe three things your family might do at home to prevent carbon monoxide poisoning, assuming that you use fuel-burning devices. (5.1) K/U T/I A
108. Most research scientists work for companies that pay for their research. Do you think that accidental discoveries made while doing research belong to the scientist who made the discovery or to the company for which they work? Support your answer. (5.2) T/I A
109. Acidic oxides are common air pollutants and the source of acid precipitation. Basic oxides are not. Use what you know about these oxides to explain this difference. (5.3) K/U T/I A
110. Gaseous hydrogen chloride, $HCl(g)$, reacts with ammonia gas, $NH_3(g)$, to form ammonium chloride, $NH_4Cl(s)$. (5.4) K/U T/I A
    (a) Is this reaction a neutralization reaction? Explain your answer.
    (b) Think about what forms when hydrogen chloride is added to water. Then think about the product of ammonia and water. How does this reaction differ if both gases are added to water before they react?
111. Examine all methods available for remediation. Which methods can be used to remediate land contaminated by
    (a) metal ions?
    (b) organic compounds? (5.6) K/U T/I A
112. Suppose you use mostly paper products made from recycled paper. (5.7) K/U T/I A
    (a) Which of the principles of green chemistry are included in the making of recycled paper products?
    (b) How does your purchasing choice affect the success of the recycled paper industry?

Reflect on Your Learning

113. What did you learn about the types of chemical reactions that you did not know before? K/U

114. Keep a journal for one day in which you note evidence that chemical changes occurred in yourself or your environment. What clues might you look for after completing this unit that you would not have recognized previously? K/U

115. What did you discover about various methods of getting rid of garbage? What additional questions would you like to have answered about these methods? K/U T/I

116. What did you find most interesting about the research into mercury in Arctic air? Did you find this research important? Why or why not? K/U T/I

117. Think about the products, described in Section 5.2, that were discovered accidentally. T/I A
    (a) Which of the products have you used?
    (b) For each product that you have used, explain how you used it.
    (c) How might your life be different if the importance of these products had not been recognized by research scientists?

118. (a) What are some sources of acidic oxides where you live?
    (b) What is being done to help prevent the release of acidic oxides into the environment where you live?
    (c) From what you know about how these oxides affect the environment, what effect do you expect them to have on the air and water where you live? K/U T/I A

119. The pH of swimming pools and spas must be controlled for the water to be safe for humans. It should be neutral or very slightly basic. K/U A
    (a) What might be done to restore pool water with a pH of 4 to a proper pH?
    (b) What might you add to spa water with a pH of 9 to adjust it to a proper pH?

120. In Figure 9 in Section 5.5, and in many other statistics about air and water pollution, references are made to emissions from the United States as well as emissions from Canada. T/I A
    (a) Explain why air pollutants released by neighbouring countries are important to the air and water quality of any country.
    (b) Why do you think there is less reference to water and land contamination released by neighbouring countries?

121. One common method formerly used for remediation involved burying contaminated soil and other materials. Is this approach true remediation? Explain your answer. K/U T/I A

122. Think of a business or manufacturing facility near your home. Choose two of the principles of green chemistry. Describe how each principle could be applied to make the business or facility better for the environment. K/U T/I A

Research

GO TO NELSON SCIENCE

123. For some chemical equations, more specific state symbols than (s) are required. These might include (cr) and (amor). Research the meanings of these symbols and find examples of when they might be required. Present your examples including chemical equations with the relevant state symbols. Explain what information these state symbols provide about the reaction. T/I C A

124. You are familiar with combustion reactions involving hydrocarbons. Other substances also undergo combustion reactions. Research combustion reactions, identifying several examples. Write a chemical equation for each example. Indicate the pattern common to all of your examples and explain similarities or overlaps with other types of reactions. T/I C A

125. In 1867, the British chemist Henry Deacon developed a method for producing chlorine gas. Research the method now known as the Deacon process to find out the reactants and products of the reaction. Prepare a brief report that explains the importance of this process and the uses of chlorine gas. T/I C A

126. More than 400 Canadians died from carbon monoxide poisoning between the years 2000 and 2007.
   (a) On the Internet, find a scenario in which carbon monoxide was produced by the combustion of a fuel. Why was carbon monoxide produced instead of carbon dioxide?
   (b) Research and explain how a carbon monoxide detector works.
   (c) What are some symptoms of carbon monoxide poisoning? How can carbon monoxide poisoning be treated? K/U T/I A
127. Some forms of research have as their goal the development of an effective product to be used for a specific purpose. Sometimes new products are produced accidentally during the research. Sometimes the goal product is produced and is then repurposed for uses other than the intended purpose of the research. T/I A
   (a) Find out about the development of the adhesive used on Post-it Notes. Was it an accidental discovery or a repurposing of an existing product?
   (b) Research the development of CorningWare cookware. Was it an accidental discovery or a repurposing of an existing product?
128. Buffer solutions are an interesting application of neutralization reactions. T/I A
   (a) What is a buffer solution?
   (b) Why are buffer solutions important to human health?
   (c) What is the main buffer solution in human blood?
   (d) How does this buffer solution control the pH of the blood?
129. All packing peanuts used to be made from polystyrene foam, which is a petroleum product. Many packing peanuts are now made from renewable resources. T/I A
   (a) What renewable resources are now being used to make packing peanuts?
   (b) How does the nature of this resource affect the disposal of the peanuts?
130. Mining is an important industry for Canada because of Canada's rich supply of natural resources. Canada is a worldwide leader in mining gold, among other minerals. K/U
   (a) A new open-pit gold mine is currently being dug in Malartic. Where is this mine?
   (b) What are some benefits of the mine for the community?
   (c) What are some concerns the community members have about the mine being dug?
   (d) Will the ore from the new mine be processed in a mill or a refinery? Will most of the processing of the ore be physical or chemical? Explain your answer.
131. On October 4, 2010, the Ajka waste sludge reservoir in Hungary broke. The sludge released was formed from the wastes produced by a factory that refines aluminum metal from aluminum ore. Research this environmental disaster and find answers to the following questions: K/U T/I A
   (a) How much sludge was released?
   (b) What were the main components of the sludge?
   (c) Which were potentially more harmful to the environment—the main components of the sludge or the minor components? Justify your answer.
   (d) What has been done to remediate the affected area?

UNIT 3

Quantities in Chemical Reactions

OVERALL EXPECTATIONS

- analyze processes in the home, the workplace, and the environmental sector that use chemical quantities and calculations, and assess the importance of quantitative accuracy in industrial chemical processes
- investigate quantitative relationships in chemical reactions, and solve related problems
- demonstrate an understanding of the mole concept and its significance to the quantitative analysis of chemical reactions

BIG IDEAS

- Relationships in chemical reactions can be described quantitatively.
- The efficiency of chemical reactions can be determined and optimized by applying an understanding of quantitative relationships in such reactions.

UNIT TASK **PREVIEW**

Alka-Seltzer is a popular antacid product used to relieve the discomfort caused by excess stomach acid. The active ingredient is sodium hydrogen carbonate, or baking soda. Your task is to experimentally compare the mass of baking soda in each tablet with the mass indicated on the product label.

The Unit Task is described in detail on page 352. As you work through the unit look for Unit Task Bookmarks to see how information in the section relates to the Unit Task.

FOCUS ON **STSE**

GREENING CHEMICAL PROCESSES

Manufacturing pharmaceuticals is a multibillion-dollar global industry. The production of pain relievers involves a series of chemical reactions. The products of one reaction are collected, purified, and then passed on to the next reaction and so on. The company strives to reduce waste and keep manufacturing costs down. One way to cut down on waste is to use reactions that produce more of the desired product and less waste products. This strategy helps to keep production costs down and maximize the company's profits.

Minimizing waste is one of the principles of green chemistry. Remember that a green manufacturing process is designed to have a minimal impact on the environment. In other words, the process is "benign by design."

Ibuprofen was discovered by a British pharmaceutical company in the 1950s. The company developed a process called the Brown synthesis. This process was adequate for producing small quantities of the drug, but was very wasteful for larger quantities. Six steps were required to convert the starting materials into the final product. During each step, some starting materials were unavoidably lost. Consider the analogy of preparing a meal. You will have less waste if you cook all the ingredients in one pot rather than transferring them between several different pots. The net result for the Brown synthesis was that only about 40 % of the atoms in the starting materials became part of the ibuprofen molecules. This meant that 60 % of the atoms in the reactants were "waste." For every tonne of ibuprofen produced, about 1.5 t of waste was generated.

In the mid-1980s, ibuprofen was released in North America. We see it on the pharmacy shelves as Advil and Motrin. To keep up with demand, the American pharmaceutical company BHC started researching a greener and more efficient way to make ibuprofen. The company identified and tested different chemical reactions. They were looking for a process with fewer steps in which the percentage conversion of reactants into products was high. The end result was a process that used only three steps. In this new, simpler process, 95 % of the atoms in the starting materials made it into ibuprofen. The production of ibuprofen is an example of how a production process can be environmentally benign and even more financially profitable.

Questions

1. Give two reasons why it is in a manufacturer's best interest to become greener.
2. What effect would reducing the number of steps in a manufacturing process likely have on a manufacturer's energy consumption? Why would this benefit the environment?
3. Why is it important to develop a disposal strategy when designing a new chemical process?
4. What happens to old medications? How might their disposal affect the environment?

CONCEPTS

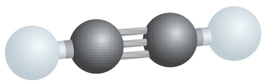
- relative atomic mass of a compound
- law of conservation of mass
- predicting products of displacement reactions using an activity series
- structural formulas of compounds
- the difference between ionic and molecular compounds

SKILLS

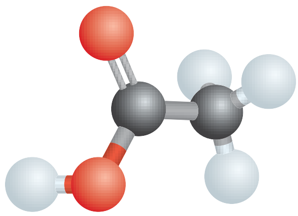
- name or write the chemical formulas of compounds
- balance chemical equations
- solve mathematical problems involving percentages, ratios, and proportions
- use correct units and the correct number of significant digits when expressing the answer to a calculation
- plan and safely conduct an investigation
- critically evaluate the results of an investigation

Concepts Review

1. (a) Determine the total number of each type of atom in the molecules modelled in **Figure 1**. (In these models, carbon atoms are black, oxygen atoms are red, and hydrogen atoms are white.)
   (b) Write the chemical formula for each compound. K/U C



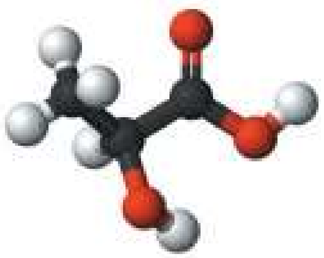
(i) ethyne (acetylene)



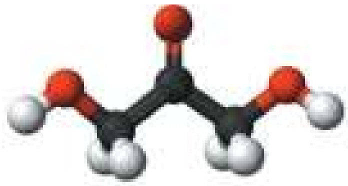
(ii) ethanoic (acetic) acid

Figure 1 Models of two molecular compounds

2. Look at the ball-and-stick models of two molecules in **Figure 2** and answer the following questions: K/U T/I C



(i) 2-hydroxypropanoic acid (lactic acid)



(ii) 1, 3-dihydroxypropan-2-one (dihydroxyacetone)

Figure 2 In these two molecular models, carbon atoms are shown in black, oxygen in red, and hydrogen in white.

   (a) Write the molecular formula of each molecule.
   (b) Are these two molecules of the same compound? Explain your answer.

3. The mass of a proton is almost the same as the mass of a neutron: 1 u. In comparison, the mass of an electron is negligible. Why, then, are the masses of the elements given on the periodic table not whole numbers? K/U

4. Classify each of the following compounds as ionic or molecular. Justify your choices. K/U
   (a) $CaCl_2$
   (b) OCl_2
   (c) CCl_4
   (d) K_2SO_4
   (e) $NaHCO_3$
   (f) H_2SO_4

5. Balance the following chemical equations: K/U C
   (a) $KClO_3 \rightarrow KCl + O_2$
   (b) $P_4 + O_2 \rightarrow P_2O_5$
   (c) $Al_2(SO_4)_3 + Ca(OH)_2 \rightarrow Al(OH)_3 + CaSO_4$
   (d) $C_2H_2 + O_2 \rightarrow CO_2 + H_2O$
   (e) $Fe + Al_2(SO_4)_3 \rightarrow FeSO_4 + Al$
   (f) $H_2O_2 \rightarrow H_2O + O_2$

6. Predict the products of the following chemical reactions. If no reaction occurs, write NR. K/U C
   (a) $Cu(s) + ZnSO_4(aq) \rightarrow$
   (b) $Mg(s) + AgNO_3(aq) \rightarrow$
   (c) $Fe(s) + Al(NO_3)_3(aq) \rightarrow$
   (d) $Cu(s) + Pb(NO_3)_2\ (aq) \rightarrow$
   (e) $Pb(s) + 2\ HCl(aq) \rightarrow$

7. **Figure 3** (on the next page) shows the formation of oxygen and water from hydrogen peroxide. K/U T/I C
   (a) Write the chemical reaction for the equation.
   (b) Classify the type of reaction.
   (c) If 6.8 g of hydrogen peroxide reacts to produce 3.2 g of oxygen, what mass of water is also produced? Explain.
   (d) If 13.6 g of hydrogen peroxide reacts, predict the masses of the products. Justify your prediction.

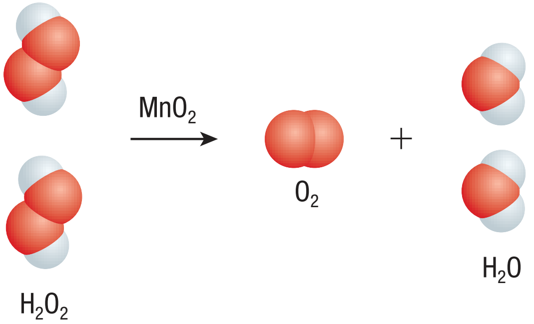


Figure 3 Decomposition of hydrogen peroxide

Skills Review

8. Name the following compounds: K/U
   (a) $Ba(ClO_3)_2$ (d) $Mn_3(PO_4)_2$
   (b) $SnSO_3$ (e) $Pb(NO_3)_2$
   (c) $FeCl_3$ (f) P_4O_{10}
9. Write the chemical formula for each of the following compounds: K/U C
   (a) copper(II) nitrate
   (b) carbon tetrachloride
   (c) carbon monoxide
   (d) iron(III) chlorate
   (e) calcium hydrogen carbonate
10. Complete the following word equations in your notebook, then write a balanced chemical equation for each reaction: K/U C
    (a) copper(II) oxide + hydrogen gas →
    (b) lead(II) nitrate + potassium iodide →
    (c) barium chloride + sodium sulfate →
    (d) sodium phosphate + calcium chloride →
11. There are 520 boys in a school of 988. K/U T/I
    (a) Calculate the percentage of boys in the school.
    (b) Determine the percentage of girls in the school using two different methods.
12. Rearrange the formula $A = \frac{B}{C}$ to solve for B and then for C. T/I C
13. Classify the following quantities as either exact (counted or defined) or measurements: K/U
    (a) There are 25 students in the class.
    (b) She poured 48.6 mL of water into a graduated cylinder.
    (c) The gravel truck hauled 5000 kg of gravel.
    (d) The subscripts in Al_2O_3 are 2 and 3 respectively.
    (e) One kilogram contains 1000 g.
    (f) The density of gold is 19.3 g/cm^3.
14. A laboratory balance is used to measure the mass of a sample of iron(III) oxide (**Figure 4**). T/I

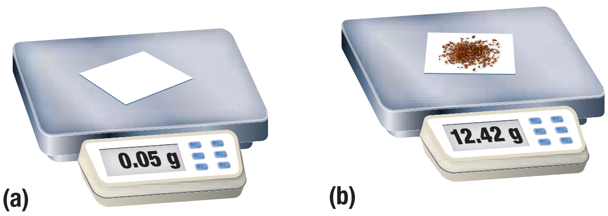


Figure 4 (a) A balance with a piece of massing paper on it (b) The same balance with a sample of iron(III) oxide on the massing paper

    (a) What is the mass of the iron(III) oxide?
    (b) Why should chemicals not be placed directly on the balance pan?
    (c) A student places the same piece of massing paper on the scale and then tares ("zeros") the scale so that 0.00 appears on the display. What would be the reading on the scale after adding the same quantity of iron(III) oxide as shown in Figure 4?
    (d) What mass would the scale display if the paper and iron(III) oxide were then removed? Why?
15. Perform the following calculations. Express the answer using the correct units and number of significant digits. T/I
    (a) $(2\text{ beakers})\left(\dfrac{45.2\text{ g}}{\text{beaker}}\right) =$
    (b) $\dfrac{9.25 \times 10^{-3}\text{g}}{4.6 \times 10^{-3}\text{mL}} =$
    (c) $(150\text{ molecules})\left(\dfrac{3\text{ atoms}}{\text{molecule}}\right) =$
    (d) $\dfrac{5.64 \times 10^{-3}\text{ g}}{\dfrac{1.2\text{ g}}{\text{L}}} =$
    (e) $\dfrac{\dfrac{1.92 \times 10^{-24}\text{ g}}{\text{atom}}}{\dfrac{12.1\text{ u}}{\text{atom}}} =$
    (f) $\dfrac{5.24\text{ g}}{\dfrac{4.1\text{ g}}{\text{mL}}} =$

CAREER PATHWAYS **PREVIEW**

Throughout this unit you will see Career Links in the margins. These links mention careers that are relevant to Quantities in Chemical Reactions. On the Chapter Summary page at the end of each chapter you will find a Career Pathways feature that shows you the educational requirements of the careers. There are also some career-related questions for you to research.

CHAPTER 6

Quantities in Chemical Formulas

KEY CONCEPTS

After completing this chapter you will be able to

- accurately analyze the quantities of chemicals in our bodies, in various foods, and in daily use
- appreciate that accurate measurements and analyses are crucial to our health and environment, as well as to industrial and manufacturing processes
- understand and use the law of definite proportions
- calculate quantities of atoms, ions, molecules, and other items using the unit "mole"
- calculate percentage composition, empirical formulas, and molecular formulas from laboratory data
- perform calculations involving atomic mass, molar mass, mass, and amount of substance

Why Is Accuracy Important when Measuring Quantities of Chemicals?

"It's the dose that makes the poison."

This statement, by the sixteenth-century Swiss physician Paracelsus, means that any substance can be toxic if you ingest, inhale, or absorb enough of it. A moderate dose of caffeine is not harmful. A much larger dose consumed in a short period of time can, however, be fatal. In 2000, an 18-year-old Irish basketball player died during a game after consuming four cans of an energy drink that contained caffeine. As a result, some European countries banned energy drinks for several years.

If a small dose is not harmful but a large dose can kill a person, it is very important to know what quantity is dangerous. Toxicity describes the degree to which a substance is poisonous to a living organism. Toxicity can be acute or chronic. Acute toxicity describes the effect of a single exposure to the substance. Chronic toxicity describes the effect of prolonged exposure. Scientists measure acute toxicity using an LD_{50} value. The LD_{50} (lethal dose for 50 %) represents the mass of the chemical required to kill 50 % of a test population. The population is often a group of laboratory animals such as rats or brine shrimp. There are LD_{50} data available for thousands of chemicals. For example, acrylamide, C_3H_5NO, forms when potatoes are deep-fried in very hot oil. The LD_{50} for this compound is 50 mg/kg for lab rats—almost as toxic as nicotine. This quantity means that, in a group of 0.5 kg rats, half would die if they each ingested 25 mg of acrylamide.

We cannot assume that just because a certain dose is lethal to rats, the same dose—accounting for our larger size—would be lethal to humans. Our bodies process chemicals differently. However, the low LD_{50} value of acrylamide serves as a useful warning. It indicates that a very small quantity may be dangerous. In fact, the health risk of acrylamide to humans remains unclear. To minimize your risk for acrylamide, Health Canada recommends reducing your intake of deep-fried foods.

"It's the dose that makes the poison" also applies to medications. The treatment of cancer with drugs is often called "chemotherapy." Many "chemo" drugs are effective because they kill cancer cells. However, they also kill normal, healthy cells, including infection-fighting white blood cells. This is why patients undergoing chemotherapy for cancer are often susceptible to infections. Calculating the correct dosage and duration of chemotherapy is critical to its success.

STARTING **POINTS**

Answer the following questions using your current knowledge. You will have a chance to revisit these questions later, applying concepts and skills from the chapter.

1. Why do you think more than one person checks the drug dosage calculations before a hospital pharmacy fills a prescription?
2. Why is it necessary for chemists to round off the result of a calculation in the correct way?
3. Government agencies establish maximum allowed quantities for environmental contaminants. Why do you think it is necessary to review these quantities periodically?
4. Why is it important to consider chronic toxicity as well as acute toxicity when deciding whether chemicals should be banned?

Mini Investigation

Comparing Salty Snacks

Skills: Performing, Observing, Communicating

SKILLS HANDBOOK A1.2, A2.2, A2.3, A3.2

In this investigation you will use a chemical test to compare the salt content of common foods.

Equipment and Materials: chemical safety goggles; lab apron; gloves; 4 test tubes; test-tube rack; scoopula; balance; 10 mL graduated cylinder; dropper bottle of distilled water; 1 cm slice of hot dog chopped into small pieces; sample of crushed potato chips; sodium chloride, $NaCl(s)$; massing paper; dropper bottle of silver nitrate, $AgNO_3(aq)$ (0.1 mol/L)

Silver nitrate stains the skin. If you spill silver nitrate on your skin, wash the affected area with lots of cool water and inform your teacher.

Never consume anything in the lab or remove food from the lab to consume later.

1. Put on your safety goggles, lab apron, and gloves.
2. Label the test tubes: 1, 2, 3, and 4.
3. Fill one-third of each test tube with warm distilled water supplied by your teacher.
4. Add the pieces of chopped hot dog to the first test tube.
5. Add a comparable volume of crushed potato chips to the second test tube.
6. Add 0.25 g of sodium chloride to the third test tube.
7. Leave the fourth test tube as the control, with nothing added.
8. Add 3 mL of distilled water, followed by 5 drops of silver nitrate solution, to each test tube. Record your observations.

A. Write a balanced chemical equation for the reaction that occurred in test tube #3. K/U C

B. Compare the appearance of the precipitate in the four test tubes. Relate your observations to the salt content of each sample. T/I

C. Predict how your observations would differ if raw potato were used in place of the crushed chips. T/I

D. Identify the controlled variables in this investigation. T/I

E. Why was it necessary for the foods to be in small pieces? T/I

F. Research your daily recommended intake of sodium. T/I

GO TO NELSON SCIENCE

6.1 Qualitative and Quantitative Analysis

Figure 1 Unfortunately, many patients going to hospital for treatment fail to correctly report the medications that they are taking. Doctors need this information when planning treatment.

Headaches, fatigue, and blurred vision—Sonja sensed that her blood sugar level was out of balance again. Sonja has diabetes, a condition that prevents the body from producing or properly using insulin. Insulin is a hormone that is needed to process glucose. As you learned in Section 2.4, glucose is a simple sugar that fuels muscle activity. Insulin helps body cells absorb glucose and keeps its concentration in blood steady. Excess blood glucose increases the risk for heart attacks, stroke, and blindness. The Public Health Agency of Canada reports that in 2005, 1 in 17 Canadians had diabetes.

Sonja's body does not produce enough insulin on its own. Instead, the medication she takes daily makes her pancreas increase its output of insulin. Unfortunately, because of flight delays during her trip in the United States, Sonja ran out of medication. She had to see an American physician to have her prescription refilled (**Figure 1**).

"What's your normal blood glucose level?" asked the doctor. "About 4.6 before meals," she responded. The doctor looked puzzled. He had expected a number about 20 times larger. "Are you sure?" he asked. "With a number that low, you should be dead."

Then he saw Sonja's home address in Canada. That explained the discrepancy in glucose levels. In Canada and Europe, blood glucose levels are measured in different units than in the United States. With this information, the doctor was able to understand Sonja's response.

Appropriate Dosage

Sonja's case illustrates the importance of accurately measuring and communicating measured quantities in an analysis. This is particularly important in medicine, where correct dosages of medication can mean the difference between life and death. Doctors usually calculate the appropriate dosage of a medication based on the patient's mass. However, other qualitative factors such as age, lifestyle, and medical history are also considered. These factors explain why two similar people with the same medical condition may be treated with different doses of the same medication.

Since the professional judgment of the doctor plays an important role in determining dosage, errors can occur. However, far more medication errors result from patients' misuse of their medications. These errors range from not taking the medication to consuming alcohol or illicit drugs while taking medication. The Ontario Ministry of Health reports that improper medication use by patients accounts for about one-third of all emergency room visits. This places a huge burden on Ontario's healthcare system. Patient misuse of medications can even be fatal. Combining medications with alcohol or illicit drugs also causes many deaths every year in Ontario.

Qualitative and Quantitative Analysis

qualitative analysis the process of identifying substances present in a sample; no measurements are involved

Qualitative analysis is the identification or detection of a specific substance. Using a urine test strip to detect excess glucose in urine is an example of qualitative analysis (**Figure 2**). Chemicals at the end of the strip change colour when dipped in urine. A dark colour indicates that there is too much glucose in the blood. This is considered a positive result. A paler colour indicates that normal levels of glucose are present, which is a negative result. Data collected during this type of analysis are called qualitative data.

Many qualitative analyses in chemistry involve colour changes. You have probably used an acid–base indicator to detect the presence of an acid or a base. This is another example of qualitative analysis.

The presence of excess glucose in blood can be a sign of diabetes. Diabetics use a device called a glucometer to regularly monitor their blood sugar levels. Typically, a tiny needle pricks the finger to produce a drop of blood. The glucometer absorbs the blood, measures the blood glucose concentration, and reports this value on a digital display (**Figure 3**). An analysis that provides numerical data is called a **quantitative analysis**. The use of a pH meter to measure changes in pH during a chemical reaction is an example of quantitative analysis. Numerical data like blood sugar concentration and pH are examples of quantitative data.

Quantitative analysis is useful in many areas beside medicine. In Unit 4 you will learn the skills necessary to conduct a quantitative analysis of some consumer products.

quantitative analysis the process of measuring the quantity of a substance in a sample, providing numerical data

CAREER **LINK**

Medical technologists perform tests on samples taken from patients. In this way they help doctors and other medical professionals in their diagnosis of diseases. To find out more about the training necessary to work as a medical technologist,

GO TO NELSON SCIENCE

Figure 2 The dark colour of a glucose urine test strip is a positive test for excess glucose in urine. Diabetic patients have abnormally high concentrations of glucose in their urine.

Figure 3 A glucometer gives a quantitative measure of the blood glucose level. A drop of blood is required for the analysis.

Mini **Investigation**

Testing for Sugar

Skills: Controlling Variables, Performing, Observing, Analyzing, Evaluating, Communicating

SKILLS HANDBOOK **A1, A2.4, A3.3**

In this activity, you will conduct qualitative tests for glucose. You will test two different glucose solutions as well as a starch suspension. Starch is made up of glucose molecules linked together in long chains. The indicator that you will use to test the solutions is called Benedict's solution.

Equipment and Materials: chemical safety goggles; lab apron; 10 mL graduated cylinder; 4 test tubes; glassware marker or labels; test-tube rack; kettle; test-tube clamp; 250 mL beaker; dropper bottle of distilled water; small beakers containing samples of 1 % starch suspension, 1 % glucose solution, 10 % glucose solution, and Benedict's solution

Benedict's solution contains sodium carbonate, which is an irritant. It also contains copper(II) sulfate, which is poisonous if ingested. Avoid skin and eye contact. If you spill Benedict's solution on your skin, wash the affected area with lots of cool water and inform your teacher.

This investigation involves the use of very hot water. Be careful not to spill the water. If you spill the water on your skin, run cold water over the area and inform your teacher.

1. Put on your chemical safety goggles and lab apron.
2. Label the test tubes 1 to 4 and place them in the test-tube rack.
3. Add liquids to the test tubes as described in **Table 1**. Rinse the graduated cylinder after each solution.

Table 1 Contents of Four Test Tubes

Test tube #	Liquid
1	5 mL distilled water
2	5 mL 1 % starch suspension
3	5 mL 1 % glucose solution
4	5 mL 10 % glucose solution

4. Add 5 mL of Benedict's solution to each test tube.
5. Your teacher will heat tap water in the kettle until it is almost boiling, and half fill your beaker with hot water.
6. Place the test tubes in the beaker of hot water for about 5 min.
7. Use a test-tube clamp to carefully remove the test tubes from the beaker and place them in a test-tube rack. Record the colour of each of the liquids.

A. Compare the colours of the two glucose solutions. Why was there a difference? T/I

B. Suggest an explanation for why the colour of test tube #2 was different from the colour of test tubes #3 and #4. T/I

C. Why was it necessary to include test tube #1? T/I

D. Diabetics can often control their symptoms by adjusting their diets. Research to find out how eating appropriately can help diabetics. T/I

GO TO NELSON SCIENCE

Figure 4 An air quality technician measures pollutants in vehicle exhaust.

Quantitative Analysis and the Environment

All human activities release substances into the environment. We use quantitative analysis to monitor whether these substances are kept within acceptable concentrations (**Figure 4**). (You will find out a lot more about water-borne pollutants in Unit 4 and airborne pollutants in Unit 5.)

INDOOR CARBON DIOXIDE LEVELS

You release carbon dioxide into your environment every time you exhale. As you can appreciate, 30 students in a classroom can produce a considerable quantity of carbon dioxide in a 70-minute period. Excess carbon dioxide can be a problem, especially in newer, airtight schools. Even though carbon dioxide is not toxic, high carbon dioxide levels can impair student reaction times and memory. Teachers sometimes observe that their students are drowsier than normal, although this is likely to be caused by other factors.

Many workplaces routinely monitor indoor levels of carbon dioxide. Carbon dioxide levels are considered a good indicator of whether or not ventilation is adequate. If carbon dioxide levels are high, then levels of other contaminants may also be dangerously high.

POLLUTANTS AND HEALTH

Quantitative analysis is also important when searching for links between human health problems and pollutants. For example, a growing number of scientists believe that diabetes may have an environmental connection. These suspicions first arose after a 1976 accident in Seveso, Italy. An explosion at a chemical plant contaminated the local area with a toxic, carcinogenic chemical called dioxin. Dioxin is one of a group of toxic compounds called persistent organic pollutants (POPs). Other examples of POPs are DDT, a pesticide, and PCBs, industrial chemicals once used in electrical transformers. These toxins are stable and relatively unreactive, so they persist in the environment for years without decomposing. Predictably, the people around Seveso are dying prematurely from cancer. However, an unusually high number also die as a result of diabetes.

Mortality data from Seveso led some researchers to question whether there is a link between POPs and diabetes. To test this question, researchers conducted a study in which they analyzed blood samples from 2000 American participants for POPs. The results supported the researchers' prediction: people with the highest blood levels of POPs were most likely to have diabetes. In fact, the correlation between high POP levels and diabetes was even stronger than the correlation between obesity and diabetes. However, the correlation between POPs and diabetes does not prove that POPs cause diabetes. In order to establish proof, additional studies are necessary.

CAREER LINK

Blood and urine samples taken from athletes are analyzed for banned substances by analytical chemists. They play a cat-and-mouse game with cheating athletes, trying to develop chemical tests to identify and measure concentrations of banned substances. To find out more about becoming an analytical chemist,

GO TO NELSON SCIENCE

Quantitative Analysis in Sport

Many organizations are dedicated to keeping performance-enhancing substances out of competitive sports. Quantitative analysis plays a key role in analyzing samples of urine or blood, taken from athletes, for these substances. Accurate measurements are critical in these analyses. An athlete's career may hinge on the outcome.

As drug-testing technologies have improved over time, strategies used by the drug users have also improved. The result has been an elaborate "cat and mouse" game in which advances by drug testers are soon countered by the drug users.

DOPING IN CYCLING

World-class cycling has a notorious history of doping—using substances to gain an unfair advantage over other competitors. Until the 1970s, there were virtually no doping regulations. The frequency of testing increased somewhat after the death of a British cyclist during the 1967 Tour de France. His death was caused by overuse of an amphetamine—a stimulant. By the 1990s, chemical detection methods had improved considerably. In order to avoid detection, some cyclists turned to biological methods to gain an advantage.

One such biological method is the use of erythropoietin (EPO). EPO is a naturally occurring hormone that stimulates the bone marrow to produce red blood cells.

Some cyclists found that EPO, and a similar substance called r-EPO, could enhance their performance. More red blood cells can deliver more oxygen to the muscles. Since erythropoietin is a naturally occurring substance, it is difficult to detect. However, its use increases a measureable characteristic of blood: the percentage by volume of red blood cells in the blood. This percentage is called the hematocrit level (HCT).

Endurance athletes such as cyclists can increase their HCT by training. An elite cyclist may have an HCT around 50 %. The use of r-EPO can push this value above 60 %. At these levels, the blood becomes so viscous that it is prone to clotting. In 1997, the international body that governs world cycling ruled that the HCT of cyclists could not exceed 50 %. However, some doped cyclists soon found a way around this rule. Immediately prior to testing, they would inject themselves with a salt solution. This would dilute the blood, thus reducing HCT to acceptable levels.

WEB LINK

To find out more about EPO, HCT, and endurance sports,

GO TO NELSON SCIENCE

6.1 Summary

- Qualitative analysis involves detecting or identifying a substance. No measurements are involved.
- Quantitative analysis involves measuring a specific quantity or characteristic of a substance.
- The accuracy of quantitative data is critical to many medical and environmental processes.

6.1 Questions

1. Classify the following observations as examples of qualitative or quantitative data: K/U
   (a) Rainwater has a pH of 5.6.
   (b) Sonja's blood glucose concentration was 4.6 mmol/L.
   (c) Carbon monoxide poisoning changes the colour of blood to cherry red.
   (d) The addition of acid to a cleaning product produces a gas with a bleach-like odour.
   (e) A glowing splint lights in a container of gas, confirming that the gas is oxygen.
   (f) A recipe requires 1.0 g of baking powder for each cup of flour.
2. Doctors, when they prescribe medications, give very specific instructions about taking the correct dosage—the right quantity of medicine at the right time. K/U
   (a) List two factors that doctors take into account when deciding on the dosage.
   (b) List two reasons why the medication could cause serious illness.
3. There are often several versions of medications, such as pain relievers, that you can buy without a prescription. What do you think are two important sources of information available at a drug store to advise which product is right for you? A
4. Imagine an electronic health record system in which every medical procedure and prescription drug that you have ever taken is documented and the information is available to your doctor. What would be an advantage of such a system? A
5. Do you think that all professional athletes should be subject to regular doping tests? Why or why not? A
6. Researchers have studied the brains of people who suffered from Alzheimer's disease and those of people without Alzheimer's, after death. The researchers found that the brains of Alzheimer's patients contained higher concentrations of aluminum than did the brains of the general population. Does this mean that aluminum causes Alzheimer's? Explain your answer in a paragraph. A
7. (a) Research the factors that increase the risk of developing diabetes. Summarize these factors in a checklist.
   (b) What lifestyle changes could you take to decrease your risk of developing diabetes?
   (c) Could you be certain that these lifestyle changes would prevent you from getting diabetes? Explain. T/I C A

GO TO NELSON SCIENCE

6.2 Explore an Issue in Chemical Quantities

SKILLS MENU

- Defining the Issue
- Researching
- Identifying Alternatives
- Analyzing
- Defending a Decision
- Communicating
- Evaluating

Overdosing on Salt

Salt is a silent killer. Millions of people worldwide die each year from complications caused by consuming too much salt. What makes these deaths so unfortunate is that most are preventable. Prevention starts with saying "no" to the salt shaker and making wiser choices at mealtime. A World Health Organization study predicts that a reduction of only 15 % in daily salt intake over a 10-year period would save 9 million lives. This is just a few grams less, per person, each day.

As you recall, table salt is mostly sodium chloride, NaCl: an ionic compound. The sodium ions in sodium chloride are the real cause of the problem. Sodium is a required nutrient. Sodium ions help regulate fluid levels in the body and allow nerves and muscles to operate effectively. However, excess sodium ions make the body retain water. Extra fluid increases blood pressure, which makes it more difficult for the heart to push blood throughout the body (**Figure 1**). This may damage the heart and increase the risk of heart attacks and cardiovascular disease.

Figure 1 A typical healthy person has a blood pressure of 120/80 mmHg (millimetres of mercury). Excess dietary salt can elevate blood pressure.

According to Health Canada, the average Canadian consumes about 3400 mg of sodium per day. This quantity is more than twice Health Canada's recommended daily intake of 1500 mg, and almost 50 % greater than the maximum tolerable limit of 2300 mg. These quantities may sound large, but they represent much less than a teaspoonful of salt (**Figure 2**).

Figure 2 (a) A Canadian's average daily consumption of salt and (b) the recommended daily intake

Using the salt shaker less frequently is one way to reduce your salt intake. However, a French study showed that only about 10 % of dietary salt is added during home cooking or at the table. The remaining 90 % is hidden in the processed foods that we eat. That is why it is so important to read product labels when choosing foods, and to choose fresh foods rather than processed products (**Figure 3**).

A recent study of Canadian Inuit verified a link between processed foods and high blood pressure. The traditional Inuit diet of fish and sea mammals is low in sodium. Consequently, the incidence of high blood pressure among Inuit has historically been below that of the general Canadian population. However, as Inuit consume more processed, store-bought foods, they are increasingly suffering from high blood pressure.

Figure 3 This meal contains most of the entire recommended daily intake of salt.

The Issue

As more students adopt healthier lifestyles, it is important that they are aware of what is in the food they consume. This applies to the school cafeteria as well as in the home.

Goal

To develop a strategy to reduce salt consumption at school (including food prepared at home or in the cafeteria)

Research

SKILLS HANDBOOK A5.1

Work in pairs or in small groups to learn more about dietary salt. Research the following topics:

- other medical conditions that are linked to an excess of dietary salt
- reasons that salt is added to food
- types of foods that are particularly high in salt
- some simple strategies that can reduce dietary salt intake
- other sources of sodium in food, beside table salt
- alternatives to salt, and their health effects

Identify Solutions

Consider a variety of strategies to find the best way of reducing students' salt intake in the school cafeteria. They could include the following:

- raising student awareness of the health risks of excess salt intake
- posting information about the nutrient content of cafeteria food
- discussions with the cafeteria management about lowering the salt content of cafeteria food
- increasing the student demand for healthier food choices

Make a Decision

Based on your discussions and research, what approach would you recommend for bringing about a reduction in the salt intake in the school cafeteria?

Communicate

Summarize your recommendations in a report that forms the guidelines for your school's action plan to "halt the salt"—or at least to reduce it.

CAREER LINK

Food chemists are involved in deciding how much salt, and other additives, is included in processed foods. If you want to know more about the career of a food chemist,

GO TO NELSON SCIENCE

WEB LINK

To start your research on salt in the diet,

GO TO NELSON SCIENCE

Plan for Action

How could you get your action plan implemented in your school? One way might be to collect evidence, showing how two similar (and great-tasting) meals can contain very different quantities of salt.

1. Plan two lunches: one consisting mainly of processed foods and another of mostly unprocessed foods. Try to make the meals as comparable as possible. For example, if you select breaded chicken strips in the processed lunch, use a similarly sized portion of chicken breast in the other. (No purchases are necessary.)
2. Use product labels or other sources of nutrition data to calculate the total mass of sodium, in milligrams, in each meal.
3. Prepare to implement your salt reduction action plan. Decide whom you need to convince, to put your plan into action. Have your evidence on hand to support your proposal. Always discuss your plans with your teacher first. Launch a campaign to make your school more "salt savvy."

GO TO NELSON SCIENCE

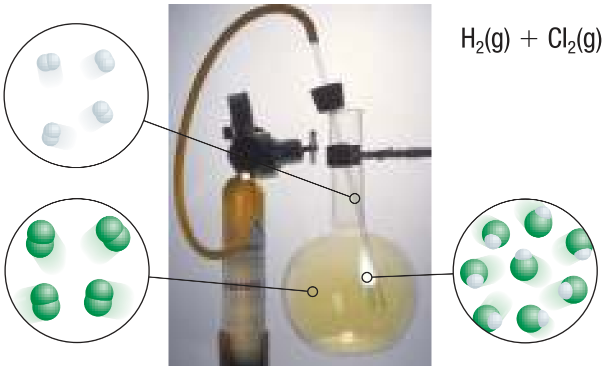
6.3 The Mole—A Unit of Counting

How is a balanced chemical equation like a recipe? A balanced chemical equation tells not only the ingredients required (the reactants), but also the proportions in which they react. Consider, for example, the equation for the synthesis of hydrogen chloride from its elements. The equation tells us that 1 molecule of hydrogen reacts with 1 molecule of chlorine to produce 2 molecules of hydrogen chloride:

$$H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g)$$

Hydrogen atoms and chlorine molecules are incredibly small: too small to be seen even with a powerful microscope. The reaction in **Figure 1**, for example, represents an astronomically large number of reacting molecules. And yet, chemists can combine just the right quantities of hydrogen and chlorine so that, when they react, both substances will be completely used up. How do you count entities as small as atoms and molecules? Counting them directly, as you would count students in a classroom, is not technically possible. And even if it were possible, there are more molecules in the flask in Figure 1 than you could count in a lifetime. Therefore, chemists need a more convenient way to determine how many entities (such as atoms or molecules) are in a sample of matter.

> **LEARNING TIP**
>
> **Entity**
> The word "entity" is a general term used to refer to particles of matter. Atoms, molecules, and ions are all entities that make up matter.



$H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g)$

Figure 1 Hydrogen gas (from the cylinder) reacts with chlorine gas (in the flask) to form hydrogen chloride. When hydrogen and chlorine molecules are combined in a 1:1 ratio, all of the reactant atoms are used to form products.

Mini **Investigation**

Counting Estimates

Skills: Performing, Observing, Analyzing, Communicating

SKILLS HANDBOOK A1.2, A3.2

How do you count things that are unimaginably small, such as atoms? In this activity you will see how measurement can be used to estimate the number of identical objects or entities in a given quantity of matter.

Equipment and Materials: balance; small plastic cup or beaker; 50 dry beans; 50 grains of dry rice

Do not consume anything in the laboratory, or remove anything to consume later.

1. Measure the mass of the empty cup.
2. Measure the mass of one bean.
3. Add 10 beans to one cup. Measure their total mass.
4. Add another 40 beans to the cup. Measure their total mass.
5. Repeat Steps 2 to 4 with grains of rice.

A. Determine the average mass of 1 bean in a sample of 10 beans (Step 3) and in a sample of 50 beans (Step 4). T/I

B. Determine the average mass of 1 grain of rice in each sample, as you did for the beans. T/I

C. How would you use your mass data to determine the number of beans in a 1.0 kg bag of beans? Which mass would you use: the mass of an individual bean, the average mass of 10 beans, or the average mass of 50 beans? Why? T/I

D. Use your data to estimate how many beans are in a 1.0 kg bag of beans. T/I

E. Use your data to estimate how many grains of rice are in a 1.0 kg bag of rice. T/I

F. Explain why your answers to D and E are estimates and not exact values. K/U

Using Mass to Estimate Numbers

In the Mini Investigation you indirectly determined the number of beans in a large sample (1.0 kg) using the total mass of the sample. All you needed to know was the average mass of one bean. Let's apply this logic to another familiar example. **Figure 2** shows the masses of a group of tennis balls and a group of golf balls.

Figure 2 We can predict the number of balls in each sample if we know the average mass of a single tennis ball and a single golf ball.

The total mass of the tennis balls is 672.7 g. If we have been told that the mass of one tennis ball is 56.0 g, then we can determine the number of tennis balls in the sample by dividing the total mass of the tennis balls by the mass of an individual ball:

$$\frac{672.7\ \text{g}}{\dfrac{56.0\ \text{g}}{1\ \text{tennis ball}}} = (672.7\ \cancel{\text{g}})\left(\frac{1\ \text{tennis ball}}{56.0\ \cancel{\text{g}}}\right)$$

$$= 12.0\ \text{tennis balls}$$

Similarly, if the total mass of the golf balls is 540.0 g and the mass of an individual ball is 45.2 g, then we can calculate the number of balls in the sample:

$$\frac{540.0\ \text{g}}{\dfrac{45.2\ \text{g}}{1\ \text{golf ball}}} = (540.0\ \cancel{\text{g}})\left(\frac{1\ \text{golf ball}}{45.2\ \cancel{\text{g}}}\right)$$

$$= 11.9\ \text{golf balls}$$

This answer was reached only after rounding to the correct number of significant digits. A calculated result such as 11.9 golf balls is not a whole number because it is derived using a measurement: in this case, mass. Of course, finding the number of balls by counting them gives exactly 12. Every measurement has some degree of uncertainty. This uncertainty is due to a number of factors, including the limitations of the measuring device and slight variation from one ball to the next. In conclusion, the use of a measurement such as mass can only give an estimate of the number of entities in a given collection. It can never give an exact value. The only way to know for sure is to physically count them. In the case of a few identical balls on a balance, the estimate is very good. The estimate may not be as good for a much larger number of objects, such as the number of grains of rice in a 1.0 kg bag.

The use of measurements such as mass to indirectly count a large number of entities has some important applications. For example, industries commonly use mass to indirectly determine the number of small parts (such as nuts and bolts) in inventory. Also, automatic coin-counting machines use measurement to sort and "count" money (**Figure 3**). In some counting machines, coins are first sorted into stacks according to type. Then the height or mass of each stack is used to determine the number of coins present. These measurements are used to calculate the total value of the money inserted into the machine.

Figure 3 Coin-counting machines stack coins and use height or mass to determine how much to pay out.

Table 1 Counting Units

Unit	Example
1 dozen	12 golf balls
1 six-pack	6 cans of pop
1 ream	500 sheets of paper
1 h	60 min

mole a unit of amount; the amount of substance containing 6.02×10^{23} entities; unit symbol mol

Avogadro's constant (N_A) the number of entities in 1 mol of a substance

LEARNING **TIP**

Scientific Notation
You will use scientific notation frequently in this unit. For a quick refresher on how to use scientific notation correctly, refer to A6.1 in the Skills Handbook.

amount (n) the quantity of a substance, measured in moles

The Mole—A Chemist's "Dozen"

A dozen is a common counting unit. A counting unit is a convenient number that makes it easier to count objects. We often count doughnuts and eggs by dozens. Similarly, we count shoes and gloves in pairs. **Table 1** gives examples of common counting units. Chemists, however, do not work with macroscopic (visible) objects like shoes and doughnuts. Instead, chemists are interested in microscopic entities: atoms and molecules. Since these entities are so small, chemists established their own practical counting unit called the mole.

What Is a Mole?

The mole (symbol: mol) is the SI base unit for the amount of a substance. One **mole** of a substance contains 602 000 000 000 000 000 000 000 or 6.02×10^{23} entities of the substance. These entities could be anything from electrons and atoms to stars. However, for practical purposes, the mole is used to count microscopic entities: atoms, ions, molecules, and subatomic particles such as electrons. The value 6.02×10^{23} is sometimes called **Avogadro's constant** (N_A) in honour of the Italian physicist Amedeo Avogadro (1776–1856). Avogadro's constant is determined experimentally. As experimental methods improve, the value of the constant becomes more precise. Currently, $N_A = 6.022\ 141\ 99 \times 10^{23}$ entities. However, for convenience, we mostly use the value 6.02×10^{23}.

You may wonder why chemists did not choose a more convenient number for their counting unit. A trillion, for example, would have been easier to remember. There is a logical reason for their choice. Like other SI base units (such as the metre and the kilogram), the mole is defined against a known standard. The standard chosen to define the mole is the number of atoms in exactly 12 g of carbon—to be more precise, in the carbon-12 isotope. Scientists have experimentally determined that exactly 12 g of carbon-12 contains 6.02×10^{23} atoms of carbon.

Remember that a mole is a counting unit just like a dozen. However, instead of counting atoms by the dozen, chemists count them by the mole. For example, **Figure 4** shows equal amounts of carbon atoms and sulfur atoms: 1 mole of each. Both samples contain 6.02×10^{23} atoms. The sulfur sample has a larger volume and a greater mass because sulfur atoms are larger and heavier than carbon atoms. Similarly, a dozen tennis balls occupy a larger volume than a dozen golf balls.

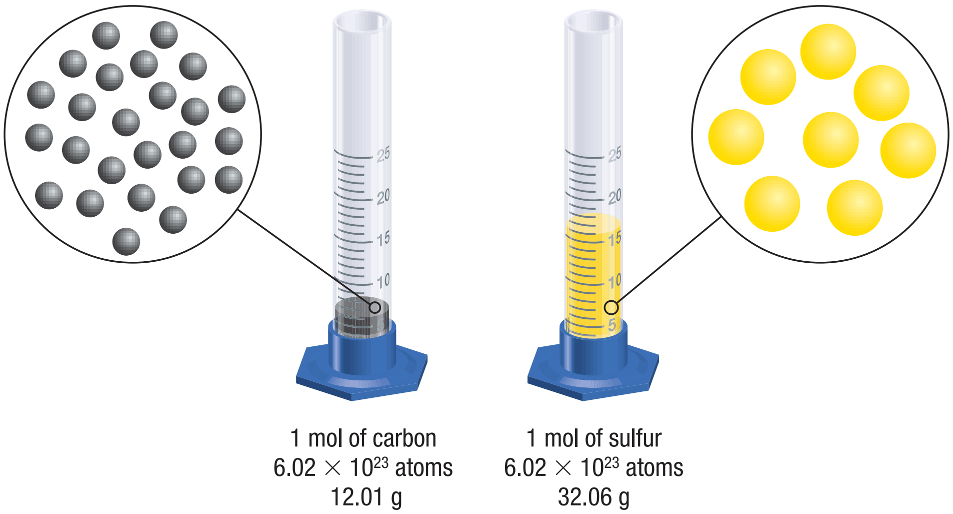


Figure 4 One mole of carbon and one mole of sulfur; both samples contain the same number of atoms. Since the atoms are different, the mass and volume of each sample are also different. The amounts of the two substances, however, are the same.

Note that, just as the volume of a substance is measured in litres, the **amount** (n) of a substance is measured in moles. Scientists use moles to communicate the amount of tiny entities such as subatomic particles, atoms, formula units, and molecules.

For example, **Figure 5** shows equal amounts of three common substances: a molecular compound, an ionic compound, and an element. Each sample contains 6.02×10^{23} entities, just as a package of a dozen doughnuts contains 12 doughnuts.

Figure 5 Samples of 1 mol each of sucrose, sodium chloride, and carbon. All three samples contain the same number of entities.

How Big Is a Mole?

It is difficult to comprehend the magnitude of huge numbers like Avogadro's constant. The following analogy may help.

THE GREEN PEA ANALOGY

The values in **Table 2** were determined using the following reasoning. One hundred green peas fill an average teacup of a known volume. If the volume of a refrigerator is known, the number of green peas that would fit in the refrigerator can be calculated without actually having to fill the refrigerator with a lot of peas and then count them all.

Table 2 Representing Quantities of Green Peas

Number of green peas	Object that holds this number of peas
100 or 10^2 (one hundred)	teacup
1 000 000 or 10^6 (one million)	refrigerator
10^9 (one billion)	an average three-bedroom home
10^{12} (one trillion)	all the homes in a small town
10^{15} (one quadrillion)	all the homes in a large city such as Hamilton
10^{18} (one quintillion)	one-half of Ontario covered 1 m deep in peas
10^{21} (one sextillion)	all the continents covered 1 m deep in peas
$10^{23}\left(\frac{1}{6}\text{ of 1 mol}\right)$	250 planets like Earth covered 1 m deep in peas

Tutorial 1 Working with Powers of 10

In the sections that follow, you will be performing calculations involving numbers expressed in scientific notation. Let's first review the multiplication and division of numbers in scientific notation. Recall that a number expressed in scientific notation is written in the form $a \times 10^x$, where the absolute value of a is 1 or greater but less than 10. For example, the number 1.6×10^{-2} is in scientific notation.

Sample Problem 1: Multiplying Numbers Expressed in Scientific Notation

Calculate the product of 2×10^5 and 7×10^{-8}.

To multiply numbers in scientific notation, multiply the coefficients and then add the exponents. Note that the final answer must also be in scientific notation.

$$\begin{aligned}(2 \times 10^5)(7 \times 10^{-8}) &= 14 \times 10^{5+(-8)} \\ &= 14 \times 10^{-3} \text{ [not in scientific notation]} \\ &= 1.4 \times 10^{-2} \text{ [in scientific notation]}\end{aligned}$$

Remember that the coefficient of a number written in scientific notation can only have a single digit to the left of the decimal point.

Sample Problem 2: Dividing Numbers Expressed in Scientific Notation

Divide 9×10^4 by 3×10^{-6}.

To divide numbers in scientific notation, divide the coefficients and subtract the exponents.

$$\frac{9 \times 10^4}{3 \times 10^{-6}} = 3 \times 10^{4-(-6)}$$
$$= 3 \times 10^{10}$$

Practice

SKILLS HANDBOOK A6.1

1. Perform the following calculations: K/U
   (a) $(2 \times 10^{-4})(3 \times 10^{-10})$ [ans: 6×10^{-14}]
   (b) $(5.0 \times 10^2)(2.4 \times 10^{-3})$ [ans: 1.2]
   (c) $\frac{1.95 \times 10^2}{1.3 \times 10^{-5}}$ [ans: 1.5×10^7]
   (d) $\frac{1.05 \times 10^{-23}}{2.5 \times 10^{-25}}$ [ans: 4.2×10^1]

6.3 Summary

- The mole is a counting unit, just like a dozen.
- The mole is the SI base unit for the amount of a substance.
- Avogadro's constant, N_A, is defined as the number of atoms in exactly 12 g of the carbon-12 isotope.
- One mole of a substance contains 6.02×10^{23} entities (atoms, ions, molecules, or formula units).

6.3 Questions

1. Why are familiar objects such as pens and paper clips not commonly counted in moles? K/U
2. Nails are usually sold by mass rather than by number, or count. Describe how you could determine the number of nails in a 500 g box. T/I
3. A chemist pours 1 mol of zinc granules into one beaker and 1 mol of zinc chloride powder into another beaker. T/I
   (a) What do the two samples have in common?
   (b) Which sample has the greater mass? Why?
4. What is the standard that Avogadro's constant is based on? K/U
5. (a) Calculate the number of doughnuts in 4 dozen doughnuts.
   (b) Use the same logic used in (a) to calculate the number of molecules in 4.0 mol of carbon dioxide.
   (c) Describe how to calculate the number of entities in a given amount (number of moles) of a substance. T/I
6. Calculate the following: T/I
   (a) $\frac{6.02 \times 10^{23}}{3.01 \times 10^{25}} =$
   (b) $\frac{4.6 \times 10^{22}}{2} =$
   (c) $\frac{6.02 \times 10^{-24}}{9.0 \times 10^{-21}} =$
7. A table of astronomical data gives the average distance from Earth to Pluto as 5.7 billion km. The average thickness of a sheet of notepaper is 1.2×10^{-4} m. Approximately how many sheets of paper would you need to make a pile reaching from Earth to Pluto? Give your answer in mol. T/I A
8. If you won a mole of dollars in a lottery, and were paid in \$100 bills, how long would it take you to count your winnings? You will have to estimate how fast you can count. T/I A
9. Suppose everybody in the world could count 100 objects in a minute, 24 hours per day, 7 days a week, without stopping. Use current data on the world's population to calculate approximately how long it would take for us all to count 1 mol of objects. T/I

GO TO NELSON SCIENCE

Molar Mass

Hydrogen is an important industrial chemical. The most common production method is by the reaction of a hydrocarbon with steam (**Figure 1**). The chemical equation for this reaction, using methane, is

$$CH_4(g) + H_2O(g) \rightarrow 3\ H_2(g) + CO(g)$$

This chemical equation tells you that 1 molecule of water reacts with every molecule of methane. Hydrogen manufacturers must have a convenient method of measuring exactly the required quantities of both reactants. Simply combining identical masses of the two substances will not work because the atomic masses of carbon, hydrogen, and oxygen are different. Therefore, 1 kg of methane contains a different number of molecules than 1 kg of water. To get around this problem, chemical manufacturers calculate the amount of each substance required (in moles). To do this, they must first determine the molar mass of each substance. The **molar mass** is the mass in grams of 1 mole or 6.02×10^{23} entities of that substance. The units of molar mass are g/mol. Chemists often use the symbol M to represent molar mass, with the chemical symbol or formula as a subscript. The symbol for the molar mass of water is therefore M_{H_2O}.

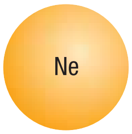
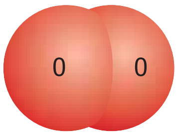
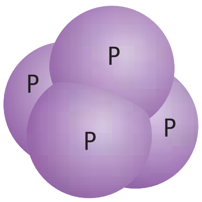
Figure 1 A researcher studies how to use steam to produce hydrogen gas from hydrocarbons. This process is called steam-reforming.

molar mass the mass of 1 mol of a substance; unit symbol g/mol

Molar Mass of Elements

The molar mass of a monatomic element such as neon is equal to the mass given on the periodic table. The molar mass of neon is therefore 20.18 g/mol. The molar mass of a molecular element such as oxygen is determined by multiplying the molar mass of the element by the number of atoms per molecule (**Table 1**).

Table 1 Calculating Molar Masses of Elements

Name and formula	Model	Molar mass calculation
neon, Ne		$M_{Ne} = 1 \times 20.18\ \frac{g}{mol} = 20.18\ \frac{g}{mol}$
oxygen, O_2		$M_{O_2} = 2 \times 16.00\ \frac{g}{mol} = 32.00\ \frac{g}{mol}$
phosphorus, P_4		$M_{P_4} = 4 \times 30.97\ \frac{g}{mol} = 123.9\ \frac{g}{mol}$

LEARNING **TIP**

Significant Digits and Counting Numbers

Some values are exact. For example, there are exactly 2 atoms of oxygen in a molecule of O_2. In calculations such as

$$2 \times 16.00\ \frac{g}{mol} = 32.00\ \frac{g}{mol}$$

you should always treat a "counting number" as if it has an infinite number of significant digits. The number of significant digits in the answer is determined by the measured quantity in the calculation. In this case, 16.00 g/mol has 4 significant digits so the answer should have 4 significant digits.

Numerical Equivalency of Atomic and Molar Masses

The mass of 1 mole of a monatomic element, expressed in g/mol, has the same numerical value as the average atomic mass of the element expressed in atomic mass units (u). How convenient! This means that, on the periodic table, only one mass value need be given for each element (**Figure 2**). This is true because of the value of Avogadro's constant. We can use carbon to illustrate this. According to the periodic

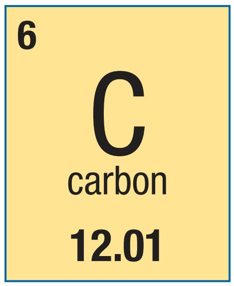


Figure 2 The atomic mass of a carbon atom, 12.01 u, has the same numerical value as the mass of 1 mole of carbon, 12.01 g.

table, 1 carbon atom has an average atomic mass of 12.01 u. Chemists define one atomic mass unit (1 u) as follows:

$$1\ \text{u} = 1.661 \times 10^{-24}\ \text{g}$$

The mass of one carbon atom in grams is therefore

$$(12.01\ \cancel{\text{u}})\left(\frac{1.661 \times 10^{-24}\ \text{g}}{1\ \cancel{\text{u}}}\right) = 1.99486 \times 10^{-23}\ \text{g}\ [\text{extra digits carried}]$$

Recall that Avogadro's constant is the number of entities in a mole of a substance. We can multiply the mass of one carbon atom by Avogadro's constant (expressed to 4 significant digits) to find the mass of 1 mol of carbon atoms:

$$\left(\frac{1.99486 \times 10^{-23}\ \text{g}}{1\ \cancel{\text{atom}_\text{c}}}\right)(6.022 \times 10^{23}\ \cancel{\text{atom}_\text{c}}) = 12.01\ \text{g}$$

The mass of 1 mol of carbon atoms is 12.01 g. The molar mass of carbon is therefore 12.01 g/mol.

As you can see, the molar mass of carbon and the mass of 1 atom of carbon have the same numerical value. Remember, however, that 12.01 g/mol and 12.01 u represent very different quantities of carbon. One mole or 12.01 g of carbon contains 6.022×10^{23} atoms—enough to fill one-quarter of a small test tube. However, 12.01 u is the mass of one carbon atom, which is too small to be visible even with the most powerful imaging technology.

LEARNING TIP

Significant Digits

If we want a fairly precise answer—say to 4 significant digits—we need to use values in the calculation with at least this many digits. Therefore, in this calculation we are using Avogadro's constant with 3 decimal places (6.022×10^{23}), instead of the more usual 2 decimal places.

Molar Masses of Compounds

We can also describe quantities of molecular and ionic compounds using moles. The molar mass of a compound is the sum of the molar mass of each entity in the compound. For example, carbon dioxide is a common molecular compound (**Figure 3(a)**). Solid carbon dioxide (dry ice) consists of vast numbers of CO_2 molecules. Each molecule contains 1 carbon atom covalently bonded to 2 oxygen atoms. Sodium chloride is a common ionic compound (**Figure 3(b)**). A crystal of sodium chloride does not contain molecules. Instead, it consists of alternating positive sodium ions and negative chloride ions. The chemical formula, NaCl, indicates that there is 1 sodium ion for every chloride ion. In Unit 1 you learned that the simplest ratio of ions in an ionic compound is called a formula unit. NaCl is the formula unit for sodium chloride.

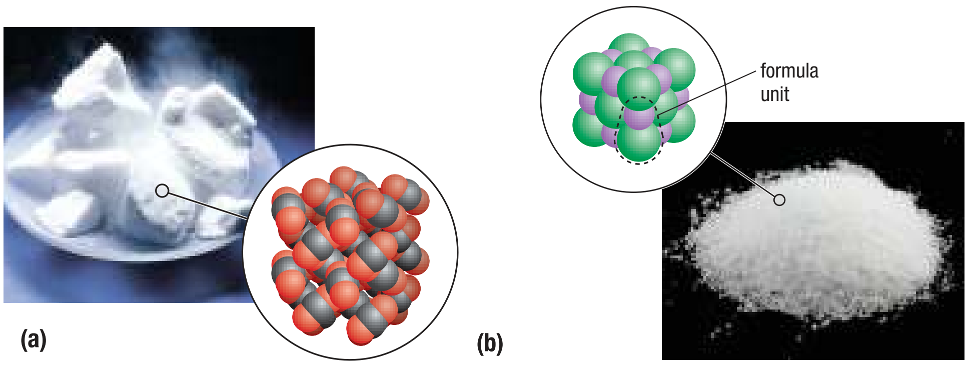


Figure 3 (a) Molecular solids are made up of molecules, while (b) ionic solids consist of a continuous pattern of alternating positive and negative ions. A formula unit for an ionic compound is the simplest ratio of positive and negative ions in the compound.

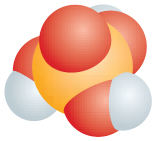
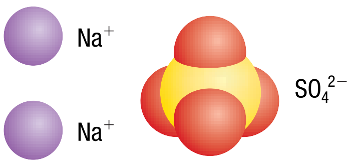
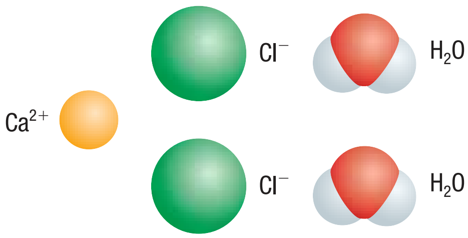
Table 2 shows molar mass calculations for common molecular and ionic compounds. Remember that you have to consider every element in the molecular formula, or formula unit, as you calculate the molar mass.

LEARNING TIP

Formula Unit and Molecular Formula

The formulas for HCl and NaCl look similar, but the entities they represent are quite different. Hydrogen chloride consists of molecules. The atoms within each molecule are joined by covalent bonds. However, there are no separate units of NaCl in a salt crystal. Rather, an ionic compound is a continuous network of positive and negative ions. The formula unit, NaCl, indicates the simplest ratio of these ions in the compound.

Table 2 Common Molecular and Ionic Compounds

Name and formula	Model	Molar mass calculation
hydrogen phosphate, H_3PO_4		$\begin{aligned} M_{H_3PO_4} &= 3M_H + M_P + 4M_O \\ &= \left(3 \times 1.01\ \frac{g}{mol}\right) + \left(30.97\ \frac{g}{mol}\right) + \left(4 \times 16.00\ \frac{g}{mol}\right) \\ &= 3.03\ \frac{g}{mol} + 30.97\ \frac{g}{mol} + 64.00\ \frac{g}{mol} \\ M_{H_3PO_4} &= 98.00\ \frac{g}{mol} \end{aligned}$
sodium chloride, NaCl		$\begin{aligned} M_{NaCl} &= M_{Na^+} + M_{Cl^-} \\ &= 22.99\ \frac{g}{mol} + 35.45\ \frac{g}{mol} \\ M_{NaCl} &= 58.44\ \frac{g}{mol} \end{aligned}$
sodium sulfate, Na_2SO_4		$\begin{aligned} M_{Na_2SO_4} &= 2M_{Na^+} + M_{SO_4^{\ 2-}} \\ &= 2M_{Na^+} + M_S + 4M_O \\ &= \left(2 \times 22.99\ \frac{g}{mol}\right) + \left(32.07\ \frac{g}{mol}\right) + \left(4 \times 16.00\ \frac{g}{mol}\right) \\ &= 45.98\ \frac{g}{mol} + 32.07\ \frac{g}{mol} + 64.00\ \frac{g}{mol} \\ M_{Na_2SO_4} &= 142.05\ \frac{g}{mol} \end{aligned}$
calcium chloride dihydrate, $CaCl_2{\bullet}2H_2O$		$\begin{aligned} M_{CaCl_2 \bullet 2H_2O} &= M_{Ca^{2+}} + 2M_{Cl^-} + 2M_{H_2O} \\ &= M_{Ca^{2+}} + 2M_{Cl^-} + 2(2M_H + M_O) \\ &= \left(40.08\ \frac{g}{mol}\right) + \left(2 \times 35.45\ \frac{g}{mol}\right) + \\ &\qquad 2\left(2 \times 1.01\ \frac{g}{mol} + 16.00\ \frac{g}{mol}\right) \\ M_{CaCl_2 \bullet 2H_2O} &= 147.02\ \frac{g}{mol} \end{aligned}$

Tutorial 1 Determining Molar Mass

The molar mass of a monatomic element is provided on the periodic table. Otherwise, determine molar mass by summing the masses of each atom or ion in the chemical formula.

Sample Problem 1: Calculating the Molar Mass of a Molecular Compound

Calculate the molar mass of carbon dioxide.

Given: carbon dioxide, CO_2

Required: molar mass of CO_2, M_{CO_2}

Solution:

Step 1. Look up the molar masses of the elements:

$$M_C = 12.01\ \frac{g}{mol};\ M_O = 16.00\ \frac{g}{mol}$$

Step 2. Add the molar masses of the elements, multiplying the molar mass of each element by the number of atoms of that element in the compound. In this case, there is 1 atom of C and 2 atoms of O.

LEARNING TIP

Periodic Table

You will refer to the periodic table frequently in this Chemistry course. The periodic table is provided on the inside back cover of this textbook and in Appendix B1.

$$M_{CO_2} = M_C + 2M_O$$
$$= \left(12.01\,\frac{g}{mol}\right) + \left(2 \times 16.00\,\frac{g}{mol}\right)$$
$$= 12.01\,\frac{g}{mol} + 32.00\,\frac{g}{mol}$$
$$M_{CO_2} = 44.01\,\frac{g}{mol}$$

Statement: The molar mass of carbon dioxide is 44.01 g/mol.

Sample Problem 2: Calculating the Molar Mass of an Ionic Compound

Calculate the molar mass of iron(III) oxide.

Given: iron(III) oxide, Fe_2O_3

Required: molar mass of Fe_2O_3, $M_{Fe_2O_3}$

Solution:

Step 1. As in Sample Problem 1, look up the molar masses of the elements.

$$M_{Fe} = 55.85\,\frac{g}{mol};\quad M_O = 16.00\,\frac{g}{mol}$$

Step 2. Add the molar masses of the elements, multiplying the molar mass of each element by the number of atoms of that element in the compound. Note that iron(III) oxide is an ionic compound containing two Fe^{3+} ions and three O^{2-} ions per formula unit. Since the mass of an electron is extremely small, we may assume that the masses of Fe^{3+} ions and O^{2-} ions are the same as the masses of Fe atoms and O atoms, respectively.

$$M_{Fe_2O_3} = 2M_{Fe^{3+}} + 3M_{O^{2-}}$$
$$= 2\left(55.85\,\frac{g}{mol}\right) + 3\left(16.00\,\frac{g}{mol}\right)$$
$$= 111.7\,\frac{g}{mol} + 48.00\,\frac{g}{mol}$$
$$M_{Fe_2O_3} = 159.7\,\frac{g}{mol}$$

Statement: The molar mass of iron(III) oxide is 159.7 g/mol.

Sample Problem 3: Calculating the Molar Mass of a Hydrate

Calculate the molar mass of iron(III) chloride hexahydrate.

Given: iron(III) chloride hexahydrate, $FeCl_3{\cdot}6H_2O$

Required: molar mass of $FeCl_3{\cdot}6H_2O$, $M_{FeCl_3\cdot 6H_2O}$

Solution:

Step 1. Look up the molar masses of the elements in the compound in the periodic table.

$$M_{Fe^{3+}} = 55.85\,\frac{g}{mol};\ M_{Cl^-} = 35.45\,\frac{g}{mol};\ M_H = 1.01\,\frac{g}{mol};\ M_O = 16.00\,\frac{g}{mol}$$

Step 2. Add the molar masses of the elements, multiplying the molar mass of each element by the number of atoms of that element in the compound.

$$M_{FeCl_3\cdot 6H_2O} = M_{Fe^{3+}} + 3M_{Cl^-} + 6(2M_H + M_O)$$
$$= \left(55.85\,\frac{g}{mol}\right) + 3\left(35.45\,\frac{g}{mol}\right) + 6\left(2 \times 1.01\,\frac{g}{mol} + 16.00\,\frac{g}{mol}\right)$$
$$M_{FeCl_3\cdot 6H_2O} = 270.32\,\frac{g}{mol}$$

Statement: The molar mass of iron(III) chloride hexahydrate is 270.32 g/mol.

Practice

1. Calculate the molar mass of each of the following substances. Express your answer to 2 decimal places. K/U T/I
   (a) S_8 (a component of gunpowder) [ans: 256.56 g/mol]
   (b) H_2S (the smell of rotten eggs) [ans: 34.09 g/mol]
   (c) NaOH (used in paper manufacture) [ans: 40.00 g/mol]
   (d) $Fe(OH)_3$ (a pigment in paint and cosmetics) [ans: 106.88 g/mol]
   (e) $(NH_4)_2S$ (used in textile production) [ans: 68.17 g/mol]
   (f) $Ca_3(PO_4)_2$ (used to make fertilizers) [ans: 310.18 g/mol]
   (g) $MgSO_4{\cdot}7H_2O$ (Epsom salts) [ans: 246.52 g/mol]

Mini Investigation

The Mole Exhibit

Skills: Performing, Observing, Analyzing, Communicating

SKILLS HANDBOOK A1, A3.2

In this activity, your teacher will ask you to measure 1.00 mol of a pure substance and display it in the classroom.

Equipment and Materials: chemical safety goggles; lab apron; balance; scoopula; 250 mL beaker; samples of substances assigned by your teacher

1. Calculate the molar mass of your assigned substance.
2. Once your teacher approves your calculations, put on your chemical safety goggles and lab apron.
3. Use the balance to measure out 1.00 mol of your substance into the beaker.

A. What is the molar mass of your substance? T/I

B. How many entities of your substance are in the beaker? K/U

Calculations Involving Molar Masses

Chemists routinely use mass to measure the quantity of a chemical needed for an investigation. As such, chemists need a convenient way to determine the amount in moles, n, in a given mass in grams, m, of a sample. As you will see, molar mass, M, is the quantity that mathematically connects the amount of a substance to its mass.

UNIT TASK BOOKMARK

You will use the concept of molar mass as you perform the calculations necessary for the Unit Task on page 352.

Converting Mass to Amount

We can use logic to develop a formula that relates amount, mass, and molar mass. Suppose a chemist needs 4.00 g of sodium hydroxide for an investigation. Earlier, you determined that the mass of one mole of sodium hydroxide is 40.00 g/mol. Therefore, 4.00 g of sodium hydroxide represents one tenth or a mole, or 0.100 mol. How did we arrive at this answer? We divided the mass (4.00 g) by the molar mass (40.00 g/mol):

$$n_{\text{NaOH}} = \frac{4.00\ \text{g}}{40.00\ \dfrac{\text{g}}{\text{mol}}}$$

$$= 4.00\ \cancel{\text{g}} \times \left(\frac{1\ \text{mol}}{40.00\ \cancel{\text{g}}}\right)$$

$$n_{\text{NaOH}} = 0.100\ \text{mol}$$

In general, the mathematical relationship between m, n and M, therefore, is

$$n = \frac{m}{M}$$

Converting Amount to Mass

If sodium hydroxide has a molar mass of 40.00 g, we can predict that 1.50 mol of sodium hydroxide has a mass of 60.0 g. We can get this answer by rearranging the equation $n = \dfrac{m}{M}$ into the form $m = nM$.

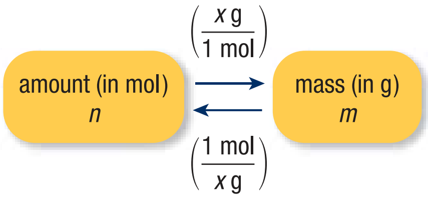


Figure 4 Molar mass is the mathematical link between the amount in moles and the mass in grams.

Figure 5 This glass of cola contains 40.0 g of sugar (sucrose).

$$m_{\mathrm{NaOH}} = 1.50\ \cancel{\mathrm{mol}} \times \left(\frac{40.00\ \mathrm{g}}{1\ \cancel{\mathrm{mol}}}\right)$$

$$m_{\mathrm{NaOH}} = 60.0\ \mathrm{g}$$

As you can see from these examples, the mass of a chemical can be converted to an amount, and vice versa, by multiplying by a conversion factor. This factor is either the molar mass (units g/mol) or the reciprocal molar mass (units mol/g) (**Figure 4**).

Tutorial 2 Converting among Amount, Mass, and Molar Mass

The Sample Problems in this tutorial illustrate how to calculate the mass, *m*, of a sample of a substance given the amount of that substance (in mol), and vice versa.

Sample Problem 1: Calculating Amount from Mass

According to the manufacturer, a typical can of cola contains 40.0 g of sucrose, $C_{12}H_{22}O_{11}$ (**Figure 5**). Calculate the amount of sucrose in 40.0 g.

Given: $m_{C_{12}H_{22}O_{11}} = 40.0\ \mathrm{g}$

Required: amount of sucrose, $n_{C_{12}H_{22}O_{11}}$

Solution:

Step 1. Calculate the molar mass of sucrose, $M_{C_{12}H_{22}O_{11}}$.

$$M_{C_{12}H_{22}O_{11}} = 12\left(12.01\ \frac{\mathrm{g}}{\mathrm{mol}}\right) + 22\left(1.01\ \frac{\mathrm{g}}{\mathrm{mol}}\right) + 11\left(16.00\ \frac{\mathrm{g}}{\mathrm{mol}}\right)$$

$$M_{C_{12}H_{22}O_{11}} = 342.34\ \frac{\mathrm{g}}{\mathrm{mol}}$$

Step 2. Use the mass of sucrose (given in the question) and the molar mass of sucrose to calculate the amount of sucrose. To do this, you need to develop a conversion factor using the molar mass. The conversion factor is either

$$\frac{1\ \mathrm{mol}}{342.34\ \mathrm{g}} \quad \text{or} \quad \frac{342.34\ \mathrm{g}}{1\ \mathrm{mol}}$$

Since you want to solve for an amount (in mol) of sucrose, multiply by the factor that has mol in the numerator, as follows:

$$n_{C_{12}H_{22}O_{11}} = 40.0\ \cancel{\mathrm{g}} \times \frac{1\ \mathrm{mol}}{342.34\ \cancel{\mathrm{g}}}$$

$$n_{C_{12}H_{22}O_{11}} = 0.117\ \mathrm{mol}$$

Statement: There is 0.117 mol of sucrose in a 40.0 sample of sucrose.

Sample Problem 2: Calculating Mass from Amount

A litre of human blood typically contains 4.0 mmol of glucose, $C_6H_{12}O_6$. Calculate the mass of this amount of glucose.

Given: $m_{C_6H_{12}O_6} = 40.0\ \mathrm{mmol}$

Required: mass of glucose, $M_{C_6H_{12}O_6}$

Solution: $4.0\ \mathrm{mmol} = 4.0 \times 10^{-3}\ \mathrm{mol}$

$$M_{C_6H_{12}O_6} = 6\left(12.01\ \frac{\mathrm{g}}{\mathrm{mol}}\right) + 12\left(1.01\ \frac{\mathrm{g}}{\mathrm{mol}}\right) + 6\left(16.00\ \frac{\mathrm{g}}{\mathrm{mol}}\right)$$

$$M_{C_6H_{12}O_6} = 180.18\ \frac{\mathrm{g}}{\mathrm{mol}}$$

$$m_{C_6H_{12}O_6} = (4.0 \times 10^{-3}\ \cancel{\mathrm{mol}})\left(\frac{180.18\ \mathrm{g}}{1\ \cancel{\mathrm{mol}}}\right)$$

$$m_{C_6H_{12}O_6} = 0.72\ \mathrm{g}$$

Statement: The mass of 4.0 mmol of glucose in a litre of blood is 0.72 g of glucose.

Practice

SKILLS HANDBOOK A6.3

1. Calculate the amount of pure substance in each of the following samples: T/I
   (a) a 500 g box of table salt (assume pure sodium chloride) [ans: 8.56 mol]
   (b) 14.2 g of aluminum in a typical pop can [ans: 0.526 mol]
   (c) 1.00 kg of calcium oxide, CaO, used to neutralize a soil sample [ans: 17.8 mol]
   (d) 1.75 kg of hydrogen chloride, HCl, in a jug of concentrated hydrochloric acid [ans: 48.0 mol]
   (e) 200.0 mg of ibuprofen, $C_{13}H_{18}O_2$, in a headache medication [ans: 9.694×10^{-4} mol]
2. Calculate the mass of the following amounts of a pure substance: T/I
   (a) 0.80 mol of hydrogen peroxide, H_2O_2, in a bottle of antiseptic [ans: 27 g]
   (b) 3.25 mol of sodium hydrogen sulfate, $NaHSO_4$, in a container of bathroom cleaner [ans: 3.90×10^2 g]
   (c) 5.0 mmol of calcium carbonate, $CaCO_3$, in an antacid tablet [ans: 0.50 g]
   (d) 1.2×10^3 mol of sodium hypochlorite, NaOCl [ans: 8.9×10^4 g]
   (e) 45 mmol of helium in a balloon [ans: 1.8×10^{-1} g]

6.4 Summary

- The molar mass and average atomic mass of an element are numerically the same. This value is given on the periodic table.
- The unit of molar mass is g/mol, while the unit of average atomic mass is u.
- The molar mass of a compound is the sum of the molar masses of each entity in the compound.
- The amount, n, mass, m, and molar mass, M, of a pure substance are related to each other through the equation $n = \frac{m}{M}$.

6.4 Questions

1. Do 1.0 mol samples of different compounds all have the same mass? Explain your answer. K/U
2. Describe how chemists use mass as a way of getting a precise estimate of the number of entities in a sample. K/U
3. Why is an amount, obtained using the equation $n = \frac{m}{M}$, always an estimate rather than an exact value? K/U
4. Calculate the molar mass of each of the following compounds: T/I
   (a) iron(III) oxide, Fe_2O_3 (in rust)
   (b) calcium carbonate, $CaCO_3$ (in blackboard chalk)
   (c) octane, C_8H_{18} (in gasoline)
   (d) calcium chlorate, $Ca(ClO_3)_2$ (in fireworks)
   (e) ammonium carbonate, $(NH_4)_2CO_3$ (in smelling salts)
5. Calcium chloride, $CaCl_2$, is used as a drying agent to protect electronics during shipping. Calculate the amount of calcium chloride in a 10.0 g sample. T/I
6. A typical energy drink contains 80.0 mg of the stimulant caffeine, $C_8H_{10}N_4O_2$. Calculate the amount of caffeine in the energy drink. T/I
7. A kernel of popping corn contains 1.22×10^{-3} mol of water. What is the mass of this water? T/I
8. Calculate the amount of silicon in a 5.8 mg sample of pure silicon used to manufacture a computer chip. T/I
9. A tank truck carries 34 000 L of sulfuric acid. The density of sulfuric acid is 1.84 kg/L. T/I
   (a) What mass of sulfuric acid is in the truck?
   (b) What amount of sulfuric acid is in the truck?
10. A bag of intravenous fluid contains 0.154 mol of sodium chloride, NaCl. What mass of sodium chloride is required to prepare this bag? T/I
11. A 2.9 g sample of a pure compound contains 0.050 mol of the compound. Use this data to calculate the molar mass of the compound. T/I
12. A sodium atom has a mass of 22.99 u. Use Avogadro's constant and the conversion factor 1.661×10^{-24} g/u to calculate the mass of 1.00 mol of sodium atoms. T/I

6.5 Mass and Number of Entities

How many molecules of water are there in a drop of water? Even the smallest visible drop of water contains more molecules than you can count in a lifetime. Because molecules are so small, chemists use mass as a convenient way to estimate the number of molecules in a sample. For example, chemists know that water has a molar mass of 18.0 g/mol. Therefore, they know that 18.0 g of water contains 1.00 mol or 6.02×10^{23} molecules of water (**Figure 1**). Similarly 36.0 g of water contains 2.00 mol of water molecules. In this section, you will see how Avogadro's constant, N_A, serves as the link, or bridge, between the mass of a pure substance and the number of entities present.

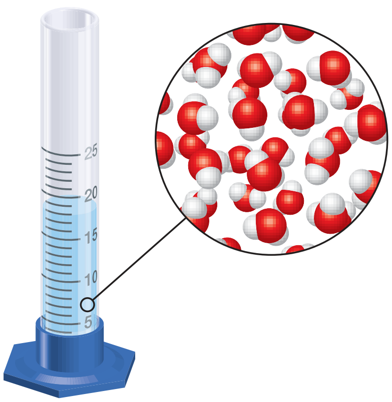


Figure 1 A mole of water has a mass of 18.0 g. Since water has a density of 1.0 g/mL at room temperature, 18.0 mL (1.00 mol) of water contains 6.02×10^{23} water molecules.

Determining the Number of Entities

We often use counting units to specify a number of entities. A dozen is a counting unit. It means "12." We can represent a larger number using dozens as the unit. For example, 5 dozen doughnuts contain 60 doughnuts. Mathematically, this can be determined using the following calculation:

$$\text{number of doughnuts} = (5\ \cancel{\text{dozen}})\left(\frac{12\ \text{doughnuts}}{1\ \cancel{\text{dozen}}}\right)$$
$$= 60\ \text{doughnuts}$$

In general:

$$\text{number of doughnuts} = (\text{number of}\ \cancel{\text{dozen}})\left(\frac{12\ \text{doughnuts}}{1\ \cancel{\text{dozen}}}\right)$$

Similarly, the number of molecules of water in 5.00 mol of water is:

$$\text{number of water molecules} = (5.00\ \cancel{\text{mol}})\left(\frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{\text{mol}}}\right)$$
$$= 30.1 \times 10^{23}\ \text{molecules}$$
$$= 3.01 \times 10^{24}\ \text{molecules}$$

The number o.f entities in a sample, N, is determined by multiplying the amount, n, by Avogadro's constant, N_A:

$N = nN_A$ (**Figure 2**)

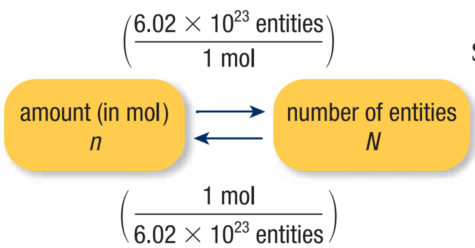


Figure 2 Avogadro's constant is the conversion factor linking an amount to the number of entities in a sample.

Tutorial 1 Calculating the Number of Entities in a Sample

A pure substance can consist of atoms, molecules, or ions. To find the number of these entities in a sample of a substance, we first need to know the amount of the substance. Then, we can find the number of entities by multiplying by a conversion factor.

Sample Problem 1: Calculating the Number of Atoms in a Sample

Calculate the number of atoms in a 1.00 kg bar of gold.

Given: $m_{Au} = 1.00$ kg

Required: number of atoms of gold, N_{Au}

Solution:

Step 1. Look up the molar mass of the element gold in the periodic table.

$$M_{Au} = 196.97\ \frac{\text{g}}{\text{mol}}$$

Step 2. Determine the amount of gold by multiplying the mass of gold (given) by an appropriate conversion factor derived from the molar mass of gold.

$$n_{Au} = (1.00 \times 10^3\ \cancel{\text{g}})\left(\frac{1\ \text{mol}}{196.97\ \cancel{\text{g}}}\right)$$
$$n_{Au} = 5.0760\ \text{mol}\ [2\ \text{extra digits carried}]$$

Step 3. Calculate the number of atoms of gold in 5.0760 mol of gold. To do this, note that 1 mol of gold contains 6×10^{23} atoms of gold, which gives two possible conversion factors:

$$\frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mol}} \text{ or } \frac{1 \text{ mol}}{6.02 \times 10^{23} \text{ atoms}}$$

Also note that the following equation may be used to calculate the number of atoms in an entity, given the amount of entity (in moles):

Since you want to solve for the number of atoms, multiply by the factor that has "number of atoms" in the numerator.

$$N_{Au} = (5.0760 \cancel{\text{mol}})\left(6.02 \times 10^{23} \frac{\text{atoms}}{\cancel{\text{mol}}}\right)$$

$$= 30.56 \times 10^{23} \text{ atoms}$$

$$= 3.056 \times 10^{24} \text{ atoms}$$

$$N_{Au} = 3.06 \times 10^{24} \text{ atoms}$$

Statement: A 1.00 kg bar of gold contains 3.06×10^{24} atoms of gold.

COMBINING STEPS IN THE CALCULATION

SKILLS HANDBOOK A6.3

Notice in Sample Problem 1 that, after determining the molar mass of gold, we calculated the number of gold atoms in two successive steps: Step 2 and Step 3. By combining these two steps we can calculate the number of atoms of gold from the given mass in one step as follows:

$$N_{Au} = (1.00 \times 10^{-3} \cancel{\text{g}})\left(\frac{1 \text{ mol}}{196.97 \cancel{\text{g}}}\right)\left(\frac{6.02 \times 10^{23} \text{ atoms}}{\cancel{1 \text{ mol}}}\right)$$

$$N_{Au} = 3.06 \times 10^{24} \text{ atoms}$$

Sample Problem 2: Calculating the Number of Molecules in a Sample

Ammonia is a pungent gas released from smelling salts. Calculate the number of molecules in a 4.00 mg sample of ammonia, NH_3.

Given: $m_{NH_3} = 4.00$ mg

$m_{NH_3} = 4.00 \times 10^{-3}$ g

Required: number of molecules of ammonia, N_{NH_3}

Solution:

Step 1. Look up the molar mass of nitrogen and hydrogen in the periodic table and calculate the molar mass of ammonia.

$$M_{NH_3} = M_N + 3M_H$$

$$= \left(14.01 \frac{\text{g}}{\text{mol}}\right) + 3\left(1.01 \frac{\text{g}}{\text{mol}}\right)$$

$$M_{NH_3} = 17.04 \frac{\text{g}}{\text{mol}}$$

Step 2. Calculate the amount of ammonia by multiplying the mass of ammonia (given) by an appropriate conversion factor derived from the molar mass of ammonia.

$$n_{NH_3} = (4.00 \times 10^{-3} \cancel{\text{g}})\left(\frac{1 \text{ mol}}{17.04 \cancel{\text{g}}}\right)$$

$$n_{NH_3} = 2.3474 \times 10^{-4} \text{ mol } [2 \text{ extra digits carried}]$$

Step 3. Calculate the number of molecules of ammonia in 2.3474×10^{-4} mol of ammonia using an appropriate conversion factor. In this case, use the conversion factor with amount in the denominator.

$$N_{NH_3} = (2.3474 \times 10^{-4}\ \cancel{mol})\left(6.02 \times 10^{23}\ \frac{\text{molecules}}{\cancel{mol}}\right)$$

$$= 14.1 \times 10^{19}\ \text{molecules}$$

$$N_{NH_3} = 1.41 \times 10^{20}\ \text{molecules}$$

Statement: A sample containing 4.00 mg of ammonia contains 1.41×10^{20} ammonia molecules.

We could perform the calculation in Sample Problem 2 in one step as follows:

$$N_{NH_3} = (4.00 \times 10^{-3}\ \cancel{g})\left(\frac{1\ \cancel{mol}}{17.04\ \cancel{g}}\right)\left(6.02 \times 10^{23}\ \frac{\text{molecules}}{1\ \cancel{mol}}\right)$$

$$= 14.1 \times 10^{19}\ \text{molecules}$$

$$N_{NH_3} = 1.41 \times 10^{20}\ \text{molecules}$$

Practice

SKILLS HANDBOOK A6.5

1. A typical car battery contains 8.16 kg of lead. Determine the number of lead atoms in this battery. T/I [ans: 2.37×10^{25} atoms]
2. A one-carat diamond has a mass of 2.00×10^{-1} g. Assuming the diamond is made of pure carbon, determine the number of carbon atoms in a one-carat diamond. T/I [ans: 1.00×10^{22} atoms]
3. Determine the number of ethylene molecules, C_2H_4, in a 7.3 g sample of ethylene. T/I [ans: 1.6×10^{23} molecules]
4. Find the number of formula units of $CaCl_2$ in a 1.5 kg pail of calcium chloride (used to melt sidewalk ice). T/I [ans: 8.1×10^{24} formula units]

Determining the Number of Entities in Compounds

Each molecular compound has a unique chemical formula, which gives the exact number of atoms of each element in a molecule. Therefore, if the number of molecules in a sample is known, chemists can calculate the number of atoms of each element present. For example, the chemical formula H_2O tells us that 1 water molecule contains 1 oxygen atom and 2 hydrogen atoms (**Figure 3**). Therefore, *x* water molecules contain *x* oxygen atoms and 2*x* hydrogen atoms.

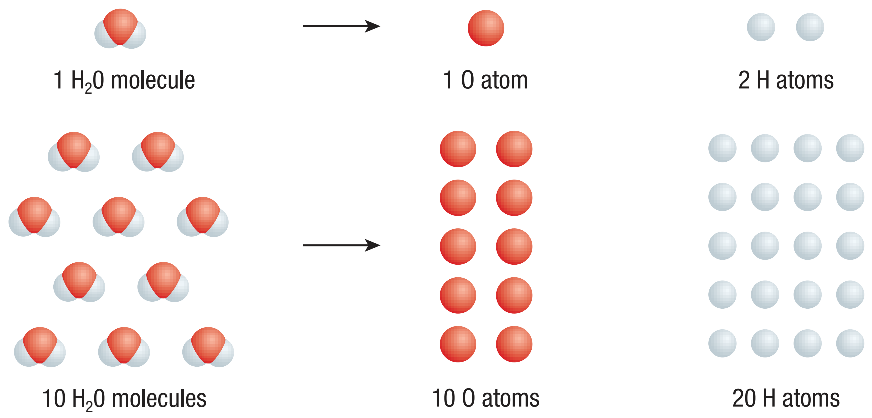


Figure 3 The number of atoms of an element in a sample of molecules is determined by multiplying the number of molecules by the number of atoms of the element per molecule.

We describe the number of ions in an ionic compound in a similar manner. The difference is that we specify the ratios of ions in a formula unit of the compound. For example, the chemical formula $Al_2(SO_4)_3$ states that there are 2 aluminum ions for every 3 sulfate ions in a formula unit of aluminum sulfate. (Remember that each sulfate ion is a polyatomic ion. It is considered a single entity within the formula unit.) Regardless of the number of aluminum sulfate formula units, aluminum ions and sulfate ions always occur in a 2:3 ratio (**Table 1**).

Table 1 Ratio of Ions in Aluminum Sulfate, $Al_2(SO_4)_3$

Number of formula units	Number of Al^{3+} cations	Number of SO_4^{2-} anions
1	$1 \times 2 = 2$	$1 \times 3 = 3$
10	$10 \times 2 = 20$	$10 \times 3 = 30$
2 000 000	$2\,000\,000 \times 2 = 4\,000\,000$	$2\,000\,000 \times 3 = 6\,000\,000$
$1 \text{ mol} = 6.02 \times 10^{23}$	$2 \text{ mol}_{Al^{3+}} = \dfrac{2 \text{ aluminum ions}}{1 \cancel{\text{formula unit}}}(6.02 \times 10^{23} \cancel{\text{formula unit}})$ $= 1.204 \times 10^{24}$ aluminum ions	$3 \text{ mol}_{SO_4^{2-}} = \dfrac{3 \text{ sulfate ions}}{1 \cancel{\text{formula unit}}}(6.02 \times 10^{23} \cancel{\text{formula unit}})$ $= 1.806 \times 10^{24}$ sulfate ions

Tutorial 2 Calculating the Number of Atoms or Ions in a Sample

Calculating the number of atoms or ions in a compound follows the same strategy as outlined in Tutorial 1. The only difference is that you now must consider the number of atoms or ions per molecule or formula unit of the compound.

Sample Problem 1: Calculating the Number of Atoms

Benzaldehyde, C_6H_5CHO, is a compound used to give prepared foods an almond flavour. Find the number of carbon, hydrogen, and oxygen atoms in a 26.5 g sample of benzaldehyde.

Given: $m_{C_6H_5CHO} = 26.5 \text{ g}$

Required: number of atoms of carbon, hydrogen and oxygen, N_C, N_H, and N_O

Solution:

Step 1. Look up the molar mass of carbon, hydrogen, and oxygen in the periodic table and calculate the molar mass of benzaldehyde.

$$M_{C_6H_5CHO} = 7M_C + 6M_H + M_O$$
$$= \left(7 \times 12.01 \frac{\text{g}}{\text{mol}}\right) + \left(6 \times 1.01 \frac{\text{g}}{\text{mol}}\right) + \left(16.00 \frac{\text{g}}{\text{mol}}\right)$$
$$M_{C_6H_5CHO} = 106.13 \frac{\text{g}}{\text{mol}}$$

Step 2. Calculate the amount of benzaldehyde by multiplying the mass of the sample (given) by an appropriate conversion factor derived from the molar mass of benzaldehyde.

$$n_{C_6H_5CHO} = 26.5 \cancel{\text{g}} \times \frac{1 \text{ mol}}{106.13 \cancel{\text{g}}}$$
$$n_{C_6H_5CHO} = 0.24969 \text{ mol [extra digits carried]}$$

Step 3. Calculate the number of molecules of benzaldehyde in the 26.5 g sample using an appropriate conversion factor. In this case, use the conversion factor with amount in the denominator.

$$N_{C_6H_5CHO} = (0.24969 \cancel{\text{mol}})\left(6.02 \times 10^{23} \frac{\text{molecules}}{1 \cancel{\text{mol}}}\right)$$
$$N_{C_6H_5CHO} = 1.503 \times 10^{23} \text{ molecules [extra digits carried]}$$

LEARNING TIP

When to Round Off Answers
Whenever the answer to one calculation will be used in a later calculation, rounding off early introduces unnecessary errors in the answer. Always round off after the last calculation.

Step 4. Determine the number of each type of atom present by multiplying the number of molecules by the number of each type of atom per molecule.

$$N_C = (1.503 \times 10^{23} \cancel{\text{molecules of benzaldehyde}})\left(\frac{7 \text{ atoms of carbon}}{1 \cancel{\text{ molecule of benzaldehyde}}}\right)$$

$$= 10.522 \times 10^{23} \text{ atoms}$$

$$N_C = 1.05 \times 10^{24} \text{ atoms}$$

$$N_H = (1.503 \times 10^{23} \cancel{\text{molecule of benzaldehyde}})\left(\frac{6 \text{ atoms of hydrogen}}{1 \cancel{\text{ molecule of benzaldehyde}}}\right)$$

$$N_H = 9.02 \times 10^{23} \text{ atoms}$$

$$N_O = (1.503 \times 10^{23} \cancel{\text{molecules of benzaldehyde}})\left(\frac{1 \text{ atoms of oxygen}}{1 \cancel{\text{ molecule of benzaldehyde}}}\right)$$

$$N_O = 1.50 \times 10^{23} \text{ atoms}$$

Statement: A 26.5 g sample of benzaldehyde contains 1.05×10^{24} atoms of carbon, 9.02×10^{23} atoms of hydrogen, and 1.50×10^{23} atoms of oxygen.

Practice

1. Find the number of hydrogen and oxygen atoms present in 15.0 g of pure hydrogen peroxide, H_2O_2. T/I [ans: 5.31×10^{23} hydrogen atoms and 5.31×10^{23} oxygen atoms]
2. Find the number of calcium ions in 2.5 g of calcium phosphate, $Ca_3(PO_4)_2$ (a nutrient in milk). T/I [ans: 1.5×10^{22} Ca^{2+} ions]
3. Calculate the mass of 8.4×10^{24} molecules of carbon dioxide (dry ice) (Hint: first calculate the number of molecules.) T/I [ans: 6.1×10^{2} g]

Mini Investigation

Determining Numbers of Everyday Entities

Skills: Planning, Performing, Observing, Analyzing, Communicating

SKILLS HANDBOOK A3.2, A3.3, A6

In this activity you will determine the number of molecules or formula units in samples of common objects.

Equipment and Materials: lab apron; blackboard; balance; dropper; graduated cylinder; chalk; water; disposable cups; penny

1. Plan procedures to determine the following quantities of matter:
   (i) the mass of chalk required to write your name. Assume that chalk is pure calcium carbonate, $CaCO_3$.
   (ii) the mass of water that you can hold in your hand
   (iii) the mass of water drops that will fit on a penny without overflowing
2. With your teacher's approval, carry out your procedures. Record your observations.

A. Calculate the amount of each substance. T/I

B. Calculate the number of molecules or formula units in each substance. T/I

6.5 Summary

- In general, the number of entities in a sample, N, is determined by multiplying the amount, n, by Avogadro's constant, N_A.
  $N = n\,N_A$
- The number of atoms or ions in a sample of a compound equals the number of atoms or ions per molecule or formula unit times the number of molecules or formula units in the sample.

6.5 Questions

1. Calculate the number of molecules present in each of the following samples: T/I
   (a) 2.0 mol of ammonia, NH_3
   (b) 1.6 mg of water, H_2O
2. Calculate the amount of substance in each of the following samples: T/I
   (a) 3.01×10^{23} atoms of copper
   (b) 6.02×10^{25} molecules of hydrogen chloride (in hydrochloric acid)
   (c) 42.0 g of calcium carbonate, $CaCO_3$ (in shellfish)
3. A sample of carbon dioxide contains 4.2×10^{24} molecules. T/I
   (a) How many atoms of carbon and how many atoms of oxygen are in the sample?
   (b) What is the mass of the sample?
4. Sodium hypochlorite, NaClO, is the active ingredient in chlorine bleaches. A sample of chlorine bleach contains 120.0 g of sodium hypochlorite. How many formula units of NaClO are in the sample? T/I
5. Acetylsalicylic acid, $C_9H_8O_4$, is the pain reliever in many headache medications. A typical tablet contains 250.0 mg of acetylsalicylic acid. T/I
   (a) Calculate the amount of acetylsalicylic acid in the tablet.
   (b) Calculate the number of hydrogen atoms in the amount of acetylsalicylic acid in the tablet determined in (a).
6. A 2.00 L birthday balloon contains 0.327 g of helium. T/I
   (a) Calculate the number of helium atoms in the balloon.
   (b) A blimp occupies a volume of 5.74×10^6 L. (**Figure 4**) Assuming that the blimp has the same number of atoms per litre as the party balloon under the same conditions, how many helium atoms are in the blimp?

Figure 4 A helium-filled blimp

7. Arsenic(III) oxide, As_2O_3, is a toxic compound once used as a wood preservative. Recently, however, this compound has been used in experimental cancer treatments. The toxicity of this compound in humans is not known. However, 14 mg of As_2O_3 is known to be lethal to half of the laboratory rats in a typical group. T/I
   (a) Calculate the number of formula units of As_2O_3 in this mass.
   (b) What number of oxide ions is present in this mass?
8. Sodium monofluorophosphate, Na_2FPO_3, is a common additive in toothpastes to help prevent tooth decay. Analysis indicates that 0.76 % of the mass of a 120.0 g tube of toothpaste is sodium monofluorophosphate. T/I
   (a) Calculate the amount of sodium monofluorophosphate in the tube.
   (b) Calculate the number of sodium ions and monofluorophosphate ions in the tube.
9. Gold medals from the 2010 Vancouver Olympic and Paralympic Games were made of silver with a thin coating of gold (**Figure 5**). A typical "gold" medal contains 509.0 g of silver and 6.0 g of gold. T/I
   (a) Calculate the number of silver atoms and the number of gold atoms in a gold medal.
   (b) Compare the costs of a gold-plated medal and a solid gold medal, each with a total mass of 515 g. Research today's cost of gold and silver on the stock market.

Figure 5 Gold medal from the 2010 Vancouver Paralympic winter games

GO TO NELSON SCIENCE

6.6 The Composition of Unknown Compounds

Figure 1 (a) A scientist in a "clean room" makes final adjustments to the Thermal and Evolved Gas Analyzer (TEGA) that the *Phoenix* used to detect water and organic compounds. (b) An artist's impression shows the robotic arm of the *Phoenix* (lower left) about to take a soil sample.

Are there signs of life on Mars? In 2008, NASA sent a robotic device called the *Phoenix Mars Lander* to the red planet to find out. The *Phoenix* searched Martian soil for water and for organic compounds that could only be produced by living things. Scientists equipped the *Phoenix* with a state-of-the-art analytical lab. A key component for this lab was a Thermal and Evolved Gas Analyzer (TEGA) (**Figure 1**). The TEGA consisted of a furnace connected to a mass spectrometer, which measures molecular masses.

In a typical analysis, the *Phoenix's* robotic arm scooped samples of Martian soil into the TEGA's furnace. Then the furnace temperature increased to vaporize any water or organic compounds present. The mass spectrometer "sniffed" the vapours to determine the molecular masses and concentrations of any compounds. Unfortunately, the TEGA's analyses showed no evidence of organic compounds; however, the *Phoenix* did find water!

Combustion Analysis

Chemists also use similar "burn and sniff" analytical tools to help identify unknown organic compounds on Earth (**Figure 2**). In this technology, a sample is burned with a plentiful oxygen supply, guaranteeing that complete combustion occurs. Any carbon dioxide and water produced during combustion is absorbed by chemicals in two successive chambers. Chemists use the change in the mass of these chambers to determine the mass of carbon dioxide and water produced. A mass spectrometer can also help to determine the molecular masses of the compounds.

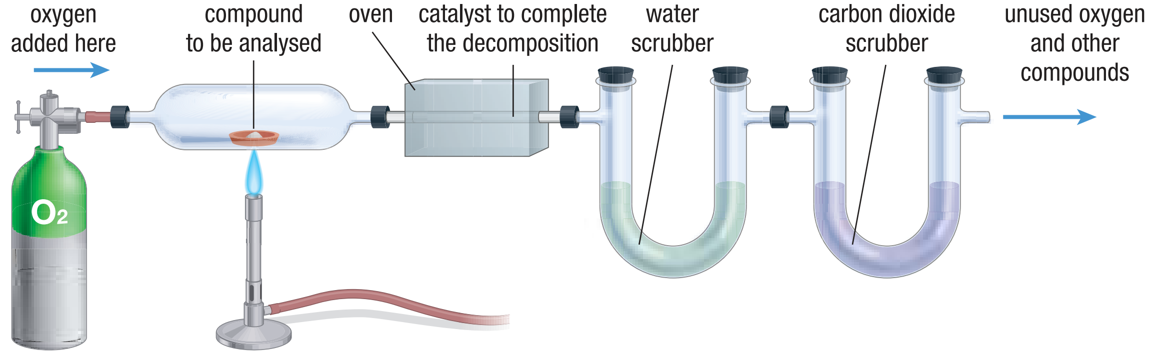


Figure 2 A combustion analysis apparatus is used to determine the composition of an unknown organic compound. The amounts of carbon dioxide and water produced in the analysis can be used to determine the percentage of carbon and hydrogen in the compound.

Percentage Composition

percentage composition the percentage, by mass, of each element in a compound

We can determine the ratio of atoms or ions in a sample of a compound if we know the relative masses of its elements. With this information we can use the molar masses to convert from the relative masses to numbers of entities. Relative masses are usually given as percentage composition. **Percentage composition** is the percentage, by mass, of each element in the compound. Note that percentage composition is not the same as the ratios of atoms of elements.

Table 1 shows the results of combustion analysis of an organic compound that contains the elements carbon, hydrogen, and oxygen. We can write the general formula for this compound as $C_xH_yO_z$.

Table 1 Combustion Analysis of $C_xH_yO_z$

Sample number	Mass of compound (g)	Mass of carbon (g)	Mass of hydrogen (g)	Mass of oxygen (g)
1	2.000	0.801	0.168	1.030
2	4.000	1.601	0.336	2.061
3	6.000	2.402	0.504	3.090

Investigation 6.6.1

Popping Percentage Composition (p. 301)

In this investigation you will measure the mass of popping corn before and after cooking, and calculate the percentage of water in the kernels.

In Sample 1, for example, 0.801 g of the 2.00 g sample is carbon. The percentage of an element in a compound is determined using the equation

$$\% \text{ element} = \frac{m_{\text{element}}}{m_{\text{sample}}} \times 100\ \%$$

Therefore the percentage of carbon in the compound is

$$\begin{aligned} \% \text{ C} &= \frac{m_{\text{carbon}}}{m_{\text{sample}}} \times 100\ \% \\ &= \frac{0.801\ \cancel{\text{g}}}{2.00\ \cancel{\text{g}}} \times 100\ \% \\ &= 40.0\ \% \end{aligned}$$

Repeating this calculation for the other elements and samples gives the percentages of carbon, hydrogen, and oxygen in the organic compound (**Table 2**).

Table 2 Combustion Analysis of $C_xH_yO_z$

Sample number	Mass of compound (g)	Percentage of carbon (%)	Percentage of hydrogen (%)	Percentage of oxygen (%)
1	2.000	40.0	8.40	51.5
2	4.000	40.0	8.40	51.5
3	6.000	40.0	8.40	51.5

Note that the percentages of carbon, hydrogen, and oxygen in this compound do not change regardless of the mass of the original compound. In fact, this result is valid for all chemical compounds and is known as the **law of definite proportions**.

law of definite proportions the statement that a compound always contains the same proportion of elements by mass

Law of Definite Proportions

The elements in a compound are always present in the same proportion by mass.

This law may seem obvious because we know that the number of atoms of each element in the compound is fixed. For example, the chemical formula H_2O states there are 2 hydrogen atoms and 1 oxygen atom in each molecule. Therefore, the proportion, by mass, of hydrogen to oxygen is always the same regardless of whether you pour a drop or a tubful of water (**Figure 3**).

However, the idea that a substance could have a fixed composition was met with considerable opposition when it was first proposed by the French chemist Joseph Proust in 1799. It only began to gain acceptance once John Dalton suggested that matter was composed of atoms in 1803.

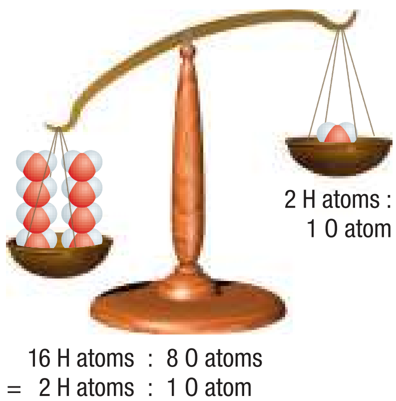


Figure 3 The mass of the water sample on the left is greater than the mass of the sample on the right. However, the proportion, by mass, of hydrogen to oxygen in each sample is the same.

Tutorial 1 Calculating Percentage Composition

You can calculate percentage composition from either lab data or the chemical formula of the compound. We will analyze both of these situations.

USING LAB DATA

The percentage of each element in a compound can be determined using the following formula:

$$\% \text{ element} = \frac{m_{\text{element}}}{m_{\text{sample}}} \times 100 \ \%$$

- Note that both masses should have identical units. They need not both be in grams.
- Note also that mass units "cancel" as a result of division. Hence the answer has no units.

Sample Problem 1: Calculating Composition from Lab Data

A 500.00 mg tablet of Aspirin, $C_9H_8O_4$, contains 300.00 mg carbon and 8.08 mg hydrogen. The remaining mass is oxygen. Determine the percentage composition of Aspirin.

Given: $m_{C_9H_8O_4} = 500.00$ mg; $m_C = 300.00$ mg; $m_H = 8.08$ mg

Required: percentage composition of each element: % C; % H; % O

Solution:

Step 1. Use the % element formula, above, to calculate the percentage composition of each element for which you know the mass.

$$\begin{aligned} \% \text{ C} &= \frac{m_{\text{carbon}}}{m_{\text{sample}}} \times 100 \ \% \\ &= \frac{300.00 \ \cancel{\text{mg}}}{500.00 \ \cancel{\text{mg}}} \times 100 \ \% \\ &= 60.000\% \end{aligned}$$

$$\begin{aligned} \% \text{ H} &= \frac{m_{\text{hydrogen}}}{m_{C_9H_8O_4}} \times 100 \ \% \\ &= \frac{8.08 \ \cancel{\text{mg}}}{500.00 \ \cancel{\text{mg}}} \times 100 \ \% \\ &= 1.62 \ \% \end{aligned}$$

Step 2. Determine the percentage of the element for which you do not know the mass by subtracting the sum of the percentages of the other elements from 100 %. In this case, the remaining mass of the Aspirin tablet is oxygen. Therefore,

$$\begin{aligned} \% \text{ O} &= 100 \ \% - (60.000 \ \% + 1.62 \ \%) \\ &= 100 \ \% - (61.62 \ \%) \\ &= 38.38 \ \% \end{aligned}$$

Statement: The percentage composition of Aspirin is 60.000 % carbon, 1.62 % hydrogen, and 38.38 % oxygen.

Practice

1. Researchers have isolated a compound that contains only nitrogen and oxygen. They found that a 4.60 g sample of this compound contains 1.40 g of nitrogen. Find the percentage composition of this compound. T/I [ans: 30.4 % N, 69.6 % O]
2. A lab technician finds that an 11.5 g sample of a liquid compound contains 6.00 g of carbon and 1.51 g of hydrogen. The remaining mass is oxygen. Determine the percentage composition of this compound. T/I [ans: 52.2 % C, 13.1 % H, 34.7 % O]

USING THE CHEMICAL FORMULA

Percentage composition has many applications in consumer products. Gardeners, for example, distinguish types of fertilizers by a three-number code (**Figure 4**). These numbers refer to the percentage of three nutrients in the product: nitrogen, phosphorus, and potassium.

Figure 4 This fertilizer contains 4 % nitrogen, 8 % phosphorus, and 12 % potassium by mass. The remaining percentage is other components.

When you are calculating percentage composition from a chemical formula, you calculate the contribution of each element toward the total mass. You will need to look up the atomic mass of each element and consider how many atoms there are in a molecule (or formula unit) of the compound. You can use either atomic mass units (u) or grams (g) as the units for mass, but remember to use the same unit for each element.

The molecular mass (or mass of one formula unit) of a compound has the same numerical value as its molar mass. Therefore, the atomic mass unit, u, or g/mol can be used in the percentage composition calculation. The atomic mass unit, u, is used in the following sample problem.

Sample Problem 2: Calculating Composition from the Chemical Formula

Determine the percentage composition of calcium hydroxide, $Ca(OH)_2$.

Given: $Ca(OH)_2$

Required: percentage composition of $Ca(OH)_2$

Solution:

Step 1. Calculate the molecular mass of the compound.

$$M_{Ca(OH)_2} = 74.10 \text{ u}$$

Step 2. Calculate the percentage of each element by dividing the total mass of the element by the molecular mass and multiplying by 100 %.

$$\% \text{ Ca} = \frac{40.08 \cancel{\text{u}}}{74.10 \cancel{\text{u}}} \times 100 \ \% = 54.09 \ \%$$

$$\% \text{ O} = \frac{32.00 \cancel{\text{u}}}{74.10 \cancel{\text{u}}} \times 100 \ \% = 43.18 \ \%$$

$$\% \text{ H} = \frac{2.02 \cancel{\text{u}}}{74.10 \cancel{\text{u}}} \times 100 \ \% = 2.73 \ \%$$

Step 3. Check that all the percentages add up to 100 %.

$$54.09 \ \% + 43.18\% + 2.73\% = 100\%$$

Statement: The percentage composition of calcium hydroxide is 54.09 % calcium, 43.18 % oxygen, and 2.73 % hydrogen.

Practice

SKILLS HANDBOOK A6.5

3. Determine the percentage composition of each of the following compounds: T/I
   (a) sodium sulfate, Na_2SO_4 [ans: (a) 32.37 % Na, 22.58 % S, 45.05 % O]
   (b) ammonium nitrate, NH_4NO_3 [ans: (b) 35.00 % N, 5.05 % H, 59.95 % O]

Percentage Composition of Alloys and Mixtures

Alloys are mixtures of different metals blended together to form useful products. Unlike compounds, the percentage composition of an alloy can vary. For example, stainless steel used to make cutlery contains about 10 % chromium by mass. Surgical stainless steel contains at least 18 % chromium. Increasing the proportion of chromium makes steel more resistant to corrosion, a critical property for implants like knee replacements.

We can also consider the percentage composition of substances in mixtures. The food industry is often concerned about the water, salt, or fat content in their products.

Investigation 6.6.2

Percentage Composition of Magnesium Oxide (p. 302)
In this controlled experiment you will calculate three different values for the percentage composition of magnesium oxide. You will then collect evidence to discover which value is correct.

Mini Investigation

Determining Percentage Composition of a Compound

Skills: Planning, Performing, Observing, Analyzing, Communicating

SKILLS HANDBOOK A2.4, A3.2, A6

In this activity you will devise a way to determine the percentage of each element in a simulated compound (**Figure 5**).

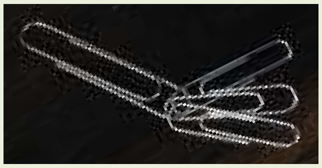


Figure 5 A paper clip "compound"

Equipment and Materials: balance; 3 molecules of the paper clip "compound"

1. Examine your compound. Make up a chemical symbol for each "element" in your compound.
2. Measure your compound's molecular mass in grams.
3. Measure the mass of 3 molecules together.
4. Plan a procedure to determine the percentage of each element in the compound.

A. Write the chemical formula of your compound. T/I C

B. Why is the unit of molecular mass for your compound in grams and not grams per mole? K/U

C. What is the percentage composition of your compound? T/I

D. What effect (if any) does increasing the numbers of molecules have on the mass of the sample and the percentage composition of the compound? T/I

6.6 Summary

- The percentage composition of a compound is the proportion of each element in the compound, by mass.
- Percentage composition can be found experimentally or from the chemical formula of the compound.
- The law of definite proportions states that a compound always has the same proportion of elements by mass. This occurs because the relative number of atoms or ions in the compound is fixed.

6.6 Questions

1. In a percentage composition investigation a compound was decomposed into its elements: 20.0 g of calcium, 6.0 g of carbon, and 24.0 g of oxygen. Determine the percentage composition of this compound. T/I
2. A researcher performed a combustion analysis on a 5.00 g sample of ethanol, the alcohol in wine. The data showed that the sample contained 2.61 g of carbon and 0.66 g of hydrogen. The researcher assumed that the remaining mass was oxygen. Calculate the percentage composition of ethanol. T/I
3. Hydrogen peroxide, H_2O_2, decomposes at room temperature to release oxygen gas. Use the law of definite proportions to explain why doubling the quantity of hydrogen peroxide doubles the quantity of oxygen produced. T/I
4. (a) A lab analysis of a 10.00 g sample of benzene indicates that it contains 9.20 g of carbon. What is the mass of hydrogen in the sample? Explain.
   (b) What is the percentage composition of benzene? T/I
5. Predict which compound in each of the following pairs has the greater percentage of carbon, by mass. Explain. T/I
   (a) C_2H_4 or C_3H_4
   (b) C_2H_4 or C_3H_6
   (c) C_2H_4 or C_2H_6
6. Calcium oxide, CaO, is a key ingredient in cement. Calcium oxide is made by heating calcium carbonate, $CaCO_3$.

   $$CaCO_3(s) \xrightarrow{\text{heat}} CaO(s) + CO_2(g)$$

   Explain what effect heat has on the following quantities: T/I
   (a) the total mass of solids in the furnace
   (b) the percentage of calcium in the solids in the furnace
7. Calculate the percentage composition of the following compounds: T/I
   (a) phosphoric acid, H_3PO_4 (gives pop a tart taste)
   (b) copper(I) sulfide, Cu_2S (found in copper ore)
   (c) iron(III) oxide, Fe_2O_3 (a component in rust)
   (d) boric acid, $B(OH)_3$ (an antiseptic)
8. Ammonium phosphate, $(NH_4)_3PO_4$, and ammonium nitrate, NH_4NO_3, are both used in the production of fertilizers. Predict which compound contains the greater percentage of nitrogen. Verify your prediction by calculating the percentage of nitrogen in each compound. T/I
9. Research, and prepare a short presentation on, gold fingerprinting. T/I C A

GO TO NELSON SCIENCE

Empirical Formulas

6.7

In the previous section you read about the *Phoenix* Mars Lander's search for signs of life. You also learned that researchers use combustion analysis to determine the percentage composition of unknown compounds. Percentage composition gives the percentage, by mass, of the elements in a compound. However, many compounds have the same percentage composition. Methanal, CH_2O, and ethanoic acid, $C_2H_4O_2$, for example, both contain 40.0 % carbon, 6.7 % hydrogen, and 53.3 % oxygen, by mass. However, these compounds have very different properties and applications (**Figure 1**).

Figure 1 (a) Solutions of methanal (formaldehyde) are used to preserve biological specimens, such as this sheep fetus. (b) Vinegar is a solution of ethanoic (acetic) acid and other substances dissolved in water.

The best way to identify an unknown compound is to determine its chemical formula. The subscripts in a chemical formula describe the number of atoms or ions of each element in the formula. How do you convert mass data to numbers of atoms in a molecule, or the number of ions in a formula unit? Earlier, you learned how to determine the number of entities, in moles, from the mass of the sample. Let's consider a compound that is 40.0 % carbon, 6.7 % hydrogen, and 53.3 % oxygen, by mass.

This compound has the same percentage of each element whatever the mass of your sample. So, we will imagine a convenient mass of the compound: 100.0 g. This sample contains 40.0 g of carbon, 6.5 g of hydrogen, and 53.5 g of oxygen. We can then determine the amount of each element (in moles) by dividing each element's mass by its molar mass (**Table 1**).

Table 1 Determining the Amount of Each Element in a 100.0 g Sample of a Compound

Amount of carbon (mol)	Amount of hydrogen (mol)	Amount of oxygen (mol)
$n_C = \dfrac{40.0\ \text{g}}{12.01\ \dfrac{\text{g}}{\text{mol}}}$	$n_H = \dfrac{6.5\ \text{g}}{1.01\ \dfrac{\text{g}}{\text{mol}}}$	$n_O = \dfrac{53.5\ \text{g}}{16.00\ \dfrac{\text{g}}{\text{mol}}}$
$= (40.0\ \cancel{\text{g}})\left(\dfrac{1\ \text{mol}}{12.01\ \cancel{\text{g}}}\right)$	$= (6.5\ \cancel{\text{g}})\left(\dfrac{1\ \text{mol}}{1.01\ \cancel{\text{g}}}\right)$	$= (53.5\ \cancel{\text{g}})\left(\dfrac{1\ \text{mol}}{16.00\ \cancel{\text{g}}}\right)$
$= 3.33$ mol	$= 6.4$ mol	$= 3.34$ mol

Note that the ratio of the amounts of carbon and oxygen is essentially 1:1. The amount of hydrogen is almost double this value. Therefore, the ratio of carbon to hydrogen to oxygen atoms in the compound is 1:2:1. We can express this ratio as the chemical formula $C_1H_2O_1$ or CH_2O.

LEARNING **TIP**

Determining Ratios
One way to determine the ratios of entities is to divide each amount by the smallest amount. The smallest amount becomes "1." The other amounts have higher values, relative to 1.

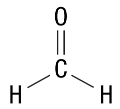
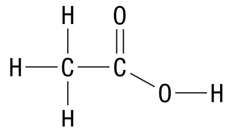
Distinguishing between Empirical Formula and Molecular Formula

empirical formula a formula that shows the simplest whole-number ratio of elements in a compound

molecular formula a formula that shows the element symbols and exact number of each type of atom in a molecular compound

CH_2O is an example of an empirical formula. The **empirical formula** gives the simplest whole-number ratio of atoms or ions in a compound. Both methanal and ethanoic acid have the same empirical formula: CH_2O. A **molecular formula** gives the exact number of each type of atom in a compound. A methanal molecule contains 1 carbon atom, 2 hydrogen atoms, and 1 oxygen atom (**Table 2**). Therefore, its empirical and molecular formulas are identical. However, an ethanoic acid molecule contains twice the number of atoms given in its empirical formula. As a result, its molecular formula is $C_2H_4O_2$.

Table 2 Comparing Empirical, Molecular, and Structural Formulas

Compound	Empirical formula	Molecular formula	Structural formula
methanal	CH_2O	CH_2O	
ethanoic acid	CH_2O	$C_2H_4O_2$	

The formula that we write for most ionic compounds is an empirical formula. For example, a tiny grain of sodium chloride (table salt) may contain quadrillions of sodium and chloride ions. Since these ions are in a 1:1 ratio, the empirical formula of sodium chloride is NaCl.

LEARNING TIP

An Analogy for Empirical Formula
Imagine a “compound” that contains exactly 2 thumbs, T, and 8 fingers, F. The molecular formula for this compound is T_2F_8. Since there is 1 thumb for every 4 fingers, the simplest ratio of thumbs to fingers is 1:4. Therefore, the empirical formula of this compound is TF_4.

Molecular formula: T_2F_8

Empirical formula: TF_4

Determining Empirical Formula

Percentage composition gives the proportion of masses of the elements in a compound. The empirical formula gives the proportion of the atoms or ions of each element. If we know the percentage composition of a compound, we can determine the empirical formula. We do this by converting the mass of each of the elements (in grams) into amount (in moles). The ratio of amounts gives the subscripts in the empirical formula.

Tutorial 1 Determining the Empirical Formula from Percentage Composition

When we know the percentage composition of a compound and want to find the empirical formula, it is helpful to consider a 100.0 g sample of the compound.

Sample Problem 1: Determining a Simple Empirical Formula

Find the empirical formula of a compound with percentage composition 35.4 % sodium and the remainder nitrogen.

Given: % Na = 35.4%, remainder is N

Required: empirical formula of compound

Solution:

Step 1. Let’s start with a convenient mass of the compound, such as 100.0 g.

Sodium makes up 35.4% of the sample; nitrogen makes up the rest:

$100.0 \text{ g} - 35.4 \text{ g} = 64.6 \text{ g}$

Step 2. Calculate the amount of each element.

$$n_{Na} = (35.4\ \cancel{g})\left(\frac{1\text{ mol}}{22.99\ \cancel{g}}\right)$$

$$= 1.540\text{ mol [extra digits carried]}$$

$$n_N = (64.6\ \cancel{g})\left(\frac{1\text{ mol}}{14.01\ \cancel{g}}\right)$$

$$= 4.6110\text{ mol [extra digits carried]}$$

Step 3. To determine the simplest ratio of the elements in the compound, divide the amount of each element by the smallest amount.

$$\frac{n_{Na}}{n_{Na}} = \frac{\cancel{1.549\text{ mol}}}{\cancel{1.5459\text{ mol}}} = 1$$

We assign sodium a value of 1.

$$\frac{n_N}{n_{Na}} = \frac{4.6110\ \cancel{\text{mol}}}{1.5459\ \cancel{\text{mol}}} = 2.98$$

According to the calculation, the ratio of sodium to nitrogen is 1:2.98. Since we cannot have a fraction of an element in a compound, the value for nitrogen is rounded off to the nearest whole number. As a result, the simplest whole-number ratio of sodium to nitrogen is 1:3.

Statement: The empirical formula of this compound is NaN_3.

EMPIRICAL FORMULAS WITH FRACTIONS

Determining the simplest ratio of the elements may give whole numbers. However, because you are working with experimental (measured) data, you should expect that your answer may include a fraction. If a number is within 0.05 of a whole number, you can round it up or down to the nearest whole number. But what do you do if one of the values is not close to a whole number? You multiply all the values by the same number to make them all whole numbers. Multiplying the subscripts by the fraction denominator gives whole numbers (**Table 3**). For example

$C_{1 \times 3}H_{1.333 \times 3} = C_3H_4$.

Sample Problem 2 illustrates this procedure.

Table 3 Ratios from Fractions

Decimal/fraction	Whole number	Multiply all subscripts by
$0.25 = \frac{1}{4}$	$4(0.25) = 1$	4
$0.33 = \frac{1}{3}$	$3(0.33) = 1$	3
$0.50 = \frac{1}{2}$	$2(0.5) = 1$	2
$0.67 = \frac{2}{3}$	$3(0.67) = 2$	3
$0.75 = \frac{3}{4}$	$4(0.75) = 3$	4

Sample Problem 2: Determining a More Complex Empirical Formula

Determine the empirical formula of a compound that contains 69.9 % iron and 30.1 % oxygen by mass (**Figure 2**).

Figure 2 An artist's impression of Mars shows the ground as reddish-orange. Scientists suggest that the rocks are this colour because of their high iron oxide content.

Given: % Fe = 69.9 %, % O = 30.1 %

Required: empirical formula of the unknown compound, Fe_xO_y

Solution:

A 100.0 g sample of this compound contains 69.9 g of iron and 30.1 g of oxygen.

$M_{Fe} = 55.85$ g/mol; $M_O = 16.00$ g/mol

Step 1. Calculate the amount of each element in the 100.0 g sample.

$$n_{Fe} = (69.9\ \cancel{g})\left(\frac{1\text{ mol}}{55.85\ \cancel{g}}\right)$$

$$= 1.2516\text{ mol [extra digits carried]}$$

$$n_O = (30.1\ \cancel{g})\left(\frac{1\text{ mol}}{16.00\ \cancel{g}}\right)$$

$$= 1.8813\text{ mol [extra digits carried]}$$

Step 2. Divide the amount of each element by the smallest amount

$$\frac{n_{Fe}}{n_{Fe}} = \frac{\cancel{1.2516\text{ mol}}}{\cancel{1.2516\text{ mol}}} = 1$$

Assign iron a value of 1.

$$\frac{n_O}{n_{Fe}} = \frac{1.8813\ \cancel{mol}}{1.2516\ \cancel{mol}} = 1.50$$

Assign oxygen a value of 1.50, relative to iron.

Step 3. These calculations give an empirical formula of $Fe_1O_{1.5}$. Subscripts in a chemical formula are normally whole numbers. Multiplying both subscripts by 2 gives Fe_2O_3.

Statement: The empirical formula of the compound is Fe_2O_3.

Sample Problem 3: Determining an Empirical Formula with Three Elements

Determine the empirical formula of a compound that contains 52.2 % carbon, 6.15 % hydrogen, and 41.7 % oxygen.

Given: % C = 52.2%, % H = 6.15 %, % O = 41.7 %

Required: empirical formula of compound, $C_xH_yO_z$

Solution:

A 100.0 g sample of this compound contains 52.2 g of carbon, 6.15 g of hydrogen, and 41.7 g of oxygen.

$M_C = 12.01$ g/mol; $M_H = 1.01$ g/mol; $M_O = 16.00$ g/mol

Step 1. Determine the amount of each element in the 100.0 g sample.

$$n_C = (52.2\ \cancel{g})\left(\frac{1\ \text{mol}}{12.01\ \cancel{g}}\right)$$

$$n_C = 4.3464\ \text{mol}\ [2\ \text{extra digits carried}]$$

$$n_H = (6.15\ \cancel{g})\left(\frac{1\ \text{mol}}{1.01\ \cancel{g}}\right)$$

$$n_H = 6.0891\ \text{mol}\ [2\ \text{extra digits carried}]$$

$$n_O = (41.7\ \cancel{g})\left(\frac{1\ \text{mol}}{16.00\ \cancel{g}}\right)$$

$$n_O = 2.6063\ \text{mol}\ [2\ \text{extra digits carried}]$$

Step 2. Divide the amount of each element by the smallest amount.

$$\frac{n_O}{n_O} = \frac{\cancel{2.6063\ mol}}{\cancel{2.6063\ mol}} = 1$$

Assign oxygen a value of 1.

$$\frac{n_C}{n_O} = \frac{4.3464\ \cancel{mol}}{2.6063\ \cancel{mol}} = 1.67$$

Assign carbon a value of 1.67.

$$\frac{n_H}{n_O} = \frac{6.0891\ \cancel{mol}}{2.6063\ \cancel{mol}} = 2.29$$

Assign hydrogen a value of 2.29.

Step 3. These calculations give an empirical formula of $C_{1.67}H_{2.29}O_1$. Multiplying each of the subscripts by 3 gives $C_5H_7O_3$.

Statement: The empirical formula of the compound is $C_5H_7O_3$.

Practice

SKILLS HANDBOOK A6.3, A6.5

1. Write the empirical formulas of compounds with the following percentage compositions: T/I C
   (a) 20.2 % Al, 79.8 % Cl [ans: $AlCl_3$]
   (b) 18.4 % C, 21.5 % N, and the rest K [ans: KCN]
   (c) 52.9 % Al, and the rest O [ans: Al_2O_3]
   (d) 50.85 % C, 8.47 % H, and 40.68 % O [ans: $C_5H_{10}O_3$]

6.7 Summary

- The empirical formula gives the simplest whole-number ratio of the elements in a compound.
- Most chemical formulas for ionic compounds are empirical formulas.
- The empirical formula can be determined from percentage composition data, using molar mass to find the relative amount of each element in the compound.
- The molecular formula gives the exact number of atoms of each element in a molecular compound. The molecular formula is sometimes, but not always, the same as the empirical formula.

6.7 Questions

1. Before you can determine an empirical formula,
   (a) what experimental evidence do you need?
   (b) what researched data do you need? K/U
2. Which of the following are empirical formulas? Explain. K/U

   $C_2H_4O_2$, H_2CO_3, $K_2Cr_2O_7$, $C_3H_6O_3N$
3. Which of the following pairs of compounds have the same empirical formula? Write the empirical formula in each case. T/I C
   (a) NO_2, N_2O_4
   (b) C_3H_6, C_4H_7
   (c) C_2H_2, C_6H_6
   (d) $C_{12}H_{10}O_2$, C_6H_5O
4. Identify the multiplier that converts the subscripts in these formulas into the nearest whole number. Then write the empirical formula of each compound. T/I C
   (a) $CH_{3.5}$
   (b) $C_{0.67}H_{0.67}O$
   (c) $NaSO_{1.5}$
   (d) $C_{1.34}H_{3.66}O$
5. The percentage composition of ascorbic acid (vitamin C) is 40.9 % carbon and 4.55 % hydrogen. The remainder is oxygen. Determine the empirical formula of ascorbic acid. T/I C
6. Nitrogen and oxygen form two different compounds with the following percentage compositions:

   Compound 1: 46.7 % N; 53.3 % O

   Compound 2: 30.4 % N; 69.6 % O

   Determine the empirical formula of each compound. T/I C
7. Nylon-6 is a plastic used to make strings for musical instruments. Nylon-6 consists of 63.68 % carbon, 9.80 % hydrogen, and 12.38 % nitrogen. The remainder is oxygen. Determine the empirical formula of nylon-6. T/I C
8. Aluminum carbide is a hard, abrasive substance used in some high-speed cutting tools. Aluminum makes up 75 % of the mass of this compound and carbon makes up the rest. Determine the empirical formula of aluminum carbide. T/I C
9. A 21.7 g sample of a compound of mercury and oxygen is decomposed into its elements by heating (**Figure 3**). If 20.1 g of mercury is collected, is the empirical formula of the compound HgO or Hg_2O? Why? T/I

Figure 3 Heating a compound of mercury and oxygen leaves a mirror-like coating of mercury lining the test tube.

10. A compound contains an unknown element X. This element forms an oxide with chemical formula X_2O. Experimental tests show that X_2O contains 25.8 % oxygen, by mass. T/I
    (a) What is the percentage of element X in the compound?
    (b) Use this information to identify element X.
11. Glucose and fructose are both examples of sugars. T/I C A
    (a) Research and compare their empirical, molecular, and structural formulas.
    (b) Research the correlation between obesity and these sweeteners in pop. Which sugar appears to be associated with obesity?

GO TO NELSON SCIENCE

Drug-Contaminated Currency

SKILLS HANDBOOK A4

ABSTRACT

In 1999 the majority of banknotes used in London, England were tainted with illicit drugs such as cocaine. Some become contaminated by direct contact with banknotes used to inhale cocaine or from surfaces already contaminated with cocaine. However, money-counting machines used in banks were suspected of being the major source of the contamination. An investigation was designed to verify this suspicion. In the investigation, clean banknote-sized strips of paper were passed through counting machines in banks throughout the United Kingdom. The strips were then analyzed for cocaine residue. A significant number of strips counted in the London area became contaminated with cocaine, confirming that counting machines are a potential source of cocaine contamination.

Introduction

People who buy illicit drugs such as cocaine and heroin usually pay with cash. This method of payment has been convenient and difficult to trace—until now. Analytical chemists have discovered that banknotes exchanged in a drug transaction are often contaminated with tiny amounts of drugs. In fact, according to a 1999 BBC report, 99 % of banknotes used in London are contaminated with cocaine. The amount of the illegal substance on any given banknote is invisible to the naked eye: a few micrograms or less. To put this amount into perspective, imagine a grain of salt divided into 1000 pieces.

One source of contamination is direct skin contact with someone who has handled the drug. Cocaine readily absorbs into the oils on skin. From there, it can easily be transferred to banknotes.

Another suspected source of contamination is the banking industry; specifically, the machines used by banks to count money. Researchers from the University of Bristol in the United Kingdom designed an investigation to test whether counting machines are indeed a source of the contamination.

The Design of the Investigation

Different types of paper were first examined under an electron microscope to find a type of paper whose surface best matched that of real British banknotes (**Figure 1**).

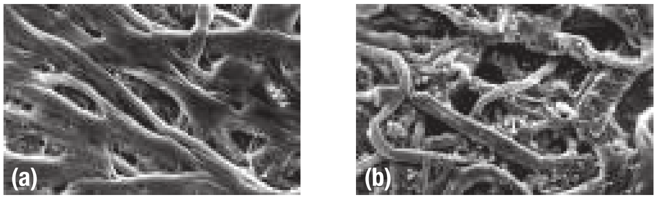


Figure 1 (a) The surface of a British banknote (magnification 500X) showing the individual paper fibres. The gaps between fibres are places where cocaine and other materials could become entrapped. (b) The surface of the paper chosen to simulate British banknotes

Technicians wore plastic gloves while handling the paper, and used clean aluminum foil to cover all surfaces that might come in contact with the paper. They cut the paper chosen for the investigation into strips the same size as banknotes and sealed them, in bundles of 50, in plastic bags. The bags were then sent to selected banks throughout the United Kingdom. Each bank passed these strips through their counting machines, following detailed instructions provided by the researchers.

Analysis Technology

The test strips were then collected and analyzed for cocaine residue using a combination of thermal desorption and mass spectroscopy (**Figure 2**). Thermal desorption involves placing each strip between two metal plates heated to 285 °C for one second. This is long enough to vaporize organic compounds such as cocaine on the surface without damaging the paper. The resulting vapour is then passed into a mass spectrometer.

Figure 2 Thermal desorption apparatus vaporizing organic compounds off a banknote

The mass spectrometer needs only a tiny quantity—a microgram or less—to determine the molecular masses of all the vaporized compounds in the sample (**Figure 3**).

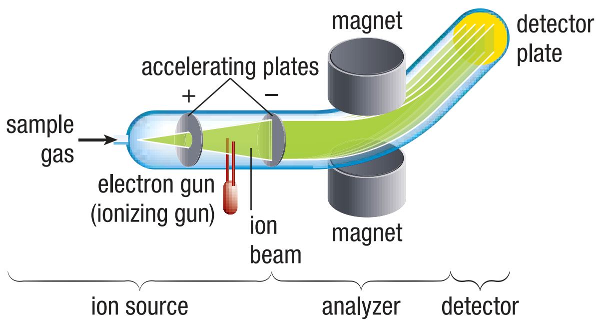


Figure 3 A mass spectrometer measures the masses of ionized entities by measuring their deflection as they pass through a strong magnetic field.

During the analysis, molecules in the sample are broken into fragments and ionized. The fragment pattern produced by the mass spectrometer is often unique to the substance, much like a fingerprint. This pattern, together with the molecular mass, is usually sufficient to determine the molecular formula of a compound and identify the compound (**Figure 4**).

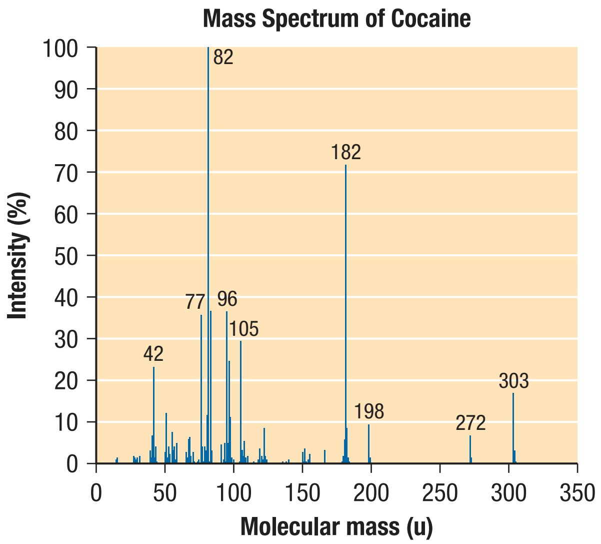


Figure 4 The largest molecular mass, 303 u, is the molecular mass of cocaine. The other spikes represents fragments of cocaine molecules. This fragment pattern identifies cocaine.

The Evidence

Based on the evidence (**Table 1**), researchers concluded that the counting machines used by banks are partially responsible for the contamination of banknotes with cocaine.

Table 1 Analysis of the Test Strips

Level of cocaine contamination	Proportion of sample notes (%)
high	14
low	46
none	40

Banks in the London area returned the highest proportion of contaminated paper. This result was expected, since police records showed that cocaine use in London was greater than in other parts of the United Kingdom. Furthermore, cocaine contamination was the lowest in northern regions of the United Kingdom where cocaine use was known to be less. The amount of cocaine on the test strips was also much lower than the amount present on banknotes seized as evidence in cocaine-related crimes.

Further Reading

Carter, J.F., Sleeman, R., Parry, J. (2003). The distribution of controlled drugs on banknotes via counting machines. *Forensic Science International*, 132, 106–112.

Dixon, S.J., Brereton, R.G., Carter, J.F., Sleeman, R. (2006). Determination of cocaine contamination on banknotes using tandem mass spectrometry and pattern recognition. *Analytica Chimica Acta*, 559(1), 54–63.

GO TO NELSON SCIENCE

6.8 Questions

1. This investigation concluded that "clean" banknotes were being contaminated by cocaine. Did it identify where the cocaine originally came from? Explain your answer. K/U A
2. Why was it necessary to examine the paper used in the investigation under the microscope? K/U
3. Describe two ways in which this experiment was controlled. T/I
4. Unlike real banknotes, the paper used in this investigation contained no ink. How do you think the presence of ink might affect the outcome? T/I
5. How is a mass spectrum of a compound similar to a fingerprint? K/U
6. Heroin is an illicit drug that is less chemically stable than cocaine. How do you think this property would influence the outcome of the investigation if it were repeated to test for heroin on test strips? T/I
7. (a) Research and list at least three other uses of mass spectroscopy.
   (b) Read a scientific article related to one of these uses of mass spectroscopy. Write an abstract of the article. T/I C A

GO TO NELSON SCIENCE

6.9 Molecular Formulas

How clean is the air in your home? Despite the best air-filtering technology available, indoor air pollution is unavoidable. Some outdoor pollutants follow you in as you enter your home. However, most indoor pollutants come from one of two sources. They are either produced by chemical processes in the home, such as home heating, or released—"off-gassed"—by products that we bring into the home. A newly installed synthetic carpet, for example, releases a cocktail of volatile organic compounds (VOCs) into the air (**Figure 1**). You will learn more about indoor air pollutants in Section 11.5.

Figure 1 When new synthetic carpeting is installed, allow the area to air out for quite some time. This permits VOCs from the carpet and adhesives to escape.

You might recall that organic compounds consist mostly of carbon and hydrogen. VOCs are compounds that readily become gases at room temperature. VOCs can be natural or synthetic. For example, ethyl ethanoate is a VOC used in some nail polish removers and glues. It has relatively low toxicity and also occurs naturally in wine. Other VOCs, such as methylbenzene (toluene), are more toxic. Methylbenzene is a solvent used in glues and some flooring products. Toxic VOCs need to be handled with care to minimize exposure.

You may know someone who has allergies to "something in the air." The detection of indoor air pollutants is important in identifying possible causes of allergies. Fortunately, recent advances in analytical chemistry have made the detection of trace amounts of compounds in air possible. Scientists often use mass spectroscopy to detect and identify compounds in indoor air. As you learned in Section 6.8, a mass spectrometer is an analytical device that identifies compounds based on their molecular mass and characteristic fragment pattern. For example, **Figure 2** shows the mass spectrum of a contaminant in indoor air. This readout indicates that the compound has a molecular mass of 88 u.

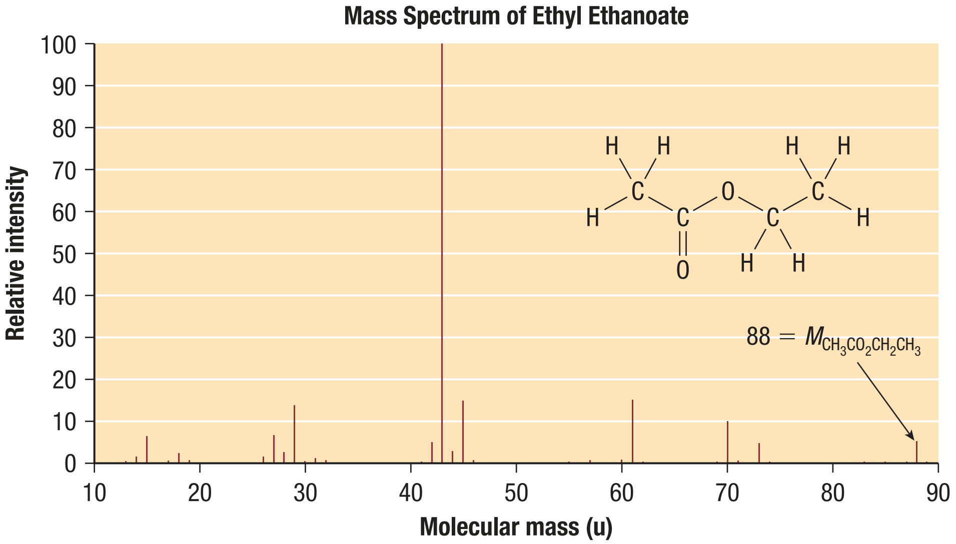
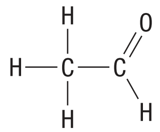
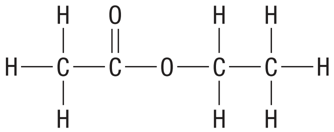
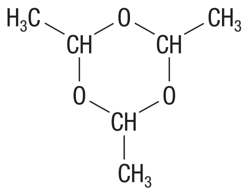


Figure 2 The largest molecular mass, 88 u, is the molecular mass of ethyl ethanoate. The other mass data are the masses of fragments of the molecule produced during the analysis. The molecular mass and fragment pattern together identify the compound.

The Importance of Molecular Formulas

Chemists use mass spectroscopy data together with an empirical formula to determine the molecular formula of a compound. Remember that the empirical formula gives only the simplest ratio of the atoms or ions of each element in the compound. The molecular formula, however, gives the exact number of atoms of each element present (Section 6.7). When we use the term "chemical formula" for a molecular compound, we are referring to the molecular formula. **Table 1** shows the molar mass and molecular formula of three organic compounds with the empirical formula C_2H_4O. However, only ethyl ethanoate has the molar mass 88.10 g/mol.

Table 1 Data for Identifying Three Organic Compounds with the Empirical Formula C_2H_4O

Compound	Molar mass (g/mol)	Molecular formula	Structural formula
ethanal (acetaldehyde) • used in the production of particle board	44.05	C_2H_4O	
ethyl ethanoate (ethyl acetate) • a solvent in glue	88.10	$C_4H_8O_2$	
2, 4, 6-trimethyl-1,3,5-trioxane (paraldehyde) • a central nervous system depressant, once used as an anesthetic	132.15	$C_6H_{12}O_3$	

Note that the molar masses of ethyl ethanoate and paraldehyde are whole-number multiples of the molecular mass of the empirical formula C_2H_4O. That explains why the subscripts in their molecular formulas are also multiples of the subscripts in C_2H_4O. For example, since 3(44.05 u) = 132.15 u, the molecular formula of paraldehyde must be $C_{2\times3}H_{4\times3}O_{1\times3}$, which equals $C_6H_{12}O_3$.

Mini **Investigation**

Comparing Molecules and Molecular Formulas

Skills: Predicting, Performing, Observing, Analyzing, Communicating — SKILLS HANDBOOK A2.4

In this activity you will construct molecules with the same empirical formula but different molecular formulas.

Equipment and Materials: molecular model kit

1. Construct 3 molecules of methanal (formaldehyde), CH_2O. Close the model kit box. For the rest of the activity, you will use only the atoms in these 3 molecules.
2. Take apart the methanal molecules and use some of the atoms to construct a molecule of a compound that has the molecular formula $C_2H_4O_2$. Note the number of starting atoms that are left over. Record the structural formula for this compound.
3. Using exactly the same atoms, construct a different molecule with the same molecular formula: $C_2H_4O_2$. Record the structural formula for this compound.
4. Construct a molecule with the molecular formula $C_3H_6O_3$. Record the structural formula for this compound.
5. Rearrange the atoms from Step 4 to form a different molecule. Record the structural formula for this compound.

A. What is the molecular mass of one molecule of methanal? T/I

B. How is the molar mass of methanal similar to its molecular mass and yet quite different? K/U

C. Predict the molecular mass of each molecule that you made in Steps 2, 3, 4, and 5. Comment on your predictions. T/I

D. Explain how the molecular masses of $C_2H_4O_2$ and $C_3H_6O_3$ can be predicted from the molecular mass of methanal. T/I

E. Why is determining the molecular formula of a compound an inadequate method of identifying the compound? Suggest a better way of identifying the compound. T/I

Investigation 6.9.1

Determining the Formula of a Hydrate (p. 304)

You will determine the formula of a sample of hydrated copper(II) sulfate by determining its percentage composition.

Determining Molecular Formulas

Chemists find that the molecular formula of a compound is far more useful than an empirical formula. This is because the molecular formula gives the exact composition of the compound, even though it does not always conclusively identify the compound. Determining the molecular formula from its empirical formula is an important step in analyzing an unknown compound.

Tutorial 1 Determining Molecular Formulas from Empirical Formulas

A compound's molar mass (based on the molecular formula) is always a whole-number multiple of the molar mass of the empirical formula. You can find this multiple, *x*, by dividing the molar mass of the compound by the molar mass of the empirical formula:

$$x = \frac{\text{molar mass of compound}}{\text{molar mass of empirical formula}}$$

You can then use this value to multiply the subscripts in the empirical formula to determine the subscripts in the molecular formula.

Sample Problem 1: Finding a Molecular Formula Given Its Empirical Formula

Determine the molecular formula of a compound with empirical formula CH_2 and molar mass 84.18 g/mol.

Given: empirical formula CH_2; molar mass of compound = 84.18 g/mol

Required: molecular formula of compound, C_xH_y

Solution:

Step 1. Find the empirical molar mass by adding the molar mass of each of the elements.

$$M_{CH_2} = 1\left(12.01\ \frac{\text{g}}{\text{mol}}\right) + 2\left(1.01\ \frac{\text{g}}{\text{mol}}\right)$$

$$M_{CH_2} = 14.03\ \frac{\text{g}}{\text{mol}}$$

Step 2. Solve for *x*, the mass multiple.

$$x = \frac{\text{molar mass of compound}}{\text{molar mass of empirical formula}}$$

$$= \frac{84.18\ \cancel{\frac{\text{g}}{\text{mol}}}}{14.03\ \cancel{\frac{\text{g}}{\text{mol}}}}$$

$$x = 6.000$$

Step 3. Therefore, the molar mass of the compound is 6 times the molar mass of the empirical formula. This means that the compound contains 6 times as many of each atom as given in the empirical formula. Multiplying each of the subscripts by 6 gives C_6H_{12}.

Statement: The molecular formula of the compound is C_6H_{12}.

Practice

SKILLS HANDBOOK A6.3, A6.5

1. Determine the molecular formulas given the following empirical formulas and molar masses: T/I
   (a) HO; 34.02 g/mol [ans: H_2O_2]
   (b) SO_2; 64.06 g/mol [ans: SO_2]
   (c) KSO_4; 270.32 g/mol [ans: $K_2S_2O_8$]
   (d) $C_3H_5O_3$; 445.40 g/mol [ans: $C_{15}H_{25}O_{15}$]

USING PERCENTAGE COMPOSITION AND MOLAR MASS DATA

You can also calculate the molecular formula of a compound if you know its percentage composition and its molar mass.

Sample Problem 2: Using Percentage Composition Data

Determine the molecular formula of vitamin C (ascorbic acid). This compound contains 40.5 % carbon, 4.6 % hydrogen, and 54.5 % oxygen. Its molar mass is 176.14 g/mol.

Given: 40.5 % C; 4.6 % H; 54.5 % O; molar mass = 176.14 g/mol

Required: molecular formula of ascorbic acid, $C_xH_yO_z$

Solution: A 100.0 g sample of this compound contains 40.5 g C, 4.6 g H, 54.5 g O.

Step 1. Determine the amount of each element in a 100.0 g sample.

$$n_C = (40.5\ \cancel{g})\left(\frac{1\ mol}{12.01\ \cancel{g}}\right)$$

$$n_C = 3.3722\ mol\ [\text{extra digits carried}]$$

$$n_H = (4.6\ \cancel{g})\left(\frac{1\ mol}{1.01\ \cancel{g}}\right)$$

$$n_H = 4.5545\ mol\ [\text{extra digits carried}]$$

$$n_O = (54.5\ \cancel{g})\left(\frac{1\ mol}{16.00\ \cancel{g}}\right)$$

$$n_O = 3.4063\ mol\ [\text{extra digits carried}]$$

A 100 g sample contains 3.37 mol of carbon, 4.55 mol of hydrogen, and 3.41 mol of oxygen.

Step 2. To determine the simplest ratio, divide the amount of each element by the smallest amount.

$$\frac{n_C}{n_C} = \frac{\cancel{3.3722\ mol}}{\cancel{3.3722\ mol}} = 1$$

$$\frac{n_O}{n_C} = \frac{3.4063\ \cancel{mol}}{3.3722\ \cancel{mol}} = 1.01$$

$$\frac{n_H}{n_C} = \frac{4.5545\ \cancel{mol}}{3.3722\ \cancel{mol}} = 1.35$$

We assign both carbon and oxygen a value of 1 and hydrogen a value of 1.35.

Step 3. These calculations give an empirical formula of $C_1H_{1.35}O_1$. Multiplying the subscripts by 3 gives $C_3H_4O_3$. Therefore, the empirical formula of the compound is $C_3H_4O_3$.

Step 4. Determine the molar mass of $C_3H_4O_3$.

$$M_{C_3H_4O_3} = 3(12.01\ g/mol) + 4(1.01\ g/mol) + 3(16.00\ g/mol)$$

$$M_{C_3H_4O_3} = 88.07\ g/mol$$

Step 5. Solve for the mass multiple.

$$x = \frac{\text{molar mass of compound}}{\text{molar mass of empirical formula}}$$

$$= \frac{176.14\ \cancel{g/mol}}{88.07\ \cancel{g/mol}}$$

$$x = 2.000$$

Step 6. The molar mass of the compound is twice the molar mass of the empirical formula. Multiplying each of the subscripts by 2 gives $C_6H_8O_6$.

Statement: The molecular formula of the compound is $C_6H_8O_6$.

Practice

SKILLS HANDBOOK A6.3, A6.5

2. Determine the molecular formula of each compound from the data provided. T/I C
   (a) A hydrocarbon containing 85.6 % carbon and the remainder hydrogen has a molar mass of 42.09 g/mol. [ans: C_3H_6]
   (b) A compound is 43.6 % phosphorus and the remainder oxygen. Its molar mass is 283.88 g/mol. [ans: P_4O_{10}]
   (c) A compound that contains 64.3 % carbon, 7.2 % hydrogen, and the remainder oxygen has a molar mass of 168.21 g/mol. [ans: $C_9H_{12}O_3$]

6.9 Summary

- The molecular formula gives the actual composition of a compound.
- The molecular formula is a whole-number multiple of the empirical formula. The multiplier can be determined using the equation

$$x = \frac{\text{molar mass of compound}}{\text{molar mass of empirical formula}}$$

- The molecular formula can be calculated from either the empirical formula and molar mass or the percentage composition and molar mass.

6.9 Questions

1. "A compound's molar mass (based on its molecular formula) is always a whole-number multiple, x, of the molar mass of its empirical formula." To test this statement, select a compound from Table 1 on page 297 and compare the molar masses of its empirical and molecular formulas. T/I
2. Determine the molecular formula for compounds with the following empirical formulas and molar masses: T/I C
   (a) NO_2; 92.02 g/mol
   (b) CH_2; 84.18 g/mol
   (c) $C_2H_3O_3$; 225.15 g/mol
   (d) CFBrO; 253.82 g/mol
3. Under what condition is the molecular formula of a compound the same as its empirical formula? K/U
4. Explain how two compounds can have the same percentage composition but different molecular masses. K/U
5. Empirical or molecular formulas can be used to describe molecular compounds. Can these concepts be used to describe ionic concepts as well? Explain your answer. K/U
6. Determine the empirical and molecular formulas of caffeine. Its molar mass is 194.19 g/mol and its percentage composition is 49.48 % carbon, 5.15 % hydrogen, 28.87 % nitrogen, and 16.49 % oxygen. T/I C
7. A substance believed to be an anabolic steroid (used by some athletes to build muscle) is seized from a professional athlete and analyzed. Combustion analysis reveals that the compound is 80.0 % carbon, 9.41 % hydrogen, and 10.6 % oxygen. A mass spectrometer analysis shows that the compound has a molar mass of 300.48 g/mol. K/U T/I C
   (a) Determine the molecular formula of the compound.
   (b) The substance is believed to be dianabol, a common performance-enhancing steroid. Is this analysis data sufficient to accuse the athlete of using dianabol? Why or why not?
8. Carbohydrates are compounds of carbon, hydrogen, and oxygen in which the ratio of hydrogen atoms to oxygen atoms is 2:1. A certain carbohydrate is known to be 40.0 % carbon. Its molar mass is approximately 180 g/mol. Determine the carbohydrate's empirical formula and molecular formula. T/I C
9. The media often highlight stories of athletes taking performance-enhancing substances such as steroids (**Figure 3**). T/I A

Figure 3 What are the effects of taking steroids?

   (a) Research the similarities and differences between anabolic steroids and human steroids.
   (b) Research the downside of "bulking up."

GO TO NELSON SCIENCE

Investigation 6.6.1 OBSERVATIONAL STUDY

SKILLS MENU

- Questioning
- Researching
- Hypothesizing
- Predicting
- Planning
- Controlling Variables
- Performing
- Observing
- Analyzing
- Evaluating
- Communicating

Popping Percentage Composition

A kernel of popping corn consists mostly of starch and small quantities of other substances such as water. All of these substances are encased in a hard shell. When a kernel is heated, its interior temperature can reach 150 °C—well above the boiling point of water. The hard exterior of the kernel prevents much of the water from boiling and turning into steam. Water that remains a liquid above its boiling point is said to be "superheated." However, some of the water does vaporize. The molecules in the vapour push against the hard exterior of the kernel. Eventually, the pressure builds up so much that the kernel explodes or "pops." This allows the remaining water to vaporize. The expanding steam "fluffs" out the starches within the kernel into an enjoyable light snack: popcorn.

Have you ever wondered why some kernels never pop? Here's one possible explanation. Imagine trying to inflate a balloon with a hole in it. Air loss through the hole prevents the balloon from fully inflating. Similarly, a small crack in the kernel may allow water to escape before it can be superheated.

In this activity, you will determine the mass of water released by a sample of popcorn by comparing the mass of the popcorn before and after popping. This data will be used to determine the percentage by mass of water in popping corn.

Purpose

To determine the percentage (by mass) of water in popping corn

Equipment and Materials

SKILLS HANDBOOK A1.2

- hot-air popcorn maker (**Figure 1**)
- balance
- large clear plastic bag (to catch the popcorn)
- popping corn

Do not eat the popcorn. Some may have spilled on the countertops and become contaminated with chemicals. To unplug the popcorn maker, pull on the plug rather than the cord.

Procedure

SKILLS HANDBOOK A2.4, 3.2, A6

1. Plan a procedure to determine the percentage (by mass) of water in popping corn.
2. Develop a mathematical equation that you will use to calculate the percentage of water in popping corn.

Figure 1 A hot-air popcorn maker

Analyze and Evaluate

(a) Determine the percentage of water in popping corn. T/I

(b) Analyzing error is an important part of any investigation. Some inaccuracies result from human error—mistakes made by the experimenter. Experimental errors are inherent in the design of the investigation or in the tools or techniques used. Classify each of the following as either a human error or an experimental error. In each case, predict the effect that this error could have on the percentage of water calculation. T/I

(i) not all of the popping corn was massed

(ii) some kernels did not pop

(iii) a student had to hold the bag to prevent it from falling over as its mass was measured

(iv) the initial mass of the bag was not subtracted from the total mass of bag and popcorn

(v) the age of the popcorn is unknown

(c) List the sources of error in your investigation. Suggest how you could avoid or minimize each one. T/I

Apply and Extend

(d) Repeat the investigation using a bag of unflavoured microwave popcorn in a microwave oven. Compare the sources of error between the two procedures. T/I C

Investigation 6.6.2 CONTROLLED EXPERIMENT

Percentage Composition of Magnesium Oxide

SKILLS MENU

- Questioning
- Researching
- Hypothesizing
- Predicting
- Planning
- Controlling Variables
- Performing
- Observing
- Analyzing
- Evaluating
- Communicating

In this experiment you will use experimental data to determine the percentage composition of magnesium oxide. This compound forms when magnesium metal burns:

$$2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s) + \text{energy}$$

Chemists believe that the energy released by this reaction is sufficient to allow magnesium to also react with nitrogen in the air. This reaction produces magnesium nitride, $Mg_3N_2(s)$:

$$3\ Mg(s) + N_2(g) \rightarrow Mg_3N_2(s) + \text{energy}$$

What remains unclear is whether enough magnesium nitride forms to noticeably alter the percentage composition of magnesium oxide.

Testable Question

SKILLS HANDBOOK A2.2

What effect, if any, does the formation of magnesium nitride have on the percentage composition of magnesium oxide determined in this experiment?

Hypothesis

Predict what effect the formation of magnesium nitride has on the percentage composition of magnesium oxide in this experiment. Your hypothesis should include reasons.

Variables

Identify all major variables that will be manipulated, measured, and/or controlled in this investigation. What is the responding variable?

Experimental Design

You will burn two identical strips of magnesium in ceramic crucibles. The contents of one crucible will be analyzed as they are. In the other crucible, any magnesium nitride that may have formed will be chemically converted into magnesium oxide. You will then determine the mass of product in both crucibles and use this value to calculate the percentage composition of magnesium oxide. You will compare the two calculated values.

Equipment and Materials

SKILLS HANDBOOK A1.1, A1.2, A3.2

- chemical safety goggles
- lab apron
- ceramic crucible and lid
- balance
- steel wool or sandpaper
- ring clamp
- Bunsen burner clamped to a retort stand
- clay triangle
- tongs
- glass stirring rod
- graduated cylinder
- spark lighter
- two 8 cm strips of magnesium ribbon, Mg(s)
- dropper bottle containing distilled water

This activity involves open flames. Tie back long hair and secure loose clothing. Always be careful when handling potentially hot materials and equipment.

In the event of a fire involving magnesium, do not use a fire extinguisher containing water or carbon dioxide. Smother the fire with sand.

Magnesium produces a hot and very bright flame when it burns. Do not look directly at the flame because the intense light may cause eye damage.

Procedure

1. Put on your safety goggles and lab apron.

Part A: Not Correcting for Magnesium Nitride

2. Measure and record the mass of a clean, dry crucible and lid.
3. Polish the surface of the magnesium strips with steel wool or sandpaper.
4. Coil the magnesium loosely so that it fits in the lower half of the crucible.
5. Measure and record the mass of the magnesium, crucible, and lid.
6. Attach the ring clamp to the retort stand about 10 cm above the top of the Bunsen burner.
7. Place the crucible securely in the clay triangle resting on the ring clamp. Place the lid on the crucible so that it is slightly ajar. This is necessary to allow oxygen to enter the crucible (**Figure 1**).
8. Practice covering and uncovering the crucible using tongs.
9. Light the burner and heat the magnesium gently using a clean blue flame. You may have to raise the lid periodically, using tongs, to see if the reaction is proceeding.

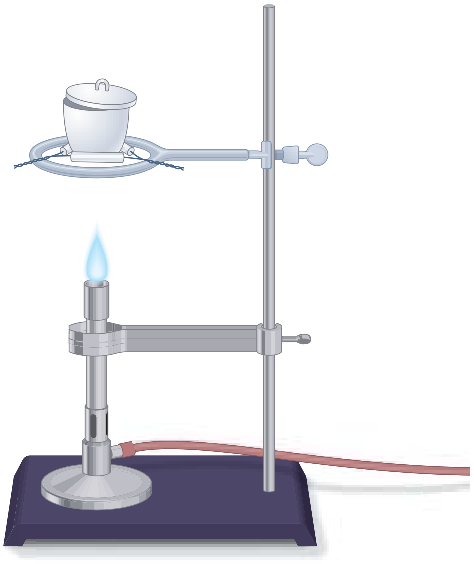


Figure 1 The apparatus for heating magnesium in a crucible

10. After about 2 min of gentle heating, increase the intensity of the flame. Heat strongly until the magnesium no longer flares.
11. Turn off the burner and allow the crucible and its contents to cool.
12. Determine and record the mass of the contents of the crucible.
13. Save the contents of the crucible for Part B.

Part B: Correcting for Magnesium Nitride

14. Using tongs, transfer the cooled crucible to the bench top. Crush the white ash in the crucible with the end of a stirring rod.
15. Add about 5 mL of distilled water to the crucible. Use some of the water to wash any residue on the rod into the crucible.
16. Heat the crucible and its contents gently for about 2 min to evaporate the water. Then heat strongly for 5 min to convert any magnesium nitride present to magnesium oxide.
17. Allow the crucible and its contents to cool.
18. Determine and record the mass of the contents of the crucible.
19. Discard the contents of the crucible as directed by your teacher.

Analyze and Evaluate

SKILLS HANDBOOK A6

(a) Identify the major variables that you measured and/or controlled in this investigation. Was there a causal relationship between any of the variables? Explain. T/I

(b) List your sources of error in this experiment. Identify one error that could significantly affect your results. What effect would it have? How could you correct for this error? T/I

(c) Calculate the theoretical percentage composition of magnesium oxide, MgO. T/I

(d) Determine the mass of oxygen that reacted with magnesium in Part A and in Part B. T/I

(e) Calculate the percentage composition of magnesium oxide using your experimental data from Part A. T/I

(f) Calculate the percentage composition of magnesium oxide using your experimental data from Part B. T/I

(g) Compare the three different values for percentage composition of magnesium oxide. Try to explain any differences. T/I

(h) Answer the Testable Question that was posed at the beginning of this investigation. Support your answer with evidence. T/I

Apply and Extend

(i) A student observes unreacted magnesium remaining in the crucible. What effect would this have on the percentage composition calculation? Why? T/I

(j) Suggest a modification of this experiment that would prevent magnesium from reacting with nitrogen. T/I

(k) Research and write the chemical reactions involved in converting magnesium nitride to magnesium oxide during Part A of this experiment. T/I C

GO TO NELSON SCIENCE

Investigation 6.9.1 OBSERVATIONAL STUDY

SKILLS MENU

• Questioning	• Planning	• Analyzing
• Researching	• Controlling Variables	• Evaluating
• Hypothesizing	• Performing	• Communicating
• Predicting	• Observing	

Determining the Formula of a Hydrate

Most hydrates are ionic compounds that contain water molecules within their crystalline structure. For example, magnesium sulfate heptahydrate, $MgSO_4{\cdot}7H_2O$, contains 7 water molecules per formula unit of magnesium sulfate. Heating some hydrates causes the water molecules within the crystal to evaporate out of the crystal. This is similar to steam leaving a kernel of popping corn. What remains is the anhydrous or "without water" form of the compound.

In this investigation you will remove the water from a hydrated form of copper(II) sulfate, $CuSO_4{\cdot}xH_2O$, in order to determine its percentage composition of water. You will also use this information to determine the chemical formula of the hydrate.

Purpose

- To determine the percentage composition of water in a hydrate of copper(II) sulfate
- To determine the chemical formula of $CuSO_4{\cdot}xH_2O$

Equipment and Materials

SKILLS HANDBOOK A1.1, A1.2

- chemical safety goggles
- lab apron
- heat-resistant safety gloves
- large test tube
- scoopula
- balance
- utility clamp
- Bunsen burner clamped to a retort stand
- spark lighter
- gauze pad
- 2.00 to 4.00 g of hydrated copper(II) sulfate, $CuSO_4{\cdot}xH_2O$
- 2 strips of cobalt chloride paper

This activity involves open flames. Tie back long hair and secure loose clothing. Always be careful when handling materials and equipment that could be hot.

Copper(II) sulfate is toxic and an irritant. Avoid skin and eye contact. In the case of contact, wash the affected area with lots of cool water and inform your teacher.

All the copper(II) sulfate used in this investigation must be collected and returned to your teacher at the end of the investigation.

Procedure

SKILLS HANDBOOK A2.4, A3.2

1. Put on your safety goggles and lab apron.
2. Record the mass of a large, heat-resistant test tube.
3. Your teacher will tell you what mass of $CuSO_4{\cdot}xH_2O$ your group will use. Add your assigned mass of hydrate to the test tube.
4. Attach a utility clamp to the test tube.
5. Light the Bunsen burner. Put on your safety gloves. Heat the test tube and its contents moderately by moving it from side to side for 1 min in a Bunsen burner flame (**Figure 1**). Do not localize the flame at any point on the test tube. Check for evidence of condensation forming near the mouth of the test tube.

Ensure that the mouth of the test tube is always pointing away from you and your classmates.

6. Test the condensate with a strip of cobalt chloride paper. Test a stream of tap water with the other strip.
7. Heat strongly for 2 to 3 min or until there is no further evidence of condensation. Be careful not to overheat the hydrate. If you do, it may decompose to form a black residue of copper(II) oxide.

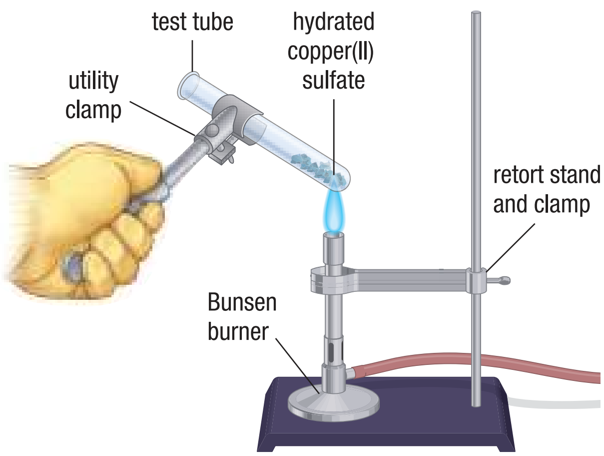


Figure 1 A test tube of hydrated copper(II) sulfate being heated in a Bunsen burner flame

8. Place the test tube on a gauze pad on the bench and allow it to cool.
9. Determine the mass of the contents of the test tube.
10. Return the anhydrous copper(II) sulfate to your teacher. Clean up as instructed.

Analyze and Evaluate

SKILLS HANDBOOK A6

(a) Are you confident that the investigation provided adequate data for analysis? Were there any significant sources of error? Describe any concerns and predict how they might affect your results. T/I

(b) What qualitative evidence suggested that the copper(II) sulfate was converted into its anhydrous form? T/I

(c) Determine the mass of water and mass of anhydrous copper(II) sulfate produced. T/I

(d) Determine the percentage of water (by mass) in your sample of hydrate. Post this value together with the mass of your hydrated copper(II) sulfate on the board. T/I

(e) Should your calculated value for the percentage of water be similar or different to those determined by other groups? Why? T/I

(f) Determine
  (i) the amount of water of hydration in your sample
  (ii) the amount of anhydrous copper(II) sulfate in your sample T/I

(g) Determine the value of *x* and write the complete chemical formula of the hydrate. T/I C

(h) Predict the effect of each of the following on the percentage by mass of water in the hydrate and the chemical formula of the hydrate: T/I A
  (i) not heating the hydrate enough
  (ii) overheating the hydrate and producing some copper(II) oxide

Apply and Extend

(i) Following Step 9, a student decided to continue heating the test tube. After cooling, the mass of the contents of the test tube was less than the mass observed in Step 9. No black residue was observed. Suggest an explanation to account for this observation. T/I

(j) Suggest a way to convert the anhydrous copper(II) sulfate back into its hydrated form. T/I

CAREER LINK

Your science teacher is an important part of your science education. How could become a science teacher? To find out, and to explore more about this career,

GO TO NELSON SCIENCE

CHAPTER 6 SUMMARY

Summary Questions

1. Create a study guide based on the points listed in the margin on page 258. For each point, create three or four sub-points that provide relevant examples, explanatory diagrams, or general equations.
2. Look back at the Starting Points questions on page 258. Answer these questions using what you have learned in this chapter.
3. Prepare a one-page summary of each of the types of calculations used in this chapter. Highlight the step in each calculation that you find most confusing. Write a one- or two-sentence summary of why you find the step confusing and how you have overcome this difficulty.
4. Make a concept map to summarize the use of the mole concept in this chapter.

Vocabulary

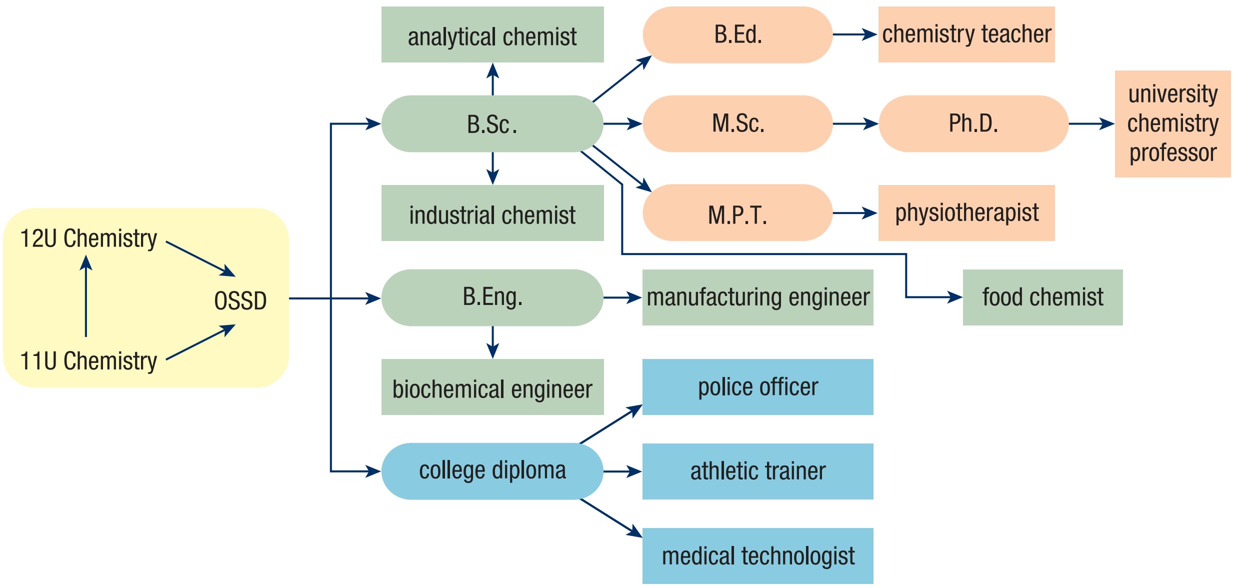
qualitative analysis (p. 260)
quantitative analysis (p. 261)
mole (p. 268)
Avogadro's constant (N_A) (p. 268)
amount (n) (p. 268)
molar mass (p. 271)
percentage composition (p. 284)
law of definite proportions (p. 285)
empirical formula (p. 290)
molecular formula (p. 290)

CAREER PATHWAYS

SKILLS HANDBOOK A7

Grade 11 Chemistry can lead to a wide range of careers. Some require a college diploma or a B.Sc. degree. Others require specialized or post-graduate degrees. Here are just a few pathways to careers mentioned in this chapter.

1. In this chapter, you have seen how a number of professions rely on quantitative aspects of chemistry and chemical reactions. Select two careers, related to quantities in chemical reactions, that you find interesting. Research the educational pathways that you would need to follow to pursue this career. What is involved in the required educational programs? Prepare a brief report of your findings.
2. For one of the two careers that you chose above, describe the career, main duties and responsibilities, working conditions, and setting. Also outline how the career benefits society and the environment.



GO TO NELSON SCIENCE

CHAPTER 6 SELF-QUIZ

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

For each question, select the best answer from the four alternatives.

1. Which of the following is an example of using quantitative analysis? (6.1) K/U
   (a) Red blood cells are present in a person's blood.
   (b) The number of red blood cells present is less than the normal amount.
   (c) In addition to other components, blood contains both red and white blood cells.
   (d) Blood looks red because it contains red blood cells.
2. Which tool might be used when performing qualitative analysis? (6.1) K/U
   (a) an eye
   (b) a ruler
   (c) a metre stick
   (d) a balance
3. If a dozen large eggs have a mass of 680 g, what is the approximate mass of one large egg? (6.3) T/I
   (a) 28 g
   (b) 57 g
   (c) 340 g
   (d) 680 g
4. Which is true about a mole of iron and a mole of sulfur? (6.3) K/U
   (a) The volume of sulfur is greater than the volume of iron because sulfur atoms are larger.
   (b) The volume of iron is greater than the volume of sulfur even though sulfur atoms are larger.
   (c) The volume of iron is greater than the volume of sulfur because iron atoms are larger.
   (d) The volumes are the same because a mole is always the same number of particles.
5. What is the molar mass of ammonia, NH_3? (6.4) K/U
   (a) 4 g/mol
   (b) 15 g/mol
   (c) 17 g/mol
   (d) 34 g/mol
6. How many atoms are present in 2.00 mol of nitrogen gas, N_2? (6.5) T/I
   (a) 3.01×10^{23} atoms
   (b) 6.02×10^{23} atoms
   (c) 1.20×10^{24} atoms
   (d) 2.40×10^{24} atoms
7. How does the percentage composition of a compound compare to the ratio of atoms of elements in a compound? (6.6) K/U
   (a) They are the same.
   (b) Percentage composition depends on both number of atoms and atomic mass; ratio of atoms depends on only the atomic mass.
   (c) Percentage composition depends on only the number of atoms; ratio of atoms depends on both number of atoms and atomic mass.
   (d) Percentage composition depends on both number of atoms and atomic mass; ratio of atoms depends on only the number of atoms.
8. What type of formula is CO_2, molecular or empirical? (6.7, 6.9) K/U
   (a) molecular only
   (b) empirical only
   (c) both molecular and empirical
   (d) neither molecular nor empirical
9. The empirical formula of a compound used to propel rockets is NO_2, and its molar mass is 92 g/mol. What is the molecular formula of the compound? (6.7, 6.9) T/I
   (a) NO (c) NO_4
   (b) NO_2 (d) N_2O_4

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

10. Quantitative analysis is used to measure the amount of a substance present. (6.1) K/U
11. Most of the salt in processed foods is added during the processing and is not originally in the food. (6.2) K/U
12. Both measurement and counting give an exact number of items. (6.3) K/U
13. The mole can be used to count only atoms, ions, molecules, and subatomic particles. (6.3) K/U
14. The unit used for molar mass is g/mol. (6.4) K/U
15. Molar mass mathematically connects the amount of a substance to its mass. (6.4) K/U
16. You can find the number of entities in a sample by dividing the amount by Avogadro's constant. (6.5) K/U
17. The percentage composition of MgO is 50 % Mg and 50 % O. (6.6) T/I
18. All compounds have different percentage compositions. (6.7) K/U
19. The mass spectrum of a compound differs from the mass spectrum of any other compound. (6.8) K/U
20. All compounds with the same molecular formula have the same properties. (6.9) K/U

To do an online self-quiz,
GO TO NELSON SCIENCE

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

Knowledge

For each question, select the best answer from the four alternatives.

1. Which of the following is an example of using quantitative analysis? (6.1) K/U
   (a) The wind is coming from the west and has a speed of 1.5 m/s.
   (b) Both cirrus and cumulonimbus clouds are present.
   (c) In summer, land masses near the Great Lakes often experience humid weather.
   (d) A thunderstorm is moving from west to east across Ontario.
2. How might a lab technician measure the amount of sodium chloride present in processed lunch meat? (6.1, 6.2) K/U
   (a) qualitative analysis only
   (b) quantitative analysis only
   (c) both qualitative analysis and quantitative analysis
   (d) neither qualitative analysis nor quantitative analysis
3. What is the base unit for the amount of a substance? (6.3) K/U
   (a) gram
   (b) molecule
   (c) Avogadro's constant
   (d) mole
4. Which isotope is the standard chosen to define the mole? (6.3) K/U
   (a) hydrogen-2
   (b) carbon-12
   (c) carbon-14
   (d) oxygen-16
5. What is the mass, in grams, of a mole of any element? (6.4) K/U
   (a) the atomic number of the element
   (b) the mass of one atom of the element
   (c) the atomic mass of the element
   (d) the mass of one molecule of the element
6. A student knows he has 2.5 mol of iron. He also knows Avogadro's constant. What must he do to these two numbers to find the number of iron atoms in the sample? (6.5) K/U
   (a) Add them.
   (b) Subtract them.
   (c) Multiply them.
   (d) Divide them.
7. What concept can be used to explain why each water molecule contains 2 hydrogen atoms and 1 oxygen atom? (6.6) K/U
   (a) the law of definite proportions
   (b) percentage composition
   (c) determination of molar mass
   (d) the comparison of an empirical formula and a molecular formula
8. $MgBr_2$ is an ionic compound. What type of chemical formula is $MgBr_2$? (6.7, 6.9) K/U
   (a) a molecular formula only
   (b) an empirical formula only
   (c) both a molecular and an empirical formula
   (d) neither a molecular nor an empirical formula
9. The formula for the compound hydrogen peroxide is H_2O_2. Which type of chemical formula is H_2O_2? (6.7, 6.9) K/U
   (a) a molecular formula only
   (b) an empirical formula only
   (c) both a molecular and an empirical formula
   (d) neither a molecular nor an empirical formula

Match each term on the left with the appropriate formula(s) on the right. You can use each term on the right more than once.

10. (a) molecular formula — (i) $MgCl_2$
    (b) empirical formula — (ii) $C_6H_{12}O_6$
    (c) formula unit — (iii) CH
    (iv) CH_4 (6.7, 6.9) K/U

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

11. Qualitative analysis is used to measure the amount of a substance present. (6.1) K/U
12. The presence of cocaine on banknotes can be determined by qualitative analysis. (6.1, 6.8) K/U
13. The mole is the base unit for the amount of something. (6.3) K/U
14. The value commonly used for Avogadro's constant is 6.02×10^{23}. (6.3) K/U
15. The unit used for molar mass is grams, g. (6.4) K/U
16. The number of entities in a sample can be found by multiplying the mass of the sample by Avogadro's constant. (6.5) K/U
17. Percentage composition is the percentage, by number, of each element in a compound. (6.6) K/U
18. A molecular formula is found by multiplying all the subscripts of an empirical formula by the same whole number. (6.7, 6.9) K/U

19. The formula $C_4H_8O_2$ is an empirical formula. (6.9) K/U

Write a short answer to each question.

20. Classify each of the following statements as an example of qualitative analysis or an example of quantitative analysis. (6.1) K/U
    (a) A leaf is green.
    (b) A board is 4.6 cm long.
21. Classify each of the following as a tool used in qualitative analysis or a tool used in quantitative analysis: (6.1) K/U
    (a) a graduated cylinder
    (b) your eyes
22. (a) How can you use the molar mass of iron and Avogadro's constant to find the mass of a single iron atom?
    (b) Use your answer to Part (a) to find the mass of a single iron atom. (6.3, 6.4) K/U T/I
23. (a) Explain how to multiply two numbers that are written in scientific notation.
    (b) Multiply 2.4×10^{14} and 4.6×10^{4}. (6.3) K/U
24. What is the result when 5.34×10^{5} is divided by 3.56×10^{7}? (6.3) K/U

Understanding

25. Bromine and mercury are the only two elements that are liquid at room temperature. One mole of mercury is in a sealed container, and one mole of bromine is in an identical sealed container. (6.3) K/U
    (a) A balance is used to determine the mass of each liquid and its container. The mass of the container is subtracted from the mass of the liquid and the container to find the mass of the liquid. How do the masses of the mercury and the bromine compare?
    (b) The volumes of the liquids are compared by measuring the depth of the liquids in the containers. How do the volumes of the two liquids compare? Explain your answer.
26. (a) Explain why the mass of 1 mol of zinc differs from the mass of 1 mol of sulfur.
    (b) How do the masses compare? (6.3) K/U
27. What is the molar mass of aluminum? Include correct units. (6.4) K/U
28. What is the molar mass of each of the following? (6.4) K/U T/I
    (a) oxygen atom (O)
    (b) oxygen gas (O_2)
    (c) ozone (O_3)
29. (a) How do you find the molar mass of a hydrate?
    (b) Find the molar mass of cobalt chloride hexahydrate, $CoCl_2 \cdot 6H_2O$. (6.4) T/I
30. Explain how finding the molar mass of a molecular compound differs from finding the molar mass of an ionic compound. Use examples in your explanation. (6.4) K/U C
31. (a) What is the molar mass of potassium permanganate, $KMnO_4$?
    (b) How many moles of potassium permanganate are in 258 g of the compound? (6.4) K/U T/I
32. How many grams of iron are in 2.7 mol of iron? (6.4) T/I
33. Calculate the mass of copper(II) nitrate, $Cu(NO_3)_2$, in 4.5 mol of the compound. (6.4) T/I
34. How many atoms are in 0.68 mol of silicon? (6.5) T/I
35. (a) Find the number of molecules in 1.4 mol of ethane, C_2H_6.
    (b) How many atoms are in the same amount of ethane? (6.5) T/I
36. The mass of a sample of pure water is 4.56 g. What is the amount of water in the sample? (6.5) T/I
37. What is percentage composition? Provide an example in your explanation. (6.6) K/U
38. A sample of methane, CH_4, contains 7.50 g of carbon. The total mass of the sample is 10.0 g. Carbon atoms make up what percentage by mass of methane? (6.6) T/I
39. What is the percentage composition of magnesium chloride, $MgCl_2$? (6.6) T/I
40. In your own words, describe the law of definite proportions. (6.6) K/U
41. A sample of potassium nitrate, KNO_3, contains 78.2 g of potassium, 28.0 g of nitrogen, and 96.0 g of oxygen. Find the percentage composition of potassium nitrate. (6.6) T/I
42. A 24.9 g sample of aluminum bromide, $AlBr_3$, contains 2.52 g of aluminum.
    (a) What mass of bromine is in the sample?
    (b) Find the percentage composition of aluminum bromide. (6.6) T/I
43. What is the percentage composition of hydrogen sulfide, H_2S? (6.6) T/I
44. How does finding the percentage composition of a compound differ from finding the percentage composition of a mixture? Use examples in your explanation. (6.6) K/U
45. Could the empirical formula of potassium fluoride be K_2F_2? Explain. (6.7) K/U
46. Find the empirical formula of a compound that has a percentage composition of 23.6 % potassium and 76.4 % iodine. (6.7) T/I

47. A compound contains 74.2 % sodium. The rest of the compound is oxygen. (6.7) T/I
    (a) What percentage of the compound is oxygen?
    (b) What is the empirical formula of the compound?
48. A compound contains 20.2 % phosphorus and 10.4 % oxygen. The rest of the compound is chlorine. What is the empirical formula of this compound? (6.7) T/I
49. What is the molecular formula of a compound that has a molar mass of 92.0 g/mol and an empirical formula of NO_2? (6.9) T/I
50. A compound has a molar mass of 208.5 g/mol. Its percentage composition is 14.9 % phosphorus and 85.1 % chlorine. What is the molecular formula of the compound? (6.9) T/I

Analysis and Application

51. A student is determining the effect of differing quantities of fertilizer on the growth of bean plants. (6.1) K/U T/I A
    (a) Describe one way the student can use qualitative analysis in his experiment.
    (b) Explain how quantitative analysis is important to the results of this experiment.
52. A student knows that some acids react with metals, releasing bubbles of hydrogen gas. She is observing what happens when magnesium and copper are placed into hydrochloric acid. She analyzes the results using the amount of hydrogen gas released by the reaction between the acid and each metal. (6.1) K/U T/I C A
    (a) How might she describe the results qualitatively?
    (b) How might she describe the results quantitatively?
53. Classify each of the following tools as being most useful in quantitative analysis or qualitative analysis: (6.1) K/U A
    (a) a metric ruler
    (b) a balance
    (c) an eye
    (d) the skin on a finger
    (e) a measuring cup
54. The average adult human circulatory system contains approximately 3.00×10^{13} red blood cells. How many moles of red blood cells does this represent? (6.3) T/I A
55. Measurements made in space are usually so large that they are expressed in scientific notation. For example, the Large Magellanic Cloud is a galaxy that is located 1.63×10^5 light years from Earth. One light year is approximately 9.5×10^{12} km, which is the distance light travels in a year. Determine the approximate distance of the Large Magellanic Cloud from Earth, in kilometres. (6.3) T/I A
56. Some systems that purify water in pools and spas use chlorine, and some use bromine. One mole of chlorine reacts in the same way as one mole of bromine. Assume you need to add 1 mol of chlorine to one pool and 1 mol of bromine to another pool. How would the mass of chlorine added compare to the mass of bromine added? Explain. (6.3) T/I C A
57. A problem in a homework assignment refers to the molar mass of bromine. One student says that the molar mass of bromine is 79.9 g/mol. Another student says that the molar mass of bromine is 159.8 g/mol. (6.4) K/U T/I C A
    (a) If no other information is known, how might either student be correct?
    (b) What else do you need to know to determine which of the two molar masses is correct for bromine?
58. Pentane has the chemical formula C_5H_{12} and is used in some refrigerants. Methylbutane also has the chemical formula C_5H_{12}, and it is used along with liquid nitrogen in situations requiring extremely low temperatures. (6.4, 6.9) K/U A
    (a) Explain how one chemical formula can represent two or more compounds with different properties.
    (b) How do the molar masses of these two compounds compare?
59. Magnesium sulfate, $MgSO_4$, forms several different hydrates. One of the hydrates has a molar mass of 246.4 g/mol. Another hydrate has a molar mass of 210.4 g/mol. What are the chemical formulas for the two hydrates? (6.4) K/U T/I
60. Aluminum oxide, Al_2O_3, is quite common in Earth's crust. If the compound is accompanied by very small concentrations of other elements, the result might be a colourful gemstone. For example, a ruby is aluminum oxide with a small quantity of chromium in it. A sapphire is aluminum oxide that contains very small quantities of iron or titanium. (6.4) T/I A
    (a) What mass of aluminum oxide is in 2.70 mol of the compound?
    (b) What amount of aluminum oxide is in 67.5 g of the compound?
61. Ethyne (acetylene), C_2H_2, produces a high-temperature flame when it burns. As a result, it is often used in welding. (6.4) K/U T/I A
    (a) How many moles of ethyne are in a tank that contains 526 g of the gas?
    (b) If a welding tank is filled with 25.7 mol of ethyne, what mass of the gas does the tank contain?

62. Nitric acid, HNO_3, is used to produce many products, such as fertilizers and explosives. It also dissolves certain metals, such as gold, that do not dissolve in other chemicals. How many atoms are in 1.25 mol of nitric acid? (6.5) T/I A

63. Some air pollutants oxidize in air to form other pollutants. (6.5) T/I A
   (a) The gas sulfur dioxide, SO_2, oxidizes to form another gas, sulfur trioxide, SO_3. How many oxygen atoms are present in 3.20 g of sulfur trioxide?
   (b) At high temperatures, nitrogen gas oxidizes to form nitrogen monoxide, NO. This compound further oxidizes to form nitrogen dioxide, NO_2. How many oxygen atoms are present in 2.50 g of NO_2?

64. Acid rain remains a problem in Canada. It damages vegetation, pollutes water, and reacts with stone and other materials. However, government regulations and concerned individuals and industries have reduced the occurrence of acid rain in recent years. One way acid rain forms is by the dissolving of the gas sulfur trioxide, SO_3, in rainwater, forming sulfuric acid, H_2SO_4. (6.5) K/U T/I A
   (a) How many molecules of sulfur trioxide are in 5.2 g of the gas?
   (b) What is the total number of atoms in 254 g of sulfuric acid?

65. Iron pyrite, FeS_2, is sometimes called “fool’s gold” because it looks like gold. A 24.3 g sample of pyrite contains 11.3 g of iron. What percentage of pyrite, by mass, is made up of iron? (6.6) T/I A

66. Ethanol, C_2H_5OH, forms when grains such as corn or wheat undergo the process of fermentation. As a result, ethanol is sometimes called grain alcohol. Ethanol is sometimes added to gasoline as a fuel. Adding it to gasoline helps conserve petroleum, which is a non-renewable resource. What is the percentage composition of ethanol? (6.6) K/U T/I A

67. Magnesium hydroxide, $Mg(OH)_2$, is a common antacid and is also used in deodorants. The percentage of magnesium in one sample of magnesium hydroxide is 41.7 %. How does the law of definite proportions help you determine the percentage of magnesium in a different sample of the same compound? (6.6) K/U T/I

68. One possible product of the reaction of iron with oxygen is iron(III) oxide, Fe_2O_3. This reaction is used in some heat packs because it releases thermal energy. Suppose the iron(III) oxide produced in one heat pack contains 5.58 g of iron and 2.40 g of oxygen. What is the percentage composition of Fe_2O_3? (6.6) T/I A

69. Marble and limestone are common building materials. The major component of marble and limestone is calcium carbonate, $CaCO_3$. One slab of marble that a sculptor wants to use to make a sculpture has a mass of 245 kg. This piece of marble contains 98.1 kg of calcium and 29.4 kg of carbon. What is the percentage composition of calcium carbonate? (6.6) T/I A

70. Tetraphosphorus trisulfide, P_4S_3, requires little energy to burn. As a result, it is one of the ingredients commonly used in match heads so that they burn readily. What is the percentage composition of tetraphosphorus trisulfide? (6.6) K/U T/I A

71. List two examples of each of the following:
   (a) a molecular formula that is also an empirical formula
   (b) a molecular formula that is not an empirical formula (6.7, 6.9) K/U

72. In determining the empirical formula of a compound from its percentage composition, Rosa chose to use a 25 g sample of the compound. Mark told her that she had to use a 100.0 g sample. (6.7) K/U T/I C
   (a) Was Mark correct?
   (b) Explain your answer to Part (a).

73. Magnesium sulfide is formed when magnesium powder reacts with the sulfur in molten steel to purify the steel. Magnesium sulfide contains 43.1 % magnesium, and the remainder of the compound is sulfur. (6.7) K/U T/I A
   (a) What percentage of magnesium sulfide is sulfur?
   (b) What is the empirical formula for magnesium sulfide?

74. Calcium chloride is sometimes used on icy roads instead of sodium chloride. Calcium chloride melts snow and ice better than an equal amount of sodium chloride does. Calcium chloride is 36.1 % calcium. The rest of the compound is chlorine. What is the empirical formula for calcium chloride? (6.7) K/U T/I A

75. Photosynthesis is the process by which green plants produce glucose and oxygen from water and carbon dioxide in the presence of sunlight and chlorophyll. Glucose contains 40.0 % carbon and 6.67 % hydrogen, by mass. The remainder of the compound is oxygen. (6.7) K/U T/I A
   (a) What percentage of glucose is oxygen?
   (b) What is the empirical formula of glucose?

76. Oxalic acid is a weak acid sometimes used to clean or bleach materials. It is especially effective in removing rust. Oxalic acid has a molar mass of 90.0 g/mol. It contains 2.20 % hydrogen and 26.7 % carbon. The rest of the compound is made up of oxygen. (6.7, 6.9) K/U T/I A
   (a) What percentage of oxalic acid is made up of oxygen?
   (b) What is the empirical formula for oxalic acid?
   (c) What is its molecular formula?
77. Phosphorus and oxygen form several different compounds when they react with each other. One of these compounds is a strong dehydrating agent because it removes water from other materials, forming phosphoric acid. This compound has a molar mass of 285 g/mol and an empirical formula of P_2O_5. What is the molecular formula for this compound? (6.9) T/I A

Evaluation

78. With a partner, debate which you think is more important to medical research, qualitative analysis or quantitative analysis. Provide specific justification for your side, accompanied by examples. Recognize that both types of analysis have value. (6.1) T/I C A
79. Ammonia is important in making many products, including fertilizers and cleaning products. Assume you work for a manufacturer that makes ammonia from hydrogen and nitrogen gases. You are in charge of ordering all the raw materials for the manufacturing process. (6.3) K/U T/I C A
   (a) Why do you need to know how much hydrogen reacts with each mole of nitrogen to make ammonia?
   (b) 1 mol of nitrogen reacts with 3 mol of hydrogen to make ammonia. What amount of nitrogen would you order to react with 270 mol of hydrogen?
   (c) How many kilograms of hydrogen do you need to order if you are ordering 840 kg of nitrogen, assuming that none of either compound is left over?
80. Recall the procedure for finding the number of entities in a sample if the amount is known. (6.5) K/U T/I C
   (a) What additional information do you need to know to use this procedure to find the amount of a substance if the number of entities is known?
   (b) Briefly explain how to find the number of moles of carbon dioxide when 1.9×10^{24} molecules of carbon dioxide are present.
   (c) How would this explanation change if you knew the number of oxygen atoms present instead of the number of carbon dioxide molecules?
81. A chemical formula can be a molecular formula or an empirical formula, or it can show a formula unit. Which two of these types of formulas cannot both apply to the same compound? Explain, using examples. (6.7, 6.9) K/U T/I C A
82. A student made the statement that any chemical formula for an ionic compound must be an empirical formula and cannot be a molecular formula. Evaluate that statement. Use at least two examples in your evaluation. (6.7, 6.9) K/U T/I C A

Reflect on Your Learning

83. Choose a product that you use every day. K/U T/I C A
   (a) How might qualitative analysis be used in the manufacturing of this product?
   (b) How might quantitative analysis be helpful in its production?
84. Recall that measurements of length can have different units, depending on what is measured. For example, the height of a door is probably written using units of metres, road distances are likely to be expressed as kilometres, and the length of an insect might have units of millimetres or centimetres. Units used to describe the amount of something can also use prefixes to make the unit best express the number of moles present. K/U C A
   (a) Describe a situation in which an amount would likely be expressed in millimoles (mmol).
   (b) When might it be most convenient to express an amount in kilomoles (kmol)?
85. Visual representations can be helpful in performing calculations that require more than one step. K/U T/I C A
   (a) Draw a concept map that shows the sequence of steps in finding the number of entities in a sample of a substance if you know the mass of the substance.
   (b) Draw another concept map that you could use to find a molecular formula if you know molar mass and percentage composition.
86. (a) A ream contains 500 pieces of paper. Suppose you are asked to find the number of reams of paper when you know the number of sheets of paper present. How is this problem similar to finding the number of moles present when you know the number of entities present and Avogadro's constant?

(b) A gross contains 144 items. Suppose you are asked to find the number of pencils present in 4.5 gross of pencils. How is this problem similar to finding the number of entities present if you know the number of moles and Avogadro's constant? K/U T/I C A

Research

GO TO NELSON SCIENCE

87. Glucose is the type of sugar found in your blood. Fructose is a sugar found in fruit and certain other foods. Use the molecular and structural formulas of these sugars to show how they are similar and how they are different. T/I C A

88. The compound ethene is commonly known as ethylene. Ethylene is the organic compound produced in the greatest quantity in Canada as well as worldwide. T/I C A
    (a) Find the molecular and empirical formulas for ethylene.
    (b) Name at least three other compounds that have the same empirical formula as ethylene.
    (c) How do the uses of these compounds compare to the uses of ethylene?

89. The procedures used to produce ethene (ethylene) in Canada vary from year to year and place to place. T/I C A
    (a) Investigate how ethene is produced.
    (b) The mass of ethene produced is usually measured in tonnes. Convert 200.0 t of ethene to kilograms.
    (c) What amount of ethene is 200.0 t? Why might it be more convenient to report the amount of ethene produced in units of kilomoles instead of moles?

90. The term *empirical* is used in instances other than in *empirical formula.* T/I A
    (a) What is meant by the term *empirical?*
    (b) How does this definition help explain what an empirical formula is?

91. Most of the dietary salt taken into the human body comes from food that is processed. T/I A
    (a) Which two elements are in salt? Which of these elements is listed on a food label to reflect the amount of salt in the product?
    (b) Use the Internet and other resources to find information on the amount of sodium present in a food that can be purchased unprocessed (fresh or frozen) or processed (dried, canned, or boxed). Compare the amount of sodium present in both forms of the food.
    (c) How does the recommended daily intake of sodium for a person change as the person ages?
    (d) What is meant by a low-sodium food and a no-sodium food?

92. Currency contaminated with drugs and diseases has been found worldwide.
    (a) Use the Internet and other resources to research what potential solution to the problem was developed by Australia in the late 1980s.
    (b) Examine a mass spectrum of heroin. How does heroin's mass spectrum differ from the mass spectrum of cocaine? Using the mass spectrum, find the molecular mass of heroin. T/I A

CHAPTER

7 Stoichiometry in Chemical Reactions

KEY CONCEPTS

After completing this chapter you will be able to

- analyze processes in the home, the workplace, and the environment that involve the use of chemical quantities and calculations
- apply the concepts of molar mass and mole ratios in chemical reactions
- predict the amount or the mass of product expected in a chemical reaction
- identify the limiting reagent and excess reagent in a chemical reaction and predict the amount or the mass of product expected
- explain how the outcome of a chemical reaction can be affected by manipulating the limiting and excess reagents
- analyze quantitative data from an experiment and determine percentage yield to compare the actual and theoretical yields
- analyze industrial chemical reactions that can affect the health and safety of local populations and environments

How Do the Quantities of Reactants Affect the Product?

Margarine is an early example of a processed food product. Margarine was invented by a French chemist in 1869 as an inexpensive alternative to butter for the French army. The original raw material for making margarine was animal fat. However, due to fat shortages in the first half of the twentieth century, margarine manufacturers began using vegetable oils.

Vegetable oils are generally liquid fats. Their molecules have long hydrocarbon chains. These chains tend to include more carbon–carbon double bonds than are present in solid fat molecules. These double bonds affect the shape of the oil molecules, preventing them from aligning neatly. The molecules in animal fats do not have so many double bonds. They can line up, allowing more intermolecular forces of attraction between molecules to form. The attractions are the reason why animal fats are usually solid at room temperature. Animal fats are "saturated" because their molecules contain no double bonds. Molecules with one or more double bonds are unsaturated.

The type of fats that we consume can affect our health. A diet rich in saturated fats increases the body's concentration of the type of cholesterol in blood associated with heart disease. Conversely, eating certain unsaturated fats and oils tends to lower cholesterol concentrations.

The food industry converts liquid corn oil into a solid by saturating it: filling its double bonds with hydrogen atoms. This is done by bubbling hydrogen gas into corn oil in the presence of a nickel catalyst. However, the stiffness of the final product varies depending on the quantity of hydrogen used. If plenty of hydrogen gas is used, the oil becomes completely saturated and a hard, off-white solid forms. To make margarine, food technicians limit the degree of saturation by reducing the quantity of hydrogen used. This makes the margarine more spreadable. The result is a partially hydrogenated margarine: some of the molecules are saturated with hydrogen and others are unsaturated. Yellow food colouring, salt, and other ingredients are then added to make the margarine look and taste more like butter.

The production of margarine illustrates how the outcome of a chemical reaction can be controlled by adjusting the quantities of reactants used.

STARTING **POINTS**

Answer the following questions using your current knowledge. You will have a chance to revisit these questions later, applying concepts and skills from the chapter.

1. Why is it important to follow recipes accurately?
2. How can you predict the amount of product expected in a chemical reaction?
3. What factors control the amount of product in a chemical reaction?
4. How can the understanding of quantities of reactants in a chemical reaction benefit a company that manufactures chemical products?

Mini Investigation

Precipitating Ratios

Skills: Performing, Observing, Analyzing, Communicating   SKILLS HANDBOOK A1, A2.4

In this investigation you will mix different quantities of reactants to see which combination gives the most product.

Equipment and Materials: chemical safety goggles; lab apron; gloves; test-tube rack; 6 numbered test tubes; 10 mL graduated cylinder; dilute (0.1 mol/L) solutions of zinc chloride, $ZnCl_2(aq)$, sodium carbonate, $Na_2CO_3(aq)$, silver nitrate, $AgNO_3(aq)$, and sodium phosphate, $Na_3PO_4(aq)$

Both sodium carbonate and sodium phosphate are irritants. Avoid skin contact. If you spill some on your skin, wash the affected area with lots of cool water and inform your teacher.

Silver nitrate stains the skin. If you spill silver nitrate on your skin, wash the affected area with lots of cool water and inform your teacher.

1. Put on your chemical safety goggles, lab apron, and gloves.
2. Place six numbered test tubes in the test-tube rack.
3. Add solutions to three test tubes according to **Table 1.**
4. While waiting for the precipitates to settle, prepare the remaining test tubes according to **Table 2.**
5. Allow the precipitates to settle for about 10 min. Record your observations.

Table 1 Contents of the First Three Test Tubes

Test tube	Volume of zinc chloride solution, $ZnCl_2(aq)$ (mL)	Volume of sodium carbonate solution, $Na_2CO_3(aq)$ (mL)
# 1	1	3
# 2	2	2
# 3	3	1

Table 2 Contents of the Second Three Test Tubes

Test tube	Volume of silver nitrate solution, $AgNO_3(aq)$ (mL)	Volume of sodium phosphate solution, $Na_3PO_4(aq)$ (mL)
# 4	1	3
# 5	2	2
# 6	3	1

A. Write chemical equations for both reactions. K/U C

B. Which test tube in each set contained the most precipitate? T/I

C. Explain your observation in B by relating the quantities of reactants to each balanced chemical equation. T/I

7.1 Mole Ratios in Chemical Equations

Figure 1 The combustion of hydrogen fuel was used to launch the space shuttle into orbit. Aboard the shuttle, hydrogen fuel cells used the same chemical reaction to provide astronauts with electricity and water.

The reaction of hydrogen and oxygen has been an important energy source for the space program (**Figure 1**). Hydrogen combines with oxygen in a synthesis reaction to form water and a great deal of energy:

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g) + \text{energy}$$

As you discovered in Investigation 4.4.1 Hydrogen Blast-Off in Unit 3, this reaction releases the most energy per mole of reactants when the volumes of hydrogen and oxygen are in a 2:1 ratio.

Ratios in Chemical Equations

A chemical equation is similar to a recipe, with the chemical formulas indicating the "ingredients" of the reaction. The coefficients in the equation give the ratio of one chemical to another chemical used in the reaction. However, these are ratios of numbers of entities rather than masses. For example, the equation

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g) + \text{energy}$$

states that hydrogen molecules and oxygen molecules combine in a 2:1 ratio. This implies that 2 molecules of hydrogen are required to completely react with 1 molecule of oxygen. But what if more hydrogen is present? The amount of oxygen required to use up this extra hydrogen is determined using the 2:1 ratio. The amount of water produced is also determined by this ratio (**Figure 2**).

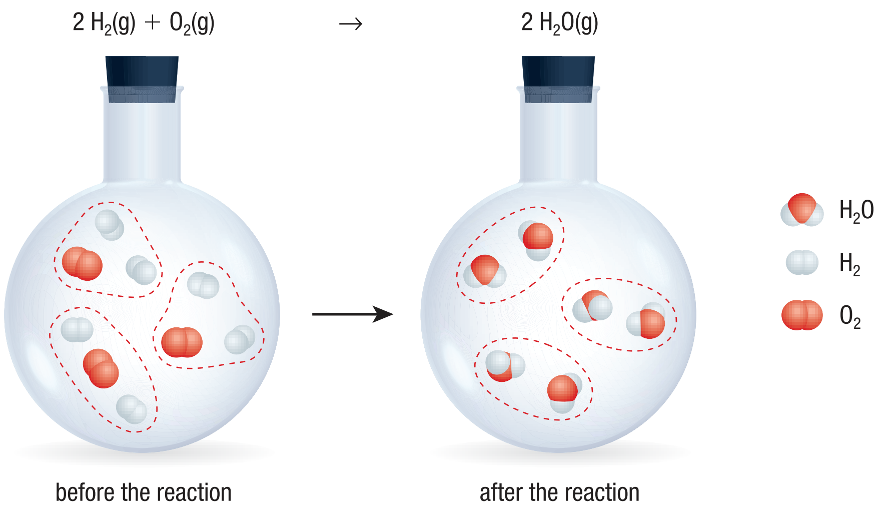


Figure 2 The synthesis of water from its elements. The coefficients in the chemical equation for the reaction give the ratio in which hydrogen and oxygen combine. In this case, 6 hydrogen molecules react with 3 oxygen molecules to form 6 water molecules.

CAREER LINK

Chemical process engineers plan which reactants should be mixed, under which conditions, to give the optimum yield and produce the smallest quantity of unwanted by-products. To find out about this career,

GO TO NELSON SCIENCE

Working with ratios is an essential skill in chemistry. The following analogy illustrates the reasoning used when working with ratios in chemical equations.

A Ratio Analogy

S'mores are sweet snacks that are fun to make over a campfire. They probably got their name from people asking, "Can I have some more?" Each one consists of two graham crackers sandwiching four chocolate chips and one marshmallow.

To make this recipe analogous to a chemical reaction, we will give each ingredient and product a chemical symbol:

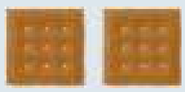
- graham cracker, Gc
- chocolate chip, Cc
- marshmallow, Mm
- s'more, Sm

Therefore, the "equation" that describes the synthesis of a s'more is

2 Gc + 4 Cc + Mm $\rightarrow$ Sm

According to the equation, the "reactants" combine in a 2:4:1 ratio to produce 1 entity of the product. You can make whatever quantity of s'mores you wish, provided the quantities of reactants are in this ratio (**Table 1**). For example, four times as many chocolate chips are required as marshmallows. Similarly, twice as many crackers are required as marshmallows. Any ingredient that is in excess of this ratio is left over after all the s'mores are made.

Table 1 Quantities in the Synthesis Reaction of S'mores

2 Gc +	4 Cc +	Mm $\rightarrow$	Sm
Number of Gc entities 	**Number of Cc entities** 	**Number of Mm entities**	**Number of Sm entities** 
2	4	1	1
4	8	2	2
50	100	25	25
2000	4000	1000	1000
$2 \times 6.02 \times 10^{23} = 2$ mol	$4 \times 6.02 \times 10^{23} = 4$ mol	$1 \times 6.02 \times 10^{23} = 1$ mol	$1 \times 6.02 \times 10^{23} = 1$ mol

Note that, in our s'more synthesis analogy, the sum of the masses of "reactants" must equal the total mass of the "products," since all the ingredients were used to make s'mores. In other words, this analogy illustrates the law of conservation of mass. Notice, however, that the total number of entities on the left side is not the same as the number of entities on the right side.

The Synthesis of Water

We can apply our s'mores reasoning to the water synthesis reaction (**Table 2**).

Table 2 Quantities in the Synthesis Reaction of Water

$2\ H_2(g)$ +	$O_2(g)$ $\rightarrow$	$2\ H_2O(g)$
Number of H_2 molecules	**Number of O_2 molecules**	**Number of H_2O molecules**
2	1	2
4	2	4
50	25	50
2000	1000	2000
$2 \times 6.02 \times 10^{23} = 2$ mol	$1 \times 6.02 \times 10^{23} = 1$ mol	$2 \times 6.02 \times 10^{23} = 2$ mol
10 mol	5 mol	10 mol

Note that each set of quantities is in the ratio 2:1:2, as given by the coefficients in the chemical equation. In fact, each set is a whole-number multiple of the coefficients 2, 1, and 2. This ratio is valid whether you are considering a few or a huge number of entities.

mole ratio the ratio of the amounts of the entities in a chemical reaction

The ratio of the amount (in moles) of one chemical to another in a chemical equation is called the **mole ratio**. The mole ratio can describe the ratio of reactants required to react. It can also describe the amount of product expected from a given amount of reactants. The mole ratio is related to the coefficients in the chemical equation.

Mini Investigation

One Plus One Does Not Always Equal Two

Skills: Predicting, Performing, Observing, Analyzing, Communicating — SKILLS HANDBOOK A2.4

In this investigation, you will use a molecular model kit to help visualize the proportions of reactants and products in a chemical reaction.

Equipment and Materials: molecular model kit

1. Build models of 1 molecule of hydrogen, H_2, and 1 molecule of fluorine, F_2. Predict the product(s) of the synthesis reaction involving these molecules. Use your models to evaluate your prediction. Record your observations.
2. Build models of 2 molecules of water, H_2O. Predict the products of the decomposition reaction involving these molecules. Use your molecules to evaluate your prediction.
3. Build models of 1 ethyne molecule, C_2H_2, and 2 fluorine molecules, F_2. Predict the molecular formula of the product of the synthesis reaction involving these molecules.

A. Write a balanced chemical equation for each reaction. K/U C

B. For each reaction, compare the total number of molecules on either side of the arrow. T/I

C. Do these reactions illustrate the law of conservation of mass? Explain your answer. K/U T/I

Determining Mole Ratios in Chemical Equations

You have seen how the amounts of reactants and products in a chemical reaction are multiples of the coefficients in the balanced chemical equation. Whole-number multiples are easy to predict, as Tables 1 and 2 illustrate, but what about fractional multiples?

Tutorial 1 Using Mole Ratios to Determine Amounts

In this Tutorial, you will learn a general process you can use to predict the amount of one chemical from a given amount of another.

Sample Problem 1: Determining Amounts Using Mole Ratios

What amount of hydrogen is produced when 5.2 mol of ammonia decomposes (**Figure 3**)?

Figure 3 Ammonia fertilizes the soil by adding nitrogen—an important plant nutrient.

Given: amount of ammonia, $n_{NH_3} = 5.2$ mol

Required: amount of hydrogen, n_{H_2}

Solution:

Step 1. Write a balanced equation for the reaction, listing the given value(s) and required value(s) below the substances being considered in the problem. Use the symbol of the required value since its value is unknown.

$$\underset{5.2\text{ mol}}{2\,NH_3(g)} \rightarrow N_2(g) + \underset{n_{H_2}}{3\,H_2(g)}$$

Step 2. Convert amount of the given substance to amount of the required substance. To do this, multiply the amount of the given substance by a suitable conversion factor derived from the mole ratio in the balanced equation.

The mole ratio between NH_3 and H_2 may be expressed as

$$\frac{2\ \text{mol}_{NH_3}}{3\ \text{mol}_{H_2}} \quad \text{or} \quad \frac{3\ \text{mol}_{H_2}}{2\ \text{mol}_{NH_3}}$$

Since we want to convert amount of ammonia to amount of hydrogen, we use the factor $\frac{3\ \text{mol}_{H_2}}{2\ \text{mol}_{NH_3}}$ as follows:

$$n_{H_2} = 5.2\ \cancel{\text{mol}_{NH_3}} \times \frac{3\ \text{mol}_{H_2}}{2\ \cancel{\text{mol}_{NH_3}}}$$

$$n_{H_2} = 7.8\ \text{mol}_{H_2}$$

Statement: The decomposition of 5.2 mol of ammonia produces 7.8 mol of hydrogen.

Sample Problem 2: Determining Amounts Using Mole Ratios

A freshly exposed surface of aluminum quickly reacts with oxygen to form a layer of aluminum oxide. What amount of oxygen is required to react completely with 6.4×10^{-2} mol of aluminum?

Given: $n_{Al} = 6.4 \times 10^{-2}$ mol

Required: amount of oxygen, n_{O_2}

Solution:

Step 1. Write a balanced equation for the reaction, listing the given value(s) and required value(s).

$$\begin{array}{lll} 4\ Al(s) & + & 3\ O_2(g) \rightarrow 2\ Al_2O_3(s) \\ 6.4 \times 10^{-2}\ \text{mol} & & n_{O_2} \end{array}$$

Step 2. Convert amount of the given substance to amount of the required substance.

$$n_{O_2} = 6.4 \times 10^{-2}\ \cancel{\text{mol}_{Al}} \times \frac{3\ \text{mol}_{O_2}}{4\ \cancel{\text{mol}_{Al}}}$$

$$n_{O_2} = 4.8 \times 10^{-2}\ \text{mol}$$

Statement: 4.8×10^{-2} mol of oxygen is required to react completely with 6.4×10^{-2} mol of aluminum.

Practice

SKILLS HANDBOOK **A6.5**

1. Phosphorus, P_4, burns in oxygen to produce diphosphorus pentoxide, P_2O_5:

   $P_4(s) + 5\ O_2(g) \rightarrow 2\ P_2O_5(g)$

   What amount of oxygen is required to produce 0.25 mol of the product? T/I

   [ans: 0.63 mol]

2. Iron can be produced from the reaction of iron(III) oxide, found in iron ore, with carbon monoxide, CO. The other product is carbon dioxide:

   $Fe_2O_3(s) + 3\ CO(g) \rightarrow 3\ CO_2(g) + 2\ Fe(s)$

   What amount of carbon monoxide is required to produce 1.4×10^3 mol of iron? T/I

   [ans: 2.1×10^3 mol]

3. What amount of potassium is produced when 0.15 mol of potassium oxide decomposes?

   $2\ K_2O(s) \rightarrow 4\ K(s) + O_2(g)$ T/I [ans: 0.30 mol]

7.1 Summary

- A mole ratio is the ratio of the amounts of reactants and/or products in a chemical reaction.
- Coefficients in a chemical equation are used to determine mole ratios.
- A mole ratio is used to predict the amount of reactant required or product produced in a chemical reaction.

7.1 Questions

1. Why is it necessary to balance a chemical equation before determining the amount of product that can be made from a given amount of reactant? T/I
2. What is conserved during a chemical reaction: mass or the number of moles? Justify your choice. K/U T/I
3. Write the mole ratio of potassium chlorate, $KClO_3$, to oxygen, O_2, in the chemical equation

   $2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$ K/U
4. Aluminum metal reacts with chlorine gas to form aluminum chloride, $AlCl_3$:

   $2\ Al(s) + 3\ Cl_2(g) \rightarrow 2\ AlCl_3(s)$

   Copy **Table 3** into your notebook. Complete the table, predicting the amounts of other chemicals that correspond to the given amount. K/U T/I

Table 3 Amounts Involved in the Synthesis of Aluminum Chloride

Amount of Al(s) (mol)	Amount of Cl_2(g) (mol)	Amount of $AlCl_3$(s) (mol)
		2
	1.5	
0.80		
		1.6
	0.45	

5. Explain what is meant by the term “relative” in the statement “chemical equations give the relative amounts of one chemical to another but not the actual amounts.” T/I
6. The chemical equation for the reaction of ammonia, NH_3, with oxygen is

   $4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$

   Predict the amounts of ammonia and oxygen required to produce the following amounts of products:
   (a) 2 mol NO
   (b) 1.8 mol H_2O
   (c) 5.2×10^{-3} mol NO T/I
7. In Unit 2 you learned that one of the hazards of the incomplete combustion of a hydrocarbon is the production of carbon monoxide. Methane is a hydrocarbon that is commonly used as a heating fuel. A possible chemical equation for the incomplete combustion of methane, CH_4, is

   $2\ CH_4(g) + 3\ O_2(g) \rightarrow 2\ CO(g) + 4\ H_2O(g) + \text{energy}$

   Copy **Table 4** into your notebook. Complete the table, predicting the amounts of other chemicals that correspond to the given amount. K/U T/I

Table 4 Amounts Involved in the Combustion of Methane

Amount of CH_4(g) (mol)	Amount of O_2(g) (mol)	Amount of CO(g) (mol)	Amount of H_2O (mol)
3			
	9		
		0.2	
			1

8. The chemical equation for the decomposition of ammonia, NH_3, into its elements is

   $2\ NH_3(g) \rightarrow 3\ H_2(g) + N_2(g)$ K/U T/I C
   (a) Sketch a diagram of four ammonia molecules inside a sealed container and another diagram of the same container after the reaction is complete, similar to Figure 2.
   (b) How many molecules of each product are there?
   (c) What is the simplest ratio of molecules of reactants and products? Relate your answer to the chemical equation.
9. Magnesium nitride is a green ionic solid that reacts with water:

   $Mg_3N_2(s) + 6\ H_2O(l) \rightarrow 3\ Mg(OH)_2(s) + 2\ NH_3(g)$
   (a) Predict the amount of magnesium nitride required to produce 1.5 mol of ammonia.
   (b) Predict the amount of water required to produce 1.5 mol of ammonia.
   (c) What amount of magnesium nitride is required to react completely with 0.25 mol of water? What amounts of the products are expected? T/I

Mass Relationships in Chemical Equations

7.2

There are many dangers aboard a spacecraft. One of the more mundane is the possibility of too much carbon dioxide in the air. Carbon dioxide can be toxic if it is allowed to build up in a confined space. Too much carbon dioxide interferes with the oxygen delivery system in the body. This would make the astronauts sleepy, cause their breathing rate to increase, and lower their blood pH. The source of the carbon dioxide in a spacecraft is the air exhaled by the astronauts. Fortunately for the astronauts, air-scrubbing technology aboard spacecraft removes much of the carbon dioxide before it can accumulate to dangerous levels (**Figure 1**).

WEB **LINK**

In 1970, when space exploration was at its height, *Apollo 13* experienced an equipment failure that could have led to a fatal buildup of carbon dioxide in the spacecraft. To find out how the astronauts solved this problem,

GO TO NELSON SCIENCE

Figure 1 An astronaut changes a lithium hydroxide canister used to scrub carbon dioxide from the interior air during space missions. Each canister lasts about 11 hours for a seven-person crew.

This technology involves passing air through canisters of lithium hydroxide, LiOH. This compound reacts with carbon dioxide in the air to form lithium carbonate and water:

$$CO_2(g) + 2\ LiOH(s) \rightarrow Li_2CO_3(aq) + 2\ H_2O(l)$$

More common hydroxides like calcium hydroxide also undergo this reaction (**Figure 2**). However, lithium hydroxide is preferred for use in space because the mass of the lithium ion is a fraction of the mass of the calcium ion. Keeping mass, and therefore weight, to a minimum is a critical consideration when deciding what to bring onto a spacecraft. Prior to each mission, chemists must determine the quantity of lithium hydroxide required—just enough to scrub exhaled carbon dioxide, plus a little extra in case of flight delays. Any more would be "excess baggage." The required amount is based on the predicted daily output of carbon dioxide by the astronauts. As you will see in this section, the balanced chemical equation for this reaction plays a key role in determining the amount of lithium hydroxide required.

Figure 2 The addition of carbon dioxide makes limewater become cloudy due to the formation of a calcium carbonate precipitate. This reaction is commonly used in high school investigations to test for carbon dioxide.

Masses and Chemical Equations

SKILLS HANDBOOK A6

The coefficients of a balanced chemical equation give the mole ratio of one chemical to another. These values give the relative amount of one chemical required to react with another. They can also be used to predict the amount of product expected.

In Chapter 6, you learned that chemists use mass as a convenient way to determine the number of entities in a sample of a pure substance. The amount present is determined by dividing the mass of the substance by its molar mass ($n = \frac{m}{M}$). The study of the relationships between the amounts of reactants and products in chemical reactions is called **stoichiometry**.

stoichiometry the study of the mass and amount relationships between reactants and products in a chemical reaction

Tutorial 1 Using Stoichiometry to Find Mass

A typical astronaut exhales 8.8×10^2 g of carbon dioxide daily. Determining the mass of lithium hydroxide required to react with this mass of CO_2 begins with an analysis of the chemical equation:

$$CO_2(g) + 2\ LiOH(s) \rightarrow Li_2CO_3(aq) + 2\ H_2O(l)$$

Given: $m_{CO_2} = 8.8 \times 10^2$ g

Required: mass of lithium hydroxide, m_{LiOH}

Solution:

Step 1. Write a balanced equation for the reaction, listing the given value(s), required value(s), and molar masses below the substances being considered in the problem. Use the symbol of the required value since its value is unknown.

$$\begin{array}{llll} CO_2(g) & +\ 2\ LiOH(s) & \rightarrow Li_2CO_3(aq) & +\ 2\ H_2O(l) \\ 8.8 \times 10^2\ g & m_{LiOH} & & \\ 44.01\ g/mol & 23.95\ g/mol & & \end{array}$$

Step 2. Convert mass of given substance(s) to amount(s).

$$n_{CO_2} = 8.80 \times 10^2\ \cancel{g} \times \frac{1\ mol}{44.01\ \cancel{g}}$$

$$n_{CO_2} = 19.995\ mol\ [2\ \text{extra digits carried}]$$

Step 3. Convert amount of given substance to amount of required substance. To do this, multiply the amount of the given substance by a suitable conversion factor derived from the mole ratio in the balanced equation. The mole ratio between CO_2 and LiOH may be expressed as

$$\frac{2\ mol_{LiOH}}{1\ mol_{CO_2}} \text{ or } \frac{1\ mol_{CO_2}}{2\ mol_{LiOH}}$$

Since we want to convert amount of carbon dioxide to amount of lithium hydroxide, use the factor $\frac{2\ mol_{LiOH}}{1\ mol_{CO_2}}$ as follows:

$$n_{LiOH} = 19.995\ \cancel{mol_{CO_2}} \times \frac{2\ mol_{LiOH}}{1\ \cancel{mol_{CO_2}}}$$

$$n_{LiOH} = 39.991\ mol\ [2\ \text{extra digits carried}]$$

Step 4. Convert amount of required substance to mass of required substance.

$$m_{LiOH} = (39.991\ \cancel{mol})\left(\frac{23.95\ g}{1\ \cancel{mol}}\right)$$

$$m_{LiOH} = 958\ g$$

Statement: 958 g of lithium hydroxide is required to react with 8.8×10^2 g carbon dioxide.

Figure 3 summarizes the steps used in this problem.

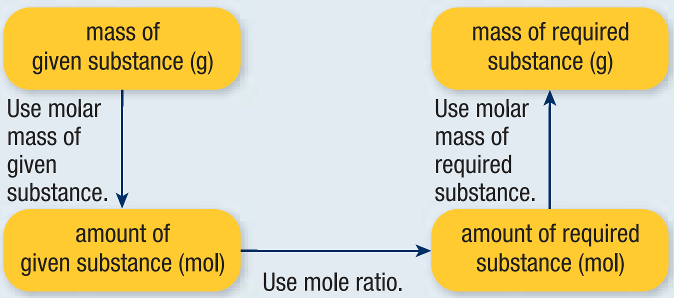


Figure 3 Solving a stoichiometry problem

The strategy used in the lithium hydroxide example can be generalized to solve most stoichiometric problems. As Figure 3 shows, key steps involved in solving stoichiometry problems involving masses are as follows:

1. Convert the given mass to an amount.
2. Apply the mole ratio of the two substances given in the question to predict the required amount of the other substance.
3. Convert the predicted amount to mass.

After completing your calculation, check that the number of significant digits in the final answer agrees with the least number of significant digits in the calculation. Also check that your final answer is numerically correct and that it has the correct unit.

Sample Problem 1: Calculating a Required Mass

An automobile airbag is inflated with nitrogen produced from the decomposition of sodium azide, NaN_3 (**Figure 4**):

$$2\ NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g)$$

The mass of nitrogen in a fully inflated airbag is 87.5 g. What mass of sodium azide is required to produce this mass of nitrogen?

Given: $m_{N_2} = 87.5$ g

Required: mass of sodium azide, m_{NaN_3}

Solution:

Step 1. Write a balanced equation listing given value(s), required value(s), and the corresponding molar masses.

$2\ NaN_3(s) \rightarrow$	$2\ Na(s) + 3\ N_2(g)$
m_{NaN_3}	87.5 g
65.02 g/mol	28.02 g/mol

Step 2. Convert mass of given substance to amount of given substance.

$$n_{N_2} = 87.5\ \cancel{g} \times \frac{1\ mol_{N_2}}{28.02\ \cancel{g}}$$

$$n_{N_2} = 3.1228\ [2\ \text{extra digits carried}]$$

Step 3. Convert amount of given substance to amount of required substance.

$$n_{NaN_3} = 3.1228\ \cancel{mol_{N_2}} \times \frac{2\ mol_{NaN_3}}{3\ \cancel{mol_{N_2}}}$$

$$n_{NaN_3} = 2.0819\ mol$$

Step 4. Convert amount of required substance to mass of required substance.

$$m_{NaN_3} = (2.0819\ \cancel{mol_{NaN_3}})\left(\frac{65.02\ g}{1\ \cancel{mol_{NaN_3}}}\right)$$

$$m_{NaN_3} = 135\ g$$

Statement: The mass of sodium azide required to inflate the airbag is 135 g.

Figure 4 An automobile airbag is designed to inflate in a fraction of a second.

WEB LINK

It is very important that the correct mass of sodium azide is present in an airbag to ensure that it expands to the optimal size. The airbag must then quickly deflate to cushion the motorist's impact. To learn more about the design and function of airbags,

GO TO NELSON SCIENCE

Practice

SKILLS HANDBOOK A6

1. A typical antacid tablet contains 0.50 g of calcium carbonate. The chemical equation for the neutralization of hydrochloric acid with calcium carbonate is

$$CaCO_3(s) + 2\ HCl(aq) \rightarrow CO_2(g) + H_2O(l) + CaCl_2(aq)$$ T/I

(a) What mass of hydrochloric acid will this mass of calcium carbonate neutralize? [ans: 0.36 g]
(b) Predict what mass of calcium chloride will be produced. [ans: 0.55 g]

2. The chemical equation for the complete combustion of propane, $C_3H_8(g)$, is

$$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$$ T/I

(a) What mass of oxygen is required to burn 8.8 kg of propane? [ans: 32 kg]
(b) Predict what mass of carbon dioxide will be produced. [ans: 26 kg]

3. Freshly cut sodium undergoes a synthesis reaction with oxygen in the air. T/I
(a) Write the balanced chemical equation for this reaction.
(b) What mass of oxygen is required to react completely with 3.45 g of sodium? [ans: 1.20 g]

Stoichiometric Amounts

In the lithium hydroxide example at the beginning of Tutorial 1, 40.0 mol of lithium hydroxide is required to absorb 20.0 mol of carbon dioxide. Note that these amounts are multiples of the amounts given in the balanced chemical equation for the reaction. The quantities 40.0 mol LiOH and 20.0 mol CO_2 are stoichiometric amounts. A **stoichiometric amount** is the predicted amount of a reactant, relative to another reactant, that will react according to the balanced chemical equation. The amounts of reactants consumed are in the same proportion as the coefficients in the balanced chemical equation. When stoichiometric amounts of reactants are available for a chemical reaction, no reactants should remain when the reaction is complete. (This assumes that the reaction proceeds to completion.)

stoichiometric amount an amount of reactants that is in the same proportion as the reactant coefficients in the balanced chemical equation

It is important to be aware of stoichiometric amounts when designing any process involving chemical reactions. For example, in Unit 2 you learned that toxic emissions result from the combustion of fossil fuels. Calcium carbonate can be used to scrub sulfur dioxide from these emissions. The chemical reaction occurring in this type of scrubber is

$$SO_2(g) + CaCO_3(s) \rightarrow CaSO_3(s) + CO_2(g)$$

Note that sulfur dioxide and calcium carbonate are in a 1:1 ratio. As a result, sulfur dioxide and calcium carbonate are present in stoichiometric amounts in any mixture containing an equal amount of each compound. (Remember that we are considering equal amounts here, not equal masses.) If less than the stoichiometric amount of calcium carbonate is used, some unused sulfur dioxide escapes into the atmosphere where it can form acid precipitation. In practice, an excess of calcium carbonate is always present to ensure all the sulfur dioxide is captured.

As you will see in Section 7.3, when non-stoichiometric amounts are combined, one reactant remains when the reaction is complete.

Investigation 7.2.1

What Is Baking Soda Doing in Your Cake? (p. 340)
In this investigation you will decompose baking soda, determine the mass of solid product, and decide which of four possible reactions actually occurred. You will test the gases produced to confirm your conclusion.

7.2 Summary

- Stoichiometry makes use of the relationships between mass and amount of the reactants and products in a chemical reaction.
- The amount of products in a chemical reaction can be predicted from the amount of reactants. The masses of products cannot be predicted directly from the masses of reactants.
- Stoichiometry problems involving masses can be solved by converting masses to amounts and using mole ratios (Figure 3).
- Reactants are present in stoichiometric amounts if they are in the same ratio as given by the balanced chemical equation. Theoretically, stoichiometric amounts of reactants are completely used up in a chemical reaction.

7.2 Questions

1. Why is a balanced chemical equation required before solving a stoichiometry problem? K/U
2. A student adds vinegar (a solution of ethanoic acid) to baking soda (sodium hydrogen carbonate). After the bubbling subsides, a clear colourless solution remains. The student then adds more vinegar, producing more bubbles. Were ethanoic acid and sodium hydrogen carbonate initially present in stoichiometric amounts? Explain. T/I
3. Aluminum undergoes a synthesis reaction with oxygen to form aluminum oxide:

   $4\ Al(s) + 3\ O_2(g) \rightarrow 2\ Al_2O_3(s)$

   Which of the following are stoichiometric amounts of the two reactants? K/U T/I

Quantity of Al(s)	Quantity of $O_2(g)$
2.0 mol	1.5 mol
0.60 mol	0.45 mol
0.13 mol	0.12 mol
108 g	96 g
4.0 g	3.0 g

4. The lithium ion battery (**Figure 5**) is an ideal power source for cellphones and digital cameras because its energy output is high for a relatively small mass. A reaction inside the battery generates electrical energy:

   $Li(s) + CoO_2(s) \rightarrow LiCoO_2(s)$ T/I

Figure 5 A lithium ion battery

   (a) When the battery is almost completely drained of power, almost 80 % of the initial amount of lithium metal is converted to $LiCoO_2(s)$. If the initial amounts of the reactants are equal, what percentage of $CoO_2(s)$ remains? Explain your answer.
   (b) When the battery is placed in a recharger, electrical energy is used to reverse the given reaction. How do the relative amounts of lithium and cobalt oxide change as the battery is recharged? Why?

5. Calcium oxide (lime), CaO(s), is a key ingredient in cement. This compound can be made by decomposing calcium hydroxide:

   $Ca(OH)_2(s) \rightarrow CaO(s) + H_2O(g)$ T/I

   (a) Predict what mass of calcium oxide will be produced when 1.00 g of calcium hydroxide is decomposed.
   (b) Predict the mass of water expected.
   (c) Is stoichiometry required to solve (b)? Explain your answer.
   (d) What law of chemical reactions do your answers to (a) and (b) illustrate? Explain.
   (e) A student incorrectly predicted that 1.00 g of calcium hydroxide produces 1.00 g of calcium oxide and 1.00 g of water. Why is this prediction incorrect?
6. Bubbling chlorine gas through a solution of potassium iodide results in the formation of elemental iodine:

   $Cl_2(g) + 2\ KI(aq) \rightarrow I_2(s) + 2\ KCl(aq)$ K/U T/I

   (a) Classify the type of reaction involved.
   (b) What mass of iodine can be produced from 1.5 kg of chlorine?
7. Iron undergoes a synthesis reaction to form iron(III) oxide when it is heated in air. K/U T/I C
   (a) Write a balanced chemical equation for this reaction.
   (b) What amount of iron reacts with 15 mol of oxygen? What amount of iron(III) oxide is produced?
   (c) If 4.60 g of iron(III) oxide is produced, what mass of iron reacted?
   (d) What is the total mass of iron and oxygen that reacted?
   (e) Is stoichiometry required to solve (d)? Explain your answer.
8. Ethanol, $C_2H_6O(l)$, is an important additive in gasoline. T/I A
   (a) Determine the mass of carbon dioxide produced during the complete combustion of 1.15 g of ethanol.
   (b) What might be the environmental consequences if insufficient oxygen were present for the combustion reaction?
9. The addition of heat causes ammonium nitrate, NH_4NO_3, to form dinitrogen monoxide (laughing gas), N_2O, and water. K/U T/I C
   (a) Write a balanced chemical equation for this reaction.
   (b) Classify this reaction.
   (c) Predict what mass of N_2O is produced when 1.00 g of ammonium nitrate reacts.
10. Nitric acid, $HNO_3(aq)$, can be manufactured from ammonia using a series of three chemical reactions called the Ostwald process. The reactions involved are

    $4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$

    $2\ NO(g) + O_2(g) \rightarrow 2\ NO_2(g)$

    $3\ NO_2(g) + H_2O(g) \rightarrow 2\ HNO_3(aq) + NO(g)$

    Determine the mass of nitric acid produced if 425 kg of ammonia reacts. Assume that plenty of oxygen is available. T/I

7.3 Which Reagent Runs Out First?

In Section 7.2, you saw that when exactly stoichiometric amounts of reactants are combined, all the reactants are used up. This produces the maximum possible amount of product. However, in practice this rarely occurs. Carefully measuring stoichiometric amounts of each reactant is tedious and often impractical. Usually, one of the reactants is used up first. The reactant that is completely used up is called the **limiting reagent**. When this reactant runs out, the reaction stops. Some of the other reactants are left over. The **excess reagent** is the substance that is present in a larger quantity than is required. Some of the excess reagent is left over after the reaction. There can only be one limiting reagent in a reaction but there could be more than one excess reagent.

limiting reagent the reactant that is completely consumed in a chemical reaction; the reactant that determines how much product will be formed

excess reagent a reactant that is still present after the reaction is complete

Limiting and Excess Reagents

Lighting a Bunsen burner is a familiar activity in which limiting and excess reagents are involved. When you light a Bunsen burner, you adjust the barrel to limit the rate at which oxygen reaches the fuel. This is because an air/fuel mixture that is rich in fuel is easier to ignite than a "lean" mixture. However, once the burner is lit and you are ready to use it, you have to make another adjustment: you open the barrel to allow more air to enter the burner. Most of the natural gas burned in a Bunsen burner is methane, CH_4. The chemical equation for the complete combustion of methane is

$$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$$

The clean blue flame of a Bunsen burner is evidence of complete combustion (**Figure 1**). This occurs only when plenty of oxygen is available. Under these conditions, methane is the limiting reagent while oxygen is the excess reagent.

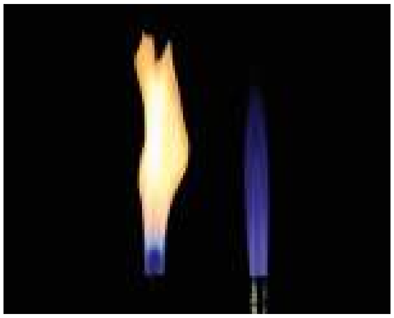


Figure 1 When you light a Bunsen burner, you start with an orange flame then adjust the air intake to get a hot, blue flame.

An Analogy for Limiting and Excess Reagents

We can return to the s'more analogy from Section 7.1 to practise the concepts of limiting and excess reagents. If you recall, the process of making a s'more was described by the "equation"

$$2\ Gc + 4\ Cc + Mm \rightarrow Sm$$

Two crackers combine with four chocolate chips and one marshmallow to make one s'more. When exact stoichiometric amounts of the "reactants" are combined, all the reactants are used up and a whole number of s'mores is produced. However, when non-stoichiometric amounts are available, such as in **Figure 2**, some reactants are in excess. The number of s'mores that can be produced is limited by the number of marshmallows.

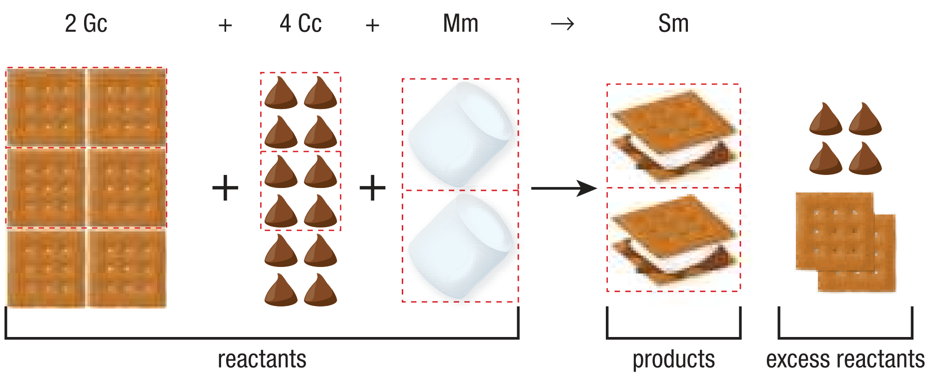


Figure 2 When these amounts of reactants are combined to make s'mores, marshmallows are the limiting "reagent" because there are not enough of them to use up all the other reactants.

Limiting Reagents in Chemical Reactions

We can apply the same logic to a chemical example. **Figure 3** shows 9 hydrogen molecules and 3 oxygen molecules in a flask. What happens when a spark ignites the mixture?

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$$

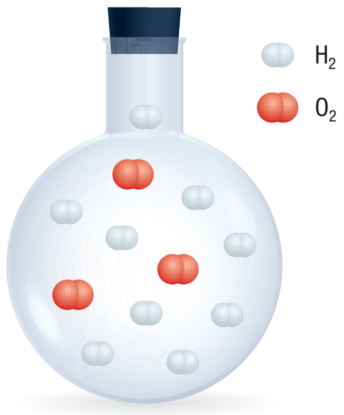


Figure 3 A flask containing 3 oxygen molecules and 9 hydrogen molecules

In **Figure 4**, the molecules are grouped together so that it is easier to see which molecules result in the formation of water molecules. Since the ratio of hydrogen to oxygen in the balanced equation is 2:1, only 6 hydrogen molecules are required to react with the 3 oxygen molecules to form 6 water molecules. As a result, 3 hydrogen molecules are left over. Oxygen, in this case, is the limiting reagent and hydrogen is in excess.

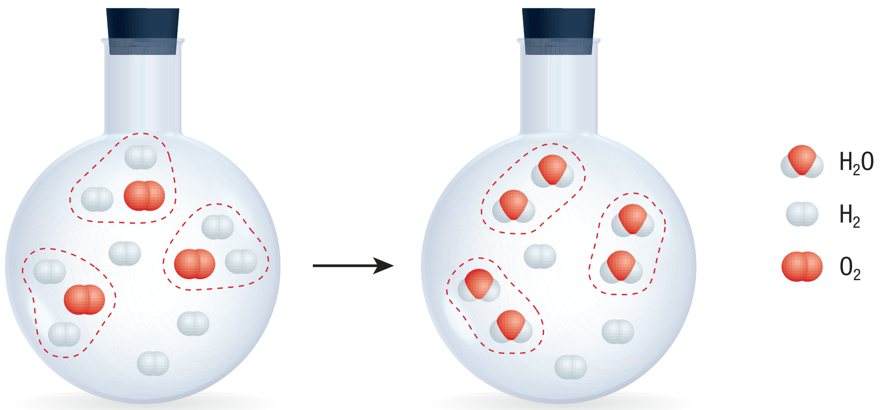


Figure 4 The synthesis of water from its elements. The limiting reagent, oxygen, is completely used up in the reaction. Some hydrogen, the excess reagent, remains after the reaction is complete.

Mini Investigation

Balloon Stoichiometry

Skills: Predicting, Planning, Controlling Variables, Performing, Observing, Analyzing, Communicating

SKILLS HANDBOOK **A1, A2.4**

Sodium hydrogen carbonate (baking soda) reacts with ethanoic acid (in vinegar) in a 1:1 ratio:

$$HC_2H_3O_2(aq) + NaHCO_3(s) \rightarrow H_2O(l) + CO_2(g) + NaC_2H_3O_2(aq)$$

In this experiment, you will explore the quantitative relationships between the two reagents to see which determines the amount of carbon dioxide gas produced.

Equipment and Materials: chemical safety goggles; lab apron; 4 medium-sized balloons; permanent marker; 4 plastic 500 mL bottles; funnel; teaspoon; sodium hydrogen carbonate (baking soda); ethanoic acid solution (vinegar)

1. Put on your safety goggles and lab apron.
2. Stretch the balloons a few times to ensure that they are flexible. Also check that they do not have any holes or tears.
3. Label the four bottles 1, 4, 7, and 10.
4. Using the dry funnel, add 1 level teaspoon (tsp) of sodium hydrogen carbonate to each balloon.
5. Add 1 tsp of ethanoic acid solution to the first bottle, 4 tsp to the second bottle, 7 tsp to the third bottle, and 10 tsp to the fourth bottle.
6. Use your finger to wet the mouth of each balloon with tap water. This will help to maintain a good seal.
7. Carefully stretch a balloon over the mouth of each bottle. Be careful not to allow the sodium hydrogen carbonate to fall into the bottle (**Figure 5(a)**).
8. Raise the balloon on Bottle 1 so that its contents fall into the bottle (**Figure 5(b)**). Swirl to mix. Repeat with the other balloons. Record your observations.

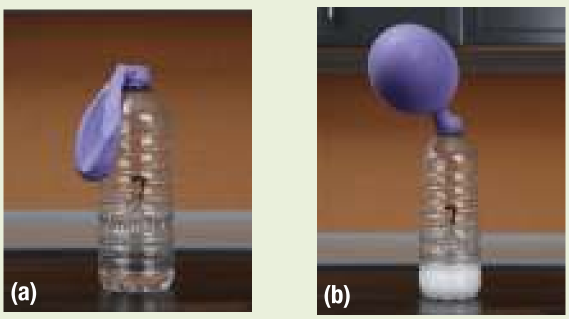


Figure 5 (a) Step 7 (b) Raising the balloon causes the baking soda to fall into the vinegar. Carbon dioxide from the resulting reaction inflates the balloon.

A. Compare the sizes of the balloons after they have inflated. T/I
B. If two or more of your balloons were equally large, explain why. Refer to the quantities of reactants in your explanation. T/I
C. Based on your observations, predict the limiting and excess reagents in each bottle. Test your prediction by planning and conducting a chemical test on the contents of the bottles after the fizzing has stopped. Your test should involve only the materials provided in this investigation. T/I

UNIT TASK BOOKMARK

In the Unit Task (page 352) you will use a technique similar to the one used in the Mini Investigation on page 327. Look for the similarities in the two investigations as you figure out the quantity of limiting reagent in the Unit Task.

Figure 6 Cisplatin (in centre of DNA strands) is an important drug used to treat various types of cancer. Cisplatin works by binding to the DNA of the cancer cell, preventing the cell from replicating.

CAREER LINK

Pharmaceutical development engineers play a crucial role in planning how new medicines will be produced. To find out about this career,

GO TO NELSON SCIENCE

Applications of Limiting Reagents

Why should we be concerned about limiting reagents? Companies that want to reduce both their costs and their impact on the environment pay attention to this decision. Chemical cleanup crews consider the proportions of reactants in their work. Also, anyone who burns hydrocarbons as a heat source should pay attention to the ratio of fuel to air during combustion.

REDUCING COSTS

Companies often choose the most expensive reactant to be the limiting reagent while a large excess of the least expensive reactant is available. This strategy helps ensure that none of the more expensive chemical is left over and wasted. Cisplatin, for example, is an important drug used in chemotherapy treatments for various types of cancer (**Figure 6**). The process to make cisplatin involves four chemical reactions. In the first reaction, a compound called potassium tetrachloroplatinate(II), K_2PtCl_4, is added to a solution containing a large excess of potassium iodide:

$$K_2PtCl_4(aq) + 4\ KI(aq) \rightarrow K_2PtI_4(aq) + 4\ KCl(aq)$$

If you own platinum jewellery, you are well aware that platinum metal is expensive. Platinum compounds are even more expensive. Potassium iodide, in comparison, is cheap. Using a large excess of potassium iodide helps to ensure that all the potassium tetrachloroplatinate(II) is used up and converted into product. A large excess also increases the likelihood that all 4 chlorine atoms in the platinum compound are replaced by iodine. Once the limiting reagent is completely used up, the reaction stops. In this step of the cisplatin production process, $K_2PtCl_4(aq)$ is the limiting reagent while KI(aq) is the excess reagent.

REDUCING ENVIRONMENTAL IMPACT

Controlling the limiting and excess reagents can make a manufacturing process greener. You already know that metal smelting is a major source of sulfur dioxide, a pollutant that reacts in the atmosphere to cause acid rain. In Chapter 5 you learned that flash smelting involves burning ore concentrate in a large excess of oxygen. The oxygen reacts with sulfur, forming sulfur dioxide. Because there is an excess of oxygen, sulfur is the limiting reagent. All of the sulfur is converted to sulfur dioxide. It is economically feasible to collect the sulfur dioxide gas and convert it into useful products such as sulfuric acid.

ENVIRONMENTAL CLEANUP

Strategies used to clean up or neutralize pollutants in the environment can involve limiting and excess reagents. Consider, for example, the treatment of water contaminated with dissolved toxic metal ions from a mining operation. One way to remove the metal ions is to precipitate them as metal hydroxides. This can be done by adding a large excess of a base such as sodium hydroxide. The metal hydroxide precipitates are then filtered out of the solution. In the following example, toxic lead ions are the limiting reagent while the added hydroxide is the excess reagent:

$$Pb^{2+}(aq) + 2\ NaOH(aq) \rightarrow Pb(OH)_2(s) + 2\ Na^+(aq)$$

Spills of concentrated acids or bases can be neutralized by adding a slight excess of sodium hydrogen carbonate (baking soda). Using an excess of baking soda ensures that the spilled compound is completely consumed in the neutralization reaction. The following equation shows the neutralization of sulfuric acid:

$$H_2SO_4(aq) + 2\ NaHCO_3(s) \rightarrow 2\ H_2O(l) + CO_2(g) + Na_2SO_4(aq)$$

A more basic substance such as sodium hydroxide could also be used to neutralize an acid spill. However, a stoichiometric amount of base must be used. The use of too much base would make the resulting mixture basic and therefore corrosive. This could affect the environment just as badly as the original spill.

IMPROVING FUEL EFFICIENCY

We usually want the most possible energy from our home furnaces, car engines, and Bunsen burners. Complete combustion releases more energy, per mole of fuel, than does incomplete combustion. We therefore want our fuels to undergo complete combustion, rather than the less efficient incomplete combustion. To ensure complete combustion, we have to be sure that enough oxygen (usually in air) is supplied. Oxygen must be present in excess for complete combustion to occur.

Complete Combustion of Different Fuels

Different fuels react in different stoichiometric ratios. For example, some barbecues burn methane (natural gas) while others burn propane. Which requires more oxygen for each mole of fuel? **Table 1** shows the equations for the complete combustion of three common fuels. Note that the amount of oxygen required per mole of fuel increases as the size of the hydrocarbon molecule increases. The combustion of a mole of propane requires more than twice the amount of oxygen required for the combustion of a mole of methane.

Table 1 Fuel Combustion Comparison

Fuel	Chemical equation for complete combustion
methane	$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$
propane	$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$
heptane (in gasoline)	$C_7H_{16}(l) + 11\ O_2(g) \rightarrow 7\ CO_2(g) + 8\ H_2O(g)$

Suppose you had a Bunsen burner designed for complete combustion of methane. What do you think would happen if the same burner were attached to a propane tank without adjusting its air/fuel ratio? The burner would likely produce a dirty orange flame, because there is insufficient oxygen for the complete combustion of propane. The reaction would change to incomplete combustion. This is one reason why devices that burn fuels, such as barbecues and vehicles, are designed for a specific fuel. As a result, a natural gas barbecue, for example, cannot be used to burn propane.

WEB LINK

Fuel efficiency researchers have to consider the ratios of various fuels to oxygen. Their aim is to obtain the maximum desired product (usually energy) from the combustion reaction. To learn more about this career,

GO TO NELSON SCIENCE

7.3 Summary

- The limiting reagent is the reactant that runs out in a chemical reaction.
- The amount of product in a chemical reaction is determined by the amount of limiting reagent.
- Excess reagents are reactants that remain after the reaction is complete.
- Controlling the limiting and excess reagents in a chemical reaction has important consumer, health, and environmental applications.

UNIT TASK BOOKMARK

You will use the concept of limiting and excess reagents as you perform the Unit Task on page 352.

7.3 Questions

1. Use an analogy of making sandwiches to illustrate the difference between limiting and excess reagents. Devise your own list of ingredients and ratios. K/U A
2. **Figure 7** shows a flask containing nitrogen and hydrogen. When the temperature of the flask is raised, these elements combine to form ammonia:

   $N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$ K/U T/I

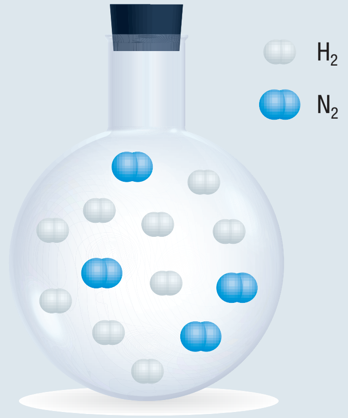


Figure 7 Nitrogen and hydrogen molecules about to react to form ammonia, NH_3

   (a) Identify the limiting and excess reagents in the flask.
   (b) What is the maximum number of ammonia molecules that can be made?
   (c) How many molecules of excess reagent remain?
3. **Figure 8** shows the results of an experiment in which different masses of magnesium were combined with identical volumes of hydrochloric acid. Hydrogen gas, produced by the reaction, inflated the balloons in each trial of the experiment. K/U T/I C

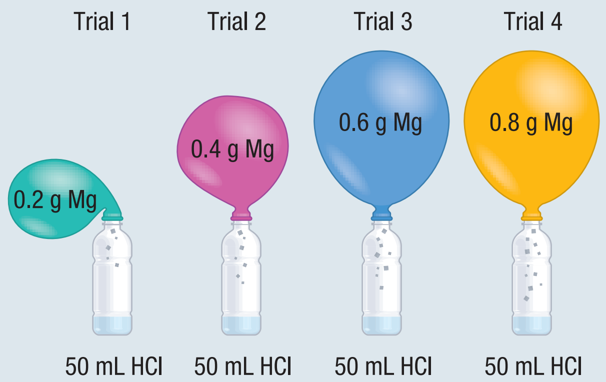


Figure 8 The four bottles contain the same quantities of hydrochloric acid but different quantities of magnesium metal.

   (a) Write a chemical equation for the reaction.
   (b) Classify the reaction.
   (c) Identify the limiting and excess reagents for each trial of the experiment. Explain your answers.
4. Use the concepts of limiting and excess reagents to explain the following observations: K/U A
   (a) The best way to extinguish a grease fire in a pot is to cover it with the lid.
   (b) A sodium sulfide solution precipitates all the iron ions from a sample of contaminated groundwater.
5. Silver is expensive. One way to recover silver from waste silver chloride produced in an investigation is to react it with zinc metal in a basic solution:

   $$2\ AgCl(s) + 4\ NaOH(aq) + Zn(s) \rightarrow Na_2Zn(OH)_4(aq) + 2\ Ag(s) + 2\ NaCl(aq)$$

   Choose the limiting and excess reagents in this process so that all of the silver is extracted from the silver chloride. Justify your choices. K/U T/I
6. Sulfur dioxide can be scrubbed from smokestack emissions with the following reaction:

   $$2\ SO_2(g) + 2\ CaCO_3(s) + O_2(g) \rightarrow 2\ CaSO_4(s) + 2\ CO_2(g)$$

   Based on this chemical equation, which reagent(s) should be the excess reagent(s) so that the most sulfur dioxide is scrubbed from the emissions? K/U T/I A
7. Gold can be extracted from ore containing gold(III) sulfide by passing hydrogen gas through the ore:

   $Au_2S_3(s) + 3\ H_2(g) \rightarrow 3\ H_2S(g) + 2\ Au(s)$

   (a) Identify the limiting and excess reagents if 5 mol gold(III) sulfide is combined with 13 mol hydrogen.
   (b) Based on the chemical equation, which reagent should be the limiting reagent in order for this process to be most profitable? Justify your choice. K/U T/I A
8. A complex system of sensors and computer equipment helps modern automobiles to burn fuel as efficiently as possible. Research the role played by an automobile's oxygen sensor in maintaining engine efficiency. Summarize your findings as a brief web article for new car owners. T/I C A
9. A medical treatment called chelation is used to remove toxic dissolved metals, such as lead, from the body. Research the role of a chemical called EDTA in chelation therapy. T/I A
10. Both diesel and gasoline are common fuels for road vehicles. Which is better for the environment? Research this complex question to identify some of the factors that a motorist should bear in mind when trying to make a "green" decision. T/I A

GO TO NELSON SCIENCE

Calculations Involving Limiting Reagents

7.4

How do you make a bicycle? Obviously, many specific parts must be assembled. Bike manufacturers must keep careful watch over their inventory during the production process. They must ensure that they have a minimum quantity of each bicycle part available at all times. If they run out of one part, the manufacturing process stops. At the same time, it is too costly to maintain a large oversupply of parts.

Similarly, chemical manufacturers must maintain careful inventory of the reactants used in a chemical process. For example, the titanium used to make bicycle frames is relatively abundant as a natural ore (**Figure 1**). However, extracting pure titanium from the ore is complicated and costly. The final step in the process is

$$TiCl_4(g) + 2\ Mg(l) \rightarrow Ti(s) + 2\ MgCl_2(l)$$

Since titanium tetrachloride and magnesium are in a 1:2 ratio in the chemical equation, at least twice the amount of magnesium must be present to ensure that all the expensive titanium tetrachloride is used up to make titanium metal. In practice, a slight excess of magnesium is always present.

Figure 1 Titanium metal is exceptionally light and very strong, making it ideal for bicycle frames.

Limiting Reagent Problems Involving Amounts

If you are given the quantities of two different reactants, you first have to figure out which one is the limiting reagent. You can then use this amount to predict what amount of product will be produced.

Tutorial 1 Solving Limiting Reagent Problems Involving Amounts

To determine the amount of product in a limiting reagent problem, follow the strategy developed in Section 7.2. The only difference is that you must first determine the limiting reagent.

Sample Problem 1: Predicting the Amount of Product

Determine the amount of titanium metal produced when 2.8 mol of titanium(IV) chloride reacts with 5.4 mol of magnesium.

$$TiCl_4(g) + 2\ Mg(l) \rightarrow Ti(s) + 2\ MgCl_2(l)$$

Given: $n_{TiCl_4} = 2.8$ mol; $n_{Mg} = 5.4$ mol

Required: mass of titanium, m_{Ti}

Solution:

Step 1. Write a balanced equation listing given value(s) and required value(s).

$$TiCl_4(g) + 2\ Mg(l) \rightarrow Ti(s) + 2\ MgCl_2(l)$$

2.8 mol 5.4 mol m_{Ti}

Step 2. To determine the limiting reagent, first use the amount of one reactant to find the stoichiometric amount of the other. As you will see, it does not matter which reactant you start with. In this case, we will convert amount of titanium(IV) chloride to amount of magnesium.

$$n_{Mg} = 2.8\ \cancel{mol_{TiCl_4}} \times \frac{2\ mol_{Mg}}{1\ \cancel{mol_{TiCl_4}}}$$

$$n_{Mg} = 5.6\ mol$$

Therefore 5.6 mol of magnesium is required to react with 2.8 mol of titanium(IV) chloride. Only 5.4 mol of magnesium is actually present. This amount is less than what is required. Therefore, magnesium is the limiting reagent and titanium(IV) chloride is the excess reagent.

LEARNING **TIP**

An Alternative Strategy

There is another way to identify the limiting reagent. For the reaction $aA + bB =$ products, if

$$\left(\frac{a}{b}\right) < \left(\frac{n_A}{n_B}\right)$$

then A is in excess, but if

$$\left(\frac{a}{b}\right) > \left(\frac{n_A}{n_B}\right)$$

then A is the limiting reagent.

You would reach the same conclusion if you initially chose titanium(IV) chloride:

$$n_{TiCl_4} = 5.4\ \cancel{mol}_{Mg} \times \frac{1\ mol_{TiCl_4}}{2\ \cancel{mol}_{Mg}}$$

$$n_{TiCl_4} = 2.7\ mol$$

Since more than 2.7 mol of titanium(IV) chloride is present initially, titanium(IV) chloride is the excess reagent. Therefore, magnesium is the limiting reagent.

Once the limiting reagent is determined, the remainder of this problem is the same as the stoichiometry problems in the previous section.

Step 3. Use the amount of limiting reagent to find the amount of required substance.

$$n_{Ti} = 5.4\ \cancel{mol}_{Mg} \times \frac{1\ mol_{Ti}}{2\ \cancel{mol}_{Mg}}$$

$$n_{Ti} = 2.7\ mol$$

Statement: When 2.8 mol of titanium(IV) chloride is combined with 5.4 mol of magnesium, 2.7 mol of titanium will be produced.

Practice

SKILLS HANDBOOK A6

1. A nitric acid spill is neutralized by adding sodium hydrogen carbonate, $NaHCO_3(s)$:

   $HNO_3(aq) + NaHCO_3(s) \rightarrow H_2O(l) + CO_2(g) + NaNO_3(aq)$

   What amount of water is produced when 2.3 mol of nitric acid is combined with 2.0 mol of sodium hydrogen carbonate? T/I [ans: 2.0 mol]
2. Chlorine can be produced in the lab by reacting hydrochloric acid with manganese(IV) oxide:

   $4\ HCl(aq) + MnO_2(s) \rightarrow Cl_2(g) + 2\ H_2O(l) + MnCl_2(aq)$

   What amount of chlorine can be made from 5.2 mol of hydrochloric acid and 1.5 mol of manganese dioxide? T/I [ans: 1.3 mol]
3. Aluminum reacts vigorously with iodine in a synthesis reaction. T/I C
   (a) Write a balanced chemical equation for this reaction.
   (b) Predict the amount of product that can be made from 0.50 mol of aluminum and 0.60 mol of iodine. [ans: 0.40 mol]
4. Aluminum can be used to produce iron from iron(III) oxide. T/I C
   (a) Write a balanced chemical equation for this reaction.
   (b) What amount of iron is expected when 0.26 mol of aluminum is combined with 0.10 mol of iron(III) oxide? [ans: 0.20 mol]
   (c) What amount of the other product is expected? [ans: 0.10 mol Al_2O_3]

Investigation 7.4.1

Copper Collection Stoichiometry (p. 341)

You will use stoichiometry to identify which of two possible iron compounds is formed when copper(II) sulfate solution reacts with an excess of iron.

WEB **LINK**

There are many online videos available to help you learn how to solve limiting reagent problems involving masses. Some use strategies other than those outlined here.

GO TO NELSON SCIENCE

Limiting Reagent Problems Involving Masses

Once you have identified the limiting reagent, you can predict the mass of product using the strategy outlined in Section 7.2.

Tutorial 2 Solving Limiting Reagent Problems Involving Masses

To solve limiting reagent problems involving masses, follow the same strategy as for any other stoichiometry problem involving masses. The only difference is that you first determine which reactant is the limiting reagent, then use the mass of the limiting reagent to determine the masses of product(s).

Sample Problem 1: Predicting the Mass of Product

Methanol, CH_3OH, can be made using a synthesis reaction involving carbon monoxide and hydrogen:

$$CO(g) + 2\ H_2(g) \rightarrow CH_3OH(l)$$

What mass of methanol can be produced from 9.80 g of carbon monoxide and 1.30 g of hydrogen?

Given: $m_{CO} = 9.80\ g$; $m_{H_2} = 1.30\ g$

Required: mass methanol, m_{CH_3OH}

Solution:

Step 1. Write a balanced equation listing given value(s), required value(s), and corresponding molar masses.

$CO(g)$	+	$2\ H_2(g) \rightarrow CH_3OH(l)$
9.80 g		1.30 g
28.01 g/mol		2.02 g/mol

Step 2. Convert mass of given substance to amount of given substance.

$$n_{CO} = 9.80\ \cancel{g} \times \frac{1\ mol}{28.01\ \cancel{g}}$$

$$n_{CO} = 0.34988\ mol\ [2\ \text{extra digits carried}]$$

$$n_{H_2} = 1.30\ \cancel{g} \times \frac{1\ mol}{2.02\ \cancel{g}}$$

$$n_{H_2} = 0.64356\ mol\ [2\ \text{extra digits carried}]$$

Step 3. To determine the limiting reagent, first use the amount of one reactant to find the stoichiometric amount of the other.

$$n_{CO} = 0.64356\ \cancel{mol_{H_2}} \times \frac{1\ mol_{CO}}{2\ \cancel{mol_{H_2}}}$$

$$n_{CO} = 0.32178\ mol$$

0.32178 mol of carbon monoxide is required to completely react with the given amount of hydrogen. Since the amount of carbon monoxide present initially is greater than this amount, carbon monoxide is the excess reagent. Therefore, hydrogen must be the limiting reagent.

Step 4. Use the amount of the limiting reagent to find the amount of required substance.

$$n_{CH_3OH} = 0.64356\ \cancel{mol_{H_2}} \times \frac{1\ mol_{CH_3OH}}{2\ \cancel{mol_{H_2}}}$$

$$n_{CH_3OH} = 0.32178\ mol$$

Step 5. Convert amount of required substance to mass of required substance.

$$m_{CH_3OH} = (0.32178\ \cancel{mol})\left(\frac{32.02\ g}{1\ \cancel{mol}}\right)$$

$$m_{CH_3OH} = 10.3\ g$$

Statement: When 9.80 g of carbon monoxide reacts with 1.30 g of hydrogen, 10.3 g of methanol will be produced.

Figure 2 summarizes the strategy used to solve Sample Problem 1.

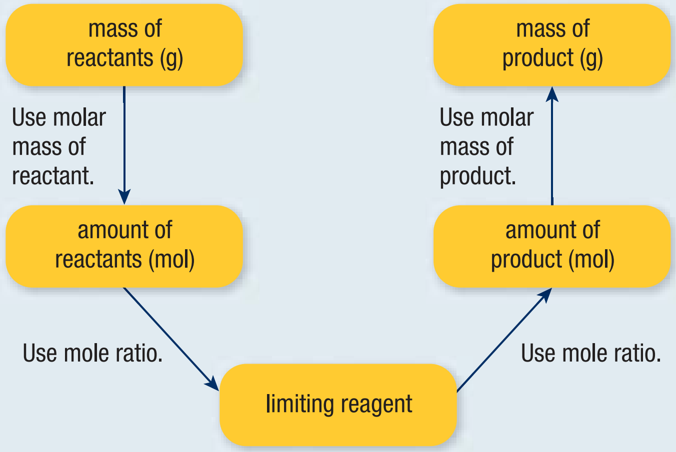


Figure 2 Strategy for solving limiting reagent problems

Alternatively, the calculation can be completed in one step, once the limiting reagent has been identified.

$$m_{CH_3OH} = \left(1.30\ \cancel{g_{H_2}} \times \frac{1\ \cancel{mol_{H_2}}}{2.02\ \cancel{g_{H_2}}}\right)\left(\frac{1\ \cancel{mol_{CH_3OH}}}{2\ \cancel{mol_{H_2}}}\right)\left(\frac{32.02\ g}{1\ \cancel{mol_{CH_3OH}}}\right)$$

$$m_{CH_3OH} = 10.3\ g$$

Practice

SKILLS HANDBOOK A6

1. Silicon carbide, SiC, also known as carborundum, is a hard, industrial abrasive used on grinding wheels to cut metal. Silicon carbide can be made by reacting silicon dioxide, SiO_2, with carbon at high temperatures:

   $SiO_2(s) + 3\ C(s) \rightarrow SiC(s) + 2\ CO_2(g)$

   Determine the mass of silicon carbide expected when 10.0 g of silicon dioxide is combined with 7.00 g of carbon. T/I [ans: 6.67 g]

2. Iron reacts with chlorine gas to form iron(III) chloride:

   $2\ Fe(s) + 3\ Cl_2(g) \rightarrow 2\ FeCl_3(s)$

   What mass of iron(III) chloride is expected if 5.00 g of iron is combined with 9.00 g of chlorine? T/I [ans: 13.7 g]

3. Ammonia, $NH_3(g)$, reacts with oxygen to form nitrogen monoxide, NO(g), and water:

   $4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$ T/I

   (a) Determine the limiting reagent if 0.34 g of ammonia combines with 1.00 g of oxygen.

   (b) What masses of nitrogen monoxide and water are produced in this reaction? [ans: 0.60 g NO; 0.54 g H_2O]

7.4 Summary

- In a limiting reagent problem, the amount of the limiting reagent determines the amount of product.
- The amount of product formed can only be predicted from the amount of the limiting reagent, not from the mass.
- Figure 2 outlines a strategy to solve limiting reagent problems.

1. Copy **Table 1** into your notebook and fill in the missing quantities. K/U T/I

Table 1 Amounts Involved in the Synthesis of Water

$2\ H_2(g)$ +	$O_2(g)$ →	$2\ H_2O(g)$	
Amount of hydrogen (mol)	**Amount of oxygen (mol)**	**Amount of water (mol)**	**Amount of excess reagent remaining (mol)**
2	2		
6	2		
0.4	0.8		
		5	3 mol H_2

2. Copy **Table 2** into your notebook and fill in the missing quantities. K/U T/I

Table 2 Amounts Involved in the Synthesis of Ammonia

$N_2(g)$ +	$3\ H_2(g)$ →	$2\ NH_3(g)$	
Amount of nitrogen (mol)	**Amount of hydrogen (mol)**	**Amount of ammonia (mol)**	**Amount of excess reagent remaining (mol)**
4	13		
0.90	0.25		
		0.16	0.22 mol N_2
		3.0	0.50 mol H_2
1.4			0.80 mol H_2

3. Determine the limiting and excess reagents for each of the following pairs of reactants: K/U T/I
   (a) 0.58 mol of magnesium and 0.20 mol of nitrogen:

   $3\ Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$

   (b) 5.3 mol of calcium and 3.8 mol of aluminum chloride:

   $3\ Ca(s) + 2\ AlCl_3(aq) \rightarrow 3\ CaCl_2(aq) + 2\ Al(s)$

   (c) 0.10 mol of iron pyrite, $FeS_2(s)$, and 0.35 mol of oxygen:

   $4\ FeS_2(s) + 11\ O_2(g) \rightarrow 2\ Fe_2O_3(s) + 8\ SO_2(g)$

4. Chemists can produce silver metal by reacting copper metal with a solution of silver nitrate:

   $Cu(s) + 2\ AgNO_3(aq) \rightarrow 2\ Ag(s) + Cu(NO_3)_2(aq)$ K/U T/I

   (a) Predict the amount of silver produced if 0.24 mol of copper were combined with 0.52 mol of silver nitrate.
   (b) Predict the amount of excess reagent remaining.

5. Aluminum chloride is an important industrial catalyst. It can be made by reacting aluminum metal with hydrochloric acid:

   $2\ Al(s) + 6\ HCl(aq) \rightarrow 2\ AlCl_3(aq) + 3\ H_2(g)$ K/U T/I

   (a) Predict the amount of aluminum chloride produced when 0.35 mol of aluminum is combined with 1.2 mol of hydrochloric acid.
   (b) Predict the amount of excess reagent remaining.

6. Determine the mass of sulfur trioxide produced when 5.8 mol of sulfur dioxide, SO_2, and 2.8 mol of oxygen combine to form sulfur trioxide (**Figure 3**):

   $2\ SO_2(g) + O_2(g) \rightarrow 2\ SO_3(g)$ T/I

Figure 3 On a typical day, the Kilauea volcano in Hawaii emits about 150 to 200 t of sulfur dioxide, most of which reacts to form sulfur trioxide.

7. Hydrogen reacts with chlorine to form gaseous hydrogen chloride. K/U T/I C
   (a) Write a balanced chemical equation for this reaction.
   (b) What mass of product is expected if 10.0 g of hydrogen mixes with 320.0 g of chlorine?

8. Aluminum hydroxide in antacid tablets neutralizes hydrochloric acid in the stomach:

   $Al(OH)_3(s) + 3\ HCl(aq) \rightarrow 3\ H_2O(l) + AlCl_3(aq)$

   If 0.50 g of aluminum hydroxide is placed in a solution containing 0.60 g of hydrochloric acid, predict what mass of aluminum chloride will form. T/I

9. The chemical equation for the combustion of butane is

   $2\ C_4H_{10}(g) + 13\ O_2(g) \rightarrow 8\ CO_2(g) + 10\ H_2O(g)$

   Predict what mass of carbon dioxide is produced from 10.0 g of butane and 30.0 g of oxygen. T/I

10. Titanium ore contains titanium(IV) oxide, TiO_2. During the production of titanium metal, this compound is first converted to titanium(IV) chloride:

    $TiO_2(s) + C(s) + 2\ Cl_2(g) \rightarrow TiCl_4(g) + CO_2(g)$ T/I

    (a) Identify the limiting reagent if 40.0 g of titanium(IV) oxide, 7.0 g of carbon, and 30.0 g of chlorine mix.
    (b) What mass of titanium(IV) chloride can be produced from these quantities?

11. Hydrogen cyanide, HCN(g), can be made in two steps:

    $4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$

    $2\ NO(g) + 2\ CH_4(g) \rightarrow 2\ HCN(g) + 2\ H_2O(g) + H_2(g)$

    If 15.0 g of ammonia, $NH_3(g)$, and 6.0 g of methane, $CH_4(g)$, are present initially with an excess of oxygen, predict what mass of hydrogen cyanide will be produced. T/I

Percentage Yield

Winemaking is one of humankind's oldest technologies based on chemical reactions. Evidence of the earliest winemakers dates back to about 6500 years ago. A key chemical process in the production of wine is fermentation (**Figure 1**). Fermentation is a complicated process during which yeast cells metabolize sugars such as glucose, $C_6H_{12}O_6$, in grape juice. The products of this reaction are an alcohol called ethanol, C_2H_5OH, and carbon dioxide:

$$C_6H_{12}O_6(aq) \rightarrow 2\ C_2H_5OH(l) + 2\ CO_2(g)$$

The toxicity of the ethanol limits the extent to which fermentation can occur. As the concentration of ethanol in the mixture increases, yeast cells begin to die. This stops fermentation before all of the glucose is converted into ethanol. As a result, the volume of ethanol in a bottle of wine rarely exceeds 14 % of the overall volume.

Figure 1 Fermentation locks filled with water are fitted to the top of each bottle, allowing carbon dioxide to escape while preventing oxygen from entering. If the bottles were sealed, the pressure from the accumulating carbon dioxide could break them.

Reaction Yield

Imagine a chemical reaction with a limiting reagent. Ideally, all of the limiting reagent is converted into the desired product. This gives the maximum possible yield of product and is defined as the theoretical yield for the reaction. The **theoretical yield** is a prediction of how much product should form based on the stoichiometry of the reaction. The theoretical yield is achieved if 100 % of the limiting reagent is converted into products. Chemists can predict the theoretical yield before observing the reaction. You predicted theoretical yields in Section 7.4, when you calculated the amount or the mass of product that could be produced.

In practice, however, theoretical yields are rarely achieved. Instead, the **actual yield**—the amount of the product collected during an experiment or industrial process—is usually less. This means that some of the limiting reagent did not become part of the collected product.

There are several factors that can account for why the actual yield is usually less than the theoretical yield. We will look at four of these factors in some detail.

theoretical yield the amount or mass of product predicted based on the stoichiometry of the chemical equation

actual yield the amount or mass of product actually collected during an experiment or industrial process

The Nature of the Reaction

The complete conversion of reactants to products is not always possible. This is sometimes because a competing reverse reaction takes place. For example, ammonia, $NH_3(g)$, is an important industrial chemical made from hydrogen and nitrogen:

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$

As the reaction proceeds, ammonia begins to decompose (**Figure 2**):

$$2\ NH_3(g) \rightarrow N_2(g) + 3\ H_2(g)$$

Chemists have determined the temperature, pressure, and catalyst that give the best yield of ammonia. However, the actual yield is never equal to the theoretical yield.

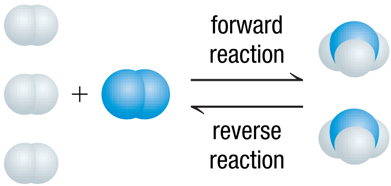


Figure 2 Some ammonia spontaneously decomposes as it forms. This reverse reaction limits the yield of ammonia.

The Experimental Procedure

Regardless of how careful chemists are, they inevitably lose small quantities of material in most investigations. If some of the mass of a reactant is lost after it is measured, or some of the mass of the product is lost before it is measured, the actual yield will be less than the theoretical yield. Materials are lost through spillage during the transfer of solutions, by splattering during heating, and losses during the isolation and purification of the product. Chemists can reduce these losses by improving their skills, using better equipment, and reducing the number of steps in the procedure. Having fewer steps reduces the chances of accidental loss of product.

Impurities

The chemicals used in investigations are rarely 100 % pure. In fact, they come in a wide range of grades (or purities). The mass of the reactant in any sample of starting chemical is therefore less than the measured mass. Failure to take into account the impurities when determining the theoretical yield results in a predicted amount of product that is impossible to achieve. As chemicals age, they may become more impure—particularly if they are not stored properly. Sodium hydroxide, for example, must be stored in an airtight container because it readily absorbs water. Magnesium forms an oxide layer if it is exposed to the air. This increases the mass of a strip of "magnesium" about to be used in an investigation.

Competing Side Reactions

Sometimes, competing reactions prevent some of the reactants from being converted into products. For example, during the synthesis of cisplatin, the chemotherapy drug introduced in Section 7.3, a similar compound called transplatin is also formed (**Figure 3**). Note that cisplatin and transplatin differ only in the position of the groups of atoms attached to the central platinum atom. At first, this difference does not appear to be significant. However, researchers believe that this difference accounts for why transplatin has little to no effect as a cancer treatment. Chemists have found ways to minimize the formation of transplatin and maximize the yield of cisplatin. How they do this, though, is beyond the scope of this course.

Cisplatin Transplatin

Figure 3 Cisplatin is an important drug in the treatment of cancer. The yield of cisplatin is reduced by a competing reaction that produces transplatin.

percentage yield the ratio, expressed as a percentage, of the actual yield to the theoretical yield

Percentage Yield

Comparing the actual yield in a reaction to the theoretical yield gives an indication of how efficient or successful the reaction is at converting reactants into products. A numerical value of this efficiency is determined by calculating the percentage yield for the reaction (**Figure 4**). **Percentage yield** is defined as

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\ \%$$

Percentage yield can be calculated either by comparing the actual mass of product with the theoretical mass, or by comparing the actual amount of product with the theoretical amount.

Chemists often report the percentage yield when they synthesize a new compound or develop a new way of making an existing compound. This information gives other chemists an idea of how much product to expect, if they repeat the procedure.

Figure 4 Only 16 of the initial 20 kernels of popping corn popped. The percentage yield of popcorn is $\frac{16}{20} \times 100\ \% = 80\ \%$.

Tutorial 1 Determining Percentage Yield

In a typical percentage yield problem, you are given the masses of reactants and the mass of a desired product.

Here are some helpful tips to keep in mind when solving the problem:

- Use stoichiometry (based on the balanced chemical equation) to determine the theoretical yield of the desired product.
- If you are told which of the reactants the limiting reagent is, use its mass to calculate theoretical yield (Section 7.2). If the masses of both reactants are given, you must first determine the limiting reagent (Sections 7.3 and 7.4).
- Make sure that the actual yield and theoretical yield are expressed in the same units: either grams or moles. Note that percentage yield has no units because the units of actual and theoretical yield cancel each other in the calculation.

Sample Problem 1 follows essentially the same steps that you used in Section 7.2. The only difference is the calculation of percentage yield in the last step.

Sample Problem 1: Determining Percentage Yield from Mass Data

Methanol, CH_3OH, can be made in a synthesis reaction using carbon dioxide and hydrogen (**Figure 5**):

$$CO_2(g) + 3\ H_2(g) \rightarrow CH_3OH(l) + H_2O(g)$$

During an investigation, 20.0 g of hydrogen was reacted with excess carbon dioxide to produce 102.0 g of methanol. What is the percentage yield of this reaction?

WEB LINK

For online tutorials on determining percentage yield,

GO TO NELSON SCIENCE

Figure 5 Some high-performance race cars use methanol as fuel.

Given: $m_{H_2} = 20.0$ g; $m_{CH_3OH} = 102.0$ g; carbon dioxide is in excess

Required: percentage yield of methanol

Solution:

Step 1. Write a balanced equation listing given value(s), required value(s), and corresponding molar masses.

$$\begin{array}{llll} CO_2(g) + & 3\ H_2(g) & \rightarrow & CH_3OH(l) + H_2O(g) \\ & 20.0\ \text{g} & & 102.0\ \text{g} \\ & 2.02\ \text{g/mol} & & 32.05\ \text{g/mol} \end{array}$$

Step 2. Convert mass of given substance to amount of given substance.

$$n_{H_2} = 20.0\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{2.02\ \cancel{\text{g}}}$$

$$n_{H_2} = 9.9010\ \text{mol}\ [\text{extra digits carried}]$$

Step 3. Convert amount of given substance to amount of required substance.

$$n_{CH_3OH} = 9.9010\ \cancel{\text{mol}_{H_2}} \times \frac{1\ \text{mol}_{CH_3OH}}{3\ \cancel{\text{mol}_{H_2}}}$$

$$n_{CH_3OH} = 3.3003\ \text{mol}$$

Step 4. Convert amount of required substance to mass of required substance.

$$m_{CH_3OH} = (3.3003\ \cancel{\text{mol}})\left(\frac{32.05\ \text{g}}{1\ \cancel{\text{mol}}}\right)$$

$$m_{CH_3OH} = 105.78\ \text{g}\ [\text{extra digits carried}]$$

Step 5. Calculate the percentage yield.

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\ \%$$

$$= \frac{102.0\ \cancel{\text{g}_{CH_3OH}}}{105.78\ \cancel{\text{g}_{CH_3OH}}} \times 100\ \%$$

$$\text{percentage yield} = 96.4\ \%$$

Statement: The percentage yield in the synthesis of methanol reaction is 96.4 %.

Practice

SKILLS HANDBOOK A6

1. Impure nickel can be purified by first reacting it with carbon monoxide:

   $Ni(s) + 4\ CO(g) \rightarrow Ni(CO)_4(g)$ T/I

   (a) Calculate the theoretical yield of nickel carbonyl, $Ni(CO)_4$, in this reaction if 23.5 g of nickel reacts with excess carbon monoxide. [ans: 68.4 g]

   (b) Calculate the percentage yield if 61.0 g of nickel carbonyl is collected. [ans: 89.2 %]

2. An industrial process produces hydrogen by reacting methane with water vapour:

   $CH_4(g) + H_2O(g) \rightarrow CO(g) + 3\ H_2(g)$

   1.61 g of methane is combined with 2.00 g of water, producing 0.50 g of hydrogen. Determine the theoretical yield and percentage yield of hydrogen. T/I [ans: 0.608 g; 82 %]

7.5 Summary

- The theoretical yield is the quantity of product as predicted by the stoichiometry of the chemical equation.
- The actual yield of a reaction is the actual quantity of product collected.
- The actual yield is usually less than the theoretical yield due to material losses, the nature of the reaction, impurities, and competing side reactions.
- The percentage yield of the reaction is the ratio, expressed as a percentage, of the actual yield and theoretical yield. The value of percentage yield has no units.

$$\text{percentage yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\ \%$$

Investigation 7.5.1

What Stopped the Silver? **(p. 342)**
In this experiment you will investigate the effect that the mass of silver nitrate has on the percentage yield of pure silver. You will have to decide which reactant is the limiting reagent.

7.5 Questions

1. How do actual yield and theoretical yield differ? K/U
2. Why are actual yield and theoretical yield rarely equal? K/U
3. Why does reducing the number of steps in an experimental procedure often result in an increase in percentage yield? K/U
4. A student carefully neutralized a sample of hydrochloric acid with sodium hydrogen carbonate in an open container:

   $HCl(aq) + NaHCO_3(s) \rightarrow H_2O(l) + CO_2(g) + NaCl(aq)$

   When the reaction was complete the student heated the remaining solution to remove the water. What effect, if any, would each of the following have on the percentage yield of sodium chloride? Justify your answer in each case. T/I
   (a) not drying the mixture long enough
   (b) adding insufficient sodium hydrogen carbonate to completely react all of the acid
   (c) using an impure form of sodium hydrogen carbonate
   (d) splattering of the reaction mixture out of the container during the reaction
5. Methane reacts with chlorine to form chloromethane:

   $CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g)$

   However, when the reaction mixture is analyzed, some dichloromethane, $CH_2Cl_2(g)$, is detected as well. What effect does this have on the percentage yield of chloromethane? Why? K/U T/I
6. A chemist adds 3.00 g of zinc to a solution containing an excess of silver nitrate. Only 7.2 g of silver metal is collected by the end of the investigation. T/I C
   (a) Write a balanced chemical equation for the reaction.
   (b) Determine the theoretical yield of silver metal.
   (c) Determine the percentage yield of this reaction.
7. To manufacture large quantities of chlorine, sodium hydroxide, and hydrogen, industrial technicians pass electricity through a concentrated sodium chloride solution:

   $2\ NaCl(aq) + 2\ H_2O(l) \xrightarrow{\text{electricity}} Cl_2(g) + 2\ NaOH(aq) + H_2(g)$ T/I

   (a) If the percentage yield of chlorine in this reaction is 95 % and the theoretical yield is 142 g of chlorine, what is the actual yield of chlorine? (Assume that there is excess water present.)
   (b) What mass of sodium chloride is required to produce this yield of chlorine?
8. Coal gasification is a process that converts coal (mostly carbon) into methane, $CH_4(g)$:

   $2\ C(s) + 2\ H_2O(g) \rightarrow CH_4(g) + CO_2(g)$

   If 4.00 mol of coal is combined with excess water vapour, producing 28.0 g of methane, what is the percentage yield of the reaction? T/I
9. Ammonium sulfate, a fertilizer, is made using the following chemical reaction:

   $2\ NH_3(g) + H_2SO_4(aq) \rightarrow (NH_4)_2SO_4(aq)$ T/I

   (a) What is the theoretical yield of ammonium sulfate (in grams) from 3.4 g of ammonia and excess sulfuric acid?
   (b) Determine the percentage yield if 10.4 g of ammonium sulfate is collected.
10. Iron is extracted from the mineral magnetite, Fe_3O_4:

    $Fe_3O_4(s) + 2\ C(s) \rightarrow 3\ Fe(s) + 2\ CO_2(g)$

    A 35.0 g sample of magnetite produces 15.0 g of iron. Determine the percentage yield of this reaction. T/I
11. At 300 °C and 25 times atmospheric pressure, the percentage yield for the production of ammonia from its elements is 24.5 %. If 1.00 kg of nitrogen reacts with excess hydrogen,
    (a) write the balanced chemical equation for this reaction.
    (b) what is the actual yield of ammonia?
    (c) what amount of nitrogen remains unreacted?
    (d) what mass of nitrogen remains unreacted? T/I C
12. Ethanoic acid can be produced industrially from the reaction of methanol, $CH_3OH(g)$, with carbon monoxide, $CO(g)$:

    $CH_3OH(l) + CO(g) \rightarrow HC_2H_3O_2(l)$

    The percentage yield of this reaction is known to be 85 %. What mass of carbon monoxide is required to prepare 2.5 kg of ethanoic acid if methanol is present in excess? T/I
13. The percentage yield of the following reaction is consistently 75.2 %.

    $CH_4(g) + 4\ S(g) \rightarrow CS_2(g) + 2\ H_2S(g)$

    What mass of sulfur is required to produce 50.0 g of hydrogen sulfide if methane, CH_4, is present in excess? T/I
14. Phosphoric acid, $H_3PO_4(aq)$, is used in rust removal products and soft drinks. Phosphoric acid can be made in a two-step process:

    $P_4(s) + 5\ O_2(g) \rightarrow P_4O_{10}(s)$

    $P_4O_{10}(s) + 6\ H_2O(l) \rightarrow 4\ H_3PO_4(aq)$

    In the first step, 25.0 g phosphorus burns in excess oxygen to produce tetraphosphorus decoxide. The percentage yield of this reaction is 90.0 %. T/I
    (a) What mass of tetraphosphorus decoxide is collected from the first step?
    (b) The percentage yield of the second reaction, in excess water, is 95.2 %. What mass of phosphoric acid is produced?
15. Research "atom economy" and prepare a one-paragraph summary of how it differs from percentage yield. K/U T/I A

GO TO NELSON SCIENCE

Investigation 7.2.1 OBSERVATIONAL STUDY

SKILLS MENU

- Questioning
- Researching
- Hypothesizing
- Predicting
- Planning
- Controlling Variables
- Performing
- Observing
- Analyzing
- Evaluating
- Communicating

What Is Baking Soda Doing in Your Cake?

Baking powder is a chemical mixture used to make cakes rise. Sodium hydrogen carbonate (baking soda) is the active ingredient in baking powders. Bubbles of gas released by baking soda give baked goods their spongy consistency. In some types of baking powder, this gas is released prior to baking. In others, the gas is released during baking in the oven as a result of the decomposition of baking soda. There are four possible decomposition reactions that could theoretically occur:

sodium hydrogen carbonate → sodium carbonate (solid) + water vapour + carbon monoxide gas + oxygen gas

sodium hydrogen carbonate → sodium carbonate (solid) + water vapour + carbon dioxide gas

sodium hydrogen carbonate → sodium oxalate (solid) + water vapour + carbon dioxide gas

sodium hydrogen carbonate → sodium oxide (solid) + water vapour + carbon dioxide gas

Your task in this investigation is to determine which reaction actually occurs. You will heat a sample of sodium hydrogen carbonate in a Bunsen burner flame until it decomposes. You will then test the resulting gases to identify the products. Review these tests and be prepared to describe them to your teacher prior to the investigation. You will also determine the mass of the solid product.

Purpose

To determine which of four possible reactions occurs for the decomposition of sodium hydrogen carbonate

Equipment and Materials

SKILLS HANDBOOK A1, A3.2, A3.8

- chemical safety goggles
- lab apron
- heat-resistant test tube
- balance
- Bunsen burner clamped to a retort stand
- spark lighter
- test-tube holder
- scoopula
- test-tube rack
- wooden splint
- sodium hydrogen carbonate, $NaHCO_3(s)$
- test tube half-filled with limewater, $Ca(OH)_2(aq)$

This activity involves open flames. Tie back long hair and secure loose clothing. When heating a test tube containing chemicals, always point the open mouth of the test tube away from yourself and your classmates.

Lime water is an irritant. Avoid skin and eye contact. If you spill lime water on your skin, wash the affected area with lots of cool water and inform your teacher.

Procedure

SKILLS HANDBOOK A1, A2.4, A3

1. Write a balanced chemical equation for each of the four reactions.
2. Determine the mass of the test tube to the nearest 0.01 g. Record this value.
3. Add about 3 g of sodium hydrogen carbonate to the test tube. Record the mass to the nearest 0.01 g.
4. Use stoichiometric calculations to predict the masses of solid products expected in each reaction, using the measured mass of sodium hydrogen carbonate.
5. Plan to conduct chemical tests to identify the unknown gas produced in this reaction. List the necessary equipment and materials. Get your teacher's approval for your plans before proceeding.
6. Put on your chemical safety goggles and lab apron.
7. Assemble the materials you will need to conduct the gas tests.
8. Holding the test tube with a test-tube holder, gently heat the sodium hydrogen carbonate in a Bunsen burner flame for about 2 min. Move the test tube back and forth over the flame.
9. While heating the test tube, conduct chemical tests to identify the gas (other than water vapour) that is produced.
10. Allow the test tube to cool in the test-tube rack.
11. Record the mass of the test tube and its contents.
12. Repeat Steps 8 to 11, heating for about 1 min each time, until the mass remains constant.
13. Clean up your work area as directed by your teacher.

Analyze and Evaluate

(a) According to your evidence, which of the four possible reactions actually occurs? Justify your answer. T/I

(b) Describe how you used your observations to rule out the other possible reactions. T/I

(c) Why was it necessary to reheat the sample? T/I

Apply and Extend

SKILLS HANDBOOK A6

(d) Predict the theoretical yield (in grams) of water vapour and the other gas that will be produced in the reaction.

Investigation 7.4.1 OBSERVATIONAL STUDY

Copper Collection Stoichiometry

A typical high school generates a considerable volume of waste copper(II) sulfate solution each year. Due to the toxicity of this compound, it must be treated to remove the copper(II) ions prior to disposal. One way to do this is to convert the copper ions into elemental copper by reacting them with a reactive metal such as iron. Iron is a multivalent metal: it can form two different ions. Two possible reactions can occur:

$CuSO_4(aq) + Fe(s) \rightarrow Cu(s) + FeSO_4(aq)$

$3\ CuSO_4(aq) + 2\ Fe(s) \rightarrow 3\ Cu(s) + Fe_2(SO_4)_3(aq)$

Purpose

To determine the chemical equation for the reaction of iron with a solution of copper(II) sulfate

Equipment and Materials

SKILLS HANDBOOK A1, A2.4

- chemical safety goggles
- lab apron
- scoopula
- balance
- 50 mL graduated cylinder
- 250 mL beaker
- stirring rod
- retort stand and ring clamp
- funnel
- drying oven or hot plate
- copper(II) sulfate pentahydrate, $CuSO_4{\bullet}5H_2O$
- warm distilled water (about 40 °C)
- iron filings, Fe(s)
- squeeze bottle of distilled water
- filter paper

Copper(II) sulfate is toxic and an irritant. Avoid skin and eye contact. If you spill these chemicals on your skin, wash the affected area with lots of cool water and inform your teacher.

To unplug the hot plate, pull on the plug rather than the cord.

SKILLS MENU

- Questioning
- Researching
- Hypothesizing
- Predicting
- Planning
- Controlling Variables
- Performing
- Observing
- Analyzing
- Evaluating
- Communicating

Procedure

SKILLS HANDBOOK A3.2, A3.6

1. Calculate the stoichiometric amount of iron that would be required to completely react with 6.00 g of copper(II) sulfate pentahydrate in each of the reactions given in the introduction.
2. Select an amount of iron that would make iron the limiting reagent in *both* of the reactions.
3. Convert this amount to a mass of iron.
4. Put on your chemical safety goggles and lab apron.
5. Measure 6.00 g of copper(II) sulfate pentahydrate to the nearest 0.01 g.
6. Measure 50 mL of warm distilled water and pour it into a 250 mL beaker.
7. Dissolve the copper compound in the warm water.
8. Measure the mass of iron filings that you calculated in Step 3. (Measure to 0.01 g.) Add the iron to the solution in the beaker. Swirl the contents of the beaker to mix the reagents.
9. Filter the mixture to collect the solid residue. (See Skills Handbook A3.6 for tips on decanting and filtering mixtures.)
10. Wash the solid residue with two small volumes (about 10 mL each) of distilled water.
11. Dry the solid according to your teacher's directions.
12. Measure the dry mass of the resulting compound.
13. Follow your teacher's instructions for the disposal of waste materials.
14. Wash your hands thoroughly.

Analyze and Evaluate

SKILLS HANDBOOK A6.2

(a) You deliberately selected the limiting reagent in this investigation. What visual observations confirm that this was, indeed, the limiting reagent? T/I

(b) What is the solid product that was formed during the reaction? T/I

(c) Predict the mass of solid product that should form for each of the two possible reactions. T/I

(d) Based on your evidence, which of the two possible reactions occurred? Explain your answer. T/I

(e) If the copper residue is heated strongly in the drying process, copper(II) oxide, $CuO(s)$, may form. What effect would this reaction have on the mass of the copper residue that you collected? T/I

Apply and Extend

(f) Suggest a plan for conducting the investigation so that the other reactant is the limiting reagent. T/I

(g) What reactant masses do you recommend using for this alternative investigation? Justify your choices. T/I

(h) What problem will arise when you attempt to collect the copper residue? T/I

Investigation 7.5.1 CONTROLLED EXPERIMENT

SKILLS MENU

• Questioning	• Planning	• Observing
• Researching	• Controlling Variables	• Analyzing
• Hypothesizing	• Performing	• Evaluating
• Predicting		• Communicating

What Stopped the Silver?

Copper is above silver on the activity series. You would therefore expect a chemical reaction to occur when copper metal is placed in a solution of silver nitrate:

$$Cu(s) + 2\ AgNO_3(aq) \rightarrow 2\ Ag(s) + Cu(NO_3)_2(aq)$$

In fact, under ideal conditions, long needlelike strands of silver metal form on the surface of the copper. These "needles" appear dull grey, not silvery, largely because they do not form a smooth, flat surface that reflects light uniformly. Instead, the needles form a rough surface that reflects light in all directions.

The other reaction product, copper(II) nitrate, remains dissolved. Its ions do not combine to form $Cu(NO_3)_2(s)$. Therefore, a solution of copper(II) nitrate is really a mixture of copper(II) ions and nitrate ions. The blue colour of the solution is due to $Cu^{2+}(aq)$ ions.

The amount of silver produced, as you might expect, depends on the amount of reactants present initially. Specifically, it depends on the amount of limiting reagent.

Testable Question

SKILLS HANDBOOK A1.1, A2.2

Write a testable question that this experiment is designed to answer.

Hypothesis

Predict what effect the mass of silver nitrate used has on the percentage yield. Give reasons for your hypothesis.

Variables

Identify all major variables that will be manipulated, measured, and/or controlled in this investigation. What is the responding variable?

Experimental Design

In this experiment you will place a coiled copper wire in a solution containing a given mass of silver nitrate. You will then collect, dry, and measure the mass of the silver metal that forms on the wire and determine the mass of copper that reacts. You will repeat the experiment using different masses of silver nitrate.

Equipment and Materials

- chemical safety goggles
- lab apron
- latex or rubber gloves
- sandpaper
- balance
- 100 mL beaker
- wooden splint
- stirring rod
- 250 mL beaker
- 30 cm length of copper wire
- silver nitrate, $AgNO_3(s)$
- wash bottle of distilled water
- aluminum foil

Silver nitrate stains the skin. If you spill silver nitrate on your skin, wash the affected area with lots of cool water and inform your teacher.

Procedure

SKILLS HANDBOOK A1.2, A3

1. Put on your safety goggles, lab apron, and gloves.
2. Polish the copper wire with sandpaper.
3. Measure and record the mass of the copper wire to the nearest 0.01 g.
4. Coil the copper wire into a loose coil. Use a wooden splint to suspend the coil in a clean, dry 100 mL beaker so that the coil does not touch the sides or bottom of the beaker (**Figure 1**). Remove the coil.

Figure 1 A coil of copper wire suspended from a wooden splint

5. Place between 1.00 g and 1.40 g of silver nitrate crystals, as directed by your teacher, into the 100 mL beaker. Record the mass to the nearest 0.01 g.
6. Add a small volume of distilled water to the beaker. Use a stirring rod to help dissolve the crystals. Rinse the stirring rod with about 5 mL of distilled water. Collect the rinse water in the beaker. Add more distilled water until the beaker is full to about 3 cm from the top.
7. Add the coil, still attached to the splint, to the beaker.
8. Cover the beaker with a square of aluminum foil and allow it to sit undisturbed overnight. Record your observations (**Figure 2**).

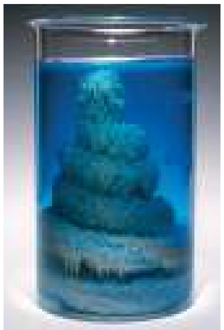


Figure 2 How does your copper coil compare to this one?

9. Determine the mass of a clean, dry 250 mL beaker to the nearest 0.01 g.
10. Use the splint to carefully lift the coil out of the 100 mL beaker and suspend it in the 250 mL beaker. Try not to lose any of the silver crystals in the process.
11. Dislodge the silver from the wire and into the beaker with a stream of distilled water. Try to avoid using more than 25 mL of water.
12. When you have removed the silver, allow the remaining copper wire to dry.
13. Measure and record the final mass of the copper wire to the nearest 0.01 g.
14. By now, most of the silver metal has settled to the bottom of the 250 mL beaker. Carefully decant as much of the liquid as you can without losing any silver. (See Skills Handbook A.3.X for decanting instructions.)
15. Wash the silver residue with about 10 mL of distilled water. Decant this liquid as well.
16. Dry the residue using a method suggested by your teacher.
17. Measure and record the mass of the dry silver residue in the beaker to the nearest 0.01 g.
18. Clean your workstation and wash your hands.
19. Post your mass data, as directed by your teacher.

Analyze and Evaluate

SKILLS HANDBOOK A6

(a) Based on your observations, which is the limiting reagent in this experiment? Justify your selection using visible evidence. K/U T/I
(b) Use the amount of limiting reagent present initially to determine the theoretical yield for this experiment. T/I
(c) Determine the percentage yield. T/I
(d) Use the class results to prepare a graph of percentage yield against the mass of silver nitrate used. C
(e) Answer the Question that you wrote at the beginning of this experiment. Support your answer with evidence. T/I
(f) Suggest two improvements to the way you conducted the experiment that would result in an increased yield. (Assume that you followed the procedure carefully.) T/I
(g) Predict the effect, if any, of the following factors on the percentage yield. Justify your prediction in each case. T/I
  (i) The silver residue was not completely dry when you measured its mass.
  (ii) The silver nitrate was old. (Silver nitrate deteriorates over time, especially if it is exposed to light.)
  (iii) Tap water was used instead of distilled water. (Tap water contains dissolved chloride ions which react with silver ions.)

Apply and Extend

(h) Copper(II) ions, $Cu^{2+}(aq)$, are poisonous. Suggest a plan to remove these ions from the blue solution remaining at the end of the experiment. Explain why your plan will be effective. K/U C
(i) To reduce costs, the silver metal collected during this experiment can be converted back into silver nitrate. Research how this recovery is done and the required safety precautions. Do not conduct the recovery of silver nitrate. T/I

UNIT TASK BOOKMARK

In the Unit Task (page 352). you will use many of the skills that you used in this experiment. Look for similarities between the two investigations.

GO TO NELSON SCIENCE

Summary Questions

1. Create a quiz of 20 questions based on the points listed in the margin on page 314. Be sure to also include correct answers.
2. Look back at the Starting Points questions on page 314. Answer these questions using what you have learned in this chapter. Compare your latest answers with those that you wrote at the beginning of the chapter. Note how your answers have changed.
3. Prepare a one-page summary of each type of numerical problem introduced in this chapter. For each type of problem, list the steps and formulas required to solve the problem.
4. Prepare a brief description of the importance of stoichiometry to a chemist. Try to incorporate the terms in the vocabulary list below.

Vocabulary

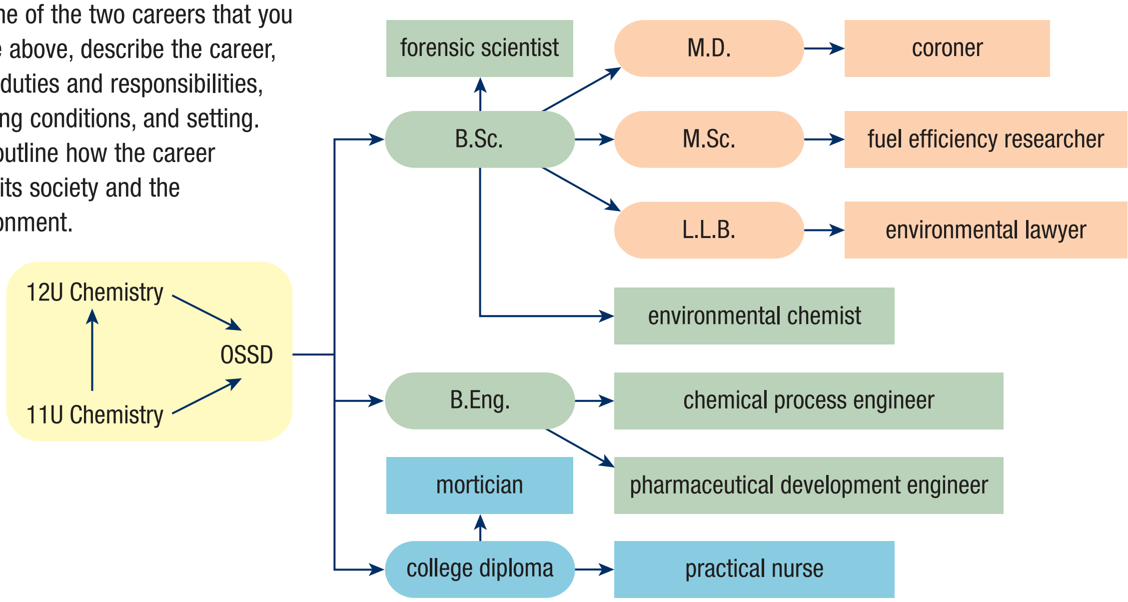
mole ratio (p. 318)	stoichiometric amount (p. 324)	excess reagent (p. 326)	actual yield (p. 336)
stoichiometry (p. 321)	limiting reagent (p. 326)	theoretical yield (p. 336)	percentage yield (p. 337)

CAREER PATHWAYS

SKILLS HANDBOOK A7

Grade 11 Chemistry can lead to a wide range of careers. Some require a college diploma or a B.Sc. degree. Others require specialized or post-graduate degrees. This graphic organizer shows a few pathways to careers mentioned in this chapter.

1. Select two careers, related to stoichiometry in chemical reactions, that you find interesting. Research the educational pathways that you would need to follow to pursue these careers. What is involved in the required educational programs? Prepare a brief report of your findings.
2. For one of the two careers that you chose above, describe the career, main duties and responsibilities, working conditions, and setting. Also outline how the career benefits society and the environment.



GO TO NELSON SCIENCE

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

For each question, select the best answer from the four alternatives.

1. Aluminum carbide powder decomposes in water:
   $Al_4C_3(s) + 12\ H_2O(l) \rightarrow 4\ Al(OH)_3(aq) + 3\ CH_4(g)$
   Which of the following pairs of quantities represent stoichiometric amounts for this reaction? (7.1, 7.2) K/U T/I
   (a) 1 mol Al_4C_3 and 1 mol H_2O
   (b) 4 mol Al_4C_3 and 24 mol H_2O
   (c) 14.4 g Al_4C_3 and 21.6 g H_2O
   (d) 7.2 g Al_4C_3 and 1.8 g H_2O
2. Aluminum hydroxide reacts with sulfuric acid:
   $2\ Al(OH)_3(s) + 3\ H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 6\ H_2O(l)$
   Which of the following quantities would be consumed or produced when 15.6 g of aluminum hydroxide reacts? (7.1, 7.2) T/I
   (a) 13.1 g H_2SO_4
   (b) 34.2 g $Al_2(SO_4)_3$
   (c) 2.4 g H_2O
   (d) 21.6 g H_2O
3. When the air vents of a lit Bunsen burner are closed,
   (a) the gas mixture is leaner
   (b) the combustion is incomplete
   (c) the flame has a clean blue appearance
   (d) the burner is more difficult to ignite (7.3) K/U
4. Hydrazine, N_2H_4, reacts with dinitrogen tetroxide, N_2O_4:
   $2\ N_2H_4(l) + N_2O_4(g) \rightarrow 4\ H_2O(l) + 3\ N_2(g)$
   When 0.10 mol of hydrazine reacts with 0.10 mol of dinitrogen tetroxide, which of the following quantities would be consumed or produced? (7.3, 7.4) T/I
   (a) 0.05 mol N_2O_4
   (b) 0.10 mol N_2O_4
   (c) 0.40 mol H_2O
   (d) 0.13 mol N_2
5. A synthesis reaction produces ammonia:
   $3\ H_2(g) + N_2(g) \rightarrow 2\ NH_3(g)$
   What is the maximum mass of ammonia that can be produced from 1.20 g of hydrogen and 11.2 g of nitrogen? (7.4) T/I
   (a) 13.6 g
   (b) 10.0 g
   (c) 6.8 g
   (d) 3.4 g
6. If 26.4 g of hydrogen, H_2, is produced from 100.0 g of methane, CH_4, reacting with excess water as shown, what is the percentage yield? (7.2, 7.5) T/I
   $CH_4(g) + H_2O(g) \rightarrow CO(g) + 3\ H_2(g)$
   (a) 100 %
   (b) 57.1 %
   (c) 23.4 %
   (d) 69.9 %
7. Ethanol, C_2H_5OH, reacts with phosphorus trichloride, PCl_3:
   $3\ C_2H_5OH(aq) + PCl_3(aq) \rightarrow 3\ C_2H_5Cl(g) + H_3PO_3(aq)$
   Which is the actual mass of chloroethane, C_2H_5Cl, produced if 138.0 g of ethanol is consumed in excess phosphorus trichloride? The percentage yield for this reaction is 93.0 %. (7.2, 7.5) T/I
   (a) 194 g
   (b) 180 g
   (c) 208 g
   (d) 106 g

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

8. The coefficients in a balanced chemical equation directly represent moles of reacting substances. (7.1) K/U
9. Astronauts use calcium oxide to remove exhaled carbon dioxide from their spacecraft. (7.2) K/U
10. Airbags that use sodium azide are inflated by nitrogen gas when they deploy during an accident. (7.2) K/U
11. When phosphorus pentachloride decomposes, the total mass of phosphorus trichloride and chlorine produced must be greater than the mass of phosphorus pentachloride that reacts. (7.2) K/U
12. Calcium carbonate can be used to scrub sulfur dioxide pollution out of gases emitted by coal-fired electrical power plants. (7.2) K/U
13. A chemical reaction comes to a halt when the excess reagent runs out. (7.3) K/U
14. Introducing more air into a rich air/fuel mixture will have no effect on the quantity of energy released when the mixture burns. (7.3) K/U
15. If methane, CH_4, is used in a gas range designed for propane, C_3H_8, the air/fuel mixture will be too rich. (7.3) K/U
16. If sodium hydroxide is used to neutralize an acid spill, the amount added should be stoichiometric. (7.3) K/U
17. Increasing the amount of an excess reagent in a chemical reaction will result in a greater amount of product. (7.3) K/U
18. The theoretical yield in a chemical reaction is calculated from the limiting reagent. (7.4, 7.5) K/U
19. The theoretical yield of a chemical reaction is always less than the actual yield. (7.5) K/U
20. If a competing reaction occurs when a particular chemical product is synthesized, the percentage yield will be decreased. (7.5) K/U

To do an online self-quiz,
GO TO NELSON SCIENCE

Knowledge

For each question, select the best answer from the four alternatives.

1. Which is a correct statement based on the balanced chemical equation given below? (7.1) K/U

   $N_2(g) + O_2(g) \rightarrow 2\ NO(g)$

   (a) If 2 mol of N_2 is consumed, 4 mol of NO will be produced.
   (b) If a sample containing 30 dozen O_2 molecules is consumed, 60 dozen NO molecules will be produced.
   (c) If a sample of NO(g) containing 6 000 000 molecules is produced, 3 000 000 $N_2(g)$ molecules will be consumed.
   (d) all of the above

2. Hydrogen, $H_2(g)$, and nitrogen, $N_2(g)$, combine to form ammonia, $NH_3(g)$:

   $3\ H_2(g) + N_2(g) \rightarrow 2\ NH_3(g)$

   Which amount of nitrogen will react with 18 mol of hydrogen? (7.1) K/U T/I

   (a) 6 mol
   (b) 9 mol
   (c) 18 mol
   (d) 54 mol

3. Why is lithium hydroxide chosen to absorb carbon dioxide exhaled by astronauts on space missions? (7.2) K/U

   (a) It can be recycled onboard the spacecraft.
   (b) It reacts more rapidly with CO_2 than other hydroxides such as NaOH or $Ca(OH)_2$.
   (c) It absorbs the most CO_2 per kilogram.
   (d) It poses no danger to astronauts upon contact.

4. Hydrogen peroxide slowly decomposes in light:

   $2\ H_2O_2(l) \rightarrow 2\ H_2O(l) + O_2(g)$

   Which pair of values are stoichiometric amounts? (7.2) K/U T/I

   (a) 7.2 g H_2O and 6.4 g O_2
   (b) 3.6 g H_2O and 1.6 g O_2
   (c) 3.6 mol H_2O and 3.2 mol O_2
   (d) all of the above

5. Which is true about blue and orange Bunsen burner flames of about the same size? (7.3) K/U A

   (a) The orange flame will occur when the barrel of the burner is fully open.
   (b) The blue flame will be hotter.
   (c) The blue flame will be richer in fuel than the orange flame.
   (d) The orange flame will be cleaner than the blue flame.

6. Aluminum and chlorine undergo a synthesis reaction to form aluminum chloride:

   $2\ Al(s) + 3\ Cl_2(g) \rightarrow 2\ AlCl_3(s)$

   If 8.0 mol of Al is reacted with 10.0 mol of Cl_2, what is the maximum amount of $AlCl_3$ that can be produced? (7.3) T/I

   (a) 6.7 mol
   (b) 8.0 mol
   (c) 12.0 mol
   (d) 15.0 mol

7. Phosphorus burns in oxygen:

   $P_4(s) + 5\ O_2(g) \rightarrow 2\ P_2O_5(s)$

   If 4 mol of molecular phosphorus is available to react with 25 mol of O_2, how many moles of the excess reactant will remain? (7.4) T/I

   (a) 1 mol
   (b) 5 mol
   (c) 20 mol
   (d) 21 mol

8. Which of the following factors limits the yield of alcohol from the fermentation of sugars by yeast? (7.5) K/U

   (a) alcohol chemically reacting with the sugars
   (b) an inability of the alcohol to mix with water
   (c) the escape of $CO_2(g)$ from the fermentation mixture
   (d) the toxicity of the alcohol

9. The actual yield in a chemical procedure is often less than the theoretical yield. What is the most likely reason? (7.5) K/U

   (a) loss of material when a solution is transferred to a new container
   (b) use of a reactant that has become impure by reacting with water or $CO_2(g)$ in the air over time
   (c) formation of an alternative product due to a competing side reaction
   (d) all of the above

10. Which situation describes a reaction with a percentage yield of 100 %? (7.5) K/U

    (a) The amounts of reactants initially present are stoichiometric.
    (b) More product is made than the limiting reagent has the potential to produce.
    (c) No competing reactions occur and the limiting reagent is fully consumed.
    (d) none of the above

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

11. The value that would complete the ratio below is 21.

    $\frac{3}{7} = \frac{?}{42}$ (7.1) T/I

12. The coefficients for the balanced chemical equation below are 2, 2, and 3.

    $3\ H_2(g) + N_2(g) \rightarrow 2\ NH_3(g)$ (7.1) K/U

13. The compound ammonia, NH_3, is commonly used as a fertilizer. (7.1) K/U
14. Stoichiometric amounts in grams are multiples of the coefficients in a balanced chemical equation. (7.2) K/U
15. For any reaction that has a strong tendency to form products, 100 % of the reactants will theoretically be converted to products when stoichiometric amounts of the reactants are initially present. (7.2) K/U
16. When the supply of a reactant in a chemical reaction runs out and the reaction stops, that reactant is said to be in excess. (7.3) K/U
17. A Bunsen burner is most easily ignited when the gas mixture is rich. (7.3) K/U
18. If a butane, C_4H_{10}, tank were attached to a stove designed for propane, C_3H_8, the air/fuel mixture would likely be somewhat leaner than the ideal mixture. (7.3) K/U
19. The quantity that directly links the amount of limiting reagent in a reaction to the amount of product is the mass ratio. (7.4) K/U

Match each term on the left with the most appropriate description on the right.

20. Aluminum metal and sulfur combine to form aluminum sulfide:

    $16\ Al(s) + 3\ S_8(s) \rightarrow 8\ Al_2S_3(s)$ (7.1, 7.2) K/U T/I

(a) amount of S_8 that will produce 32 mol of Al_2S_3	(i) 8 mol
(b) amount of Al that will produce 32 mol of Al_2S_3	(ii) 12 mol
(c) amount of Al that will react with 9 mol of S_8	(iii) 48 mol
(d) amount of Al_2S_3 produced when 16 mol of Al is consumed	(iv) 64 mol

Write a short answer to each question.

21. Aluminum undergoes a single displacement reaction with hydrochloric acid:

    $2\ Al(s) + 6\ HCl(aq) \rightarrow 2\ AlCl_3(aq) + 3\ H_2(g)$

    State each specified ratio, below, in simplest whole numbers. (7.1) K/U T/I

    (a) amount of HCl to amount of Al
    (b) amount of Al to amount of $AlCl_3$
    (c) amount of H_2 to amount of Al

22. The reaction of copper(II) oxide and aluminum metal produces molten copper that can be used to weld conductors together in electrical circuits:

    $3\ CuO(s) + 2\ Al(s) \rightarrow 3\ Cu(l) + Al_2O_3(s) + \text{heat}$ (7.1) K/U T/I

    (a) What amount of Cu is produced if 0.4 mol of Al react?
    (b) What amount of CuO is required to produce 0.2 mol of Cu?

23. What is the gas that is found in the atmosphere and also fills up sodium azide-type airbags when they deploy? (7.2) K/U
24. Give an everyday example of the reaction represented by this chemical equation:

    $CaCO_3(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$ (7.2) K/U A

25. What colour of Bunsen burner flame results when the combustion of methane is incomplete? (7.3) K/U
26. How do competing reactions affect percentage yield? (7.5) K/U

Understanding

27. Explain how a balanced chemical equation is similar to a cooking recipe. Also describe how this analogy is not accurate. (7.1) K/U
28. State for each item below whether the quantity is always conserved, sometimes conserved, or never conserved in balanced chemical equations. Give examples to support your answers. (7.1, 7.2) K/U

    (a) number of molecules
    (b) number of atoms

29. How can a researcher use limewater to test for the presence of carbon dioxide? (7.2) K/U
30. Summarize the steps required to solve a stoichiometry problem in which you are given the mass of a reactant and asked to determine the mass of a product. (7.2) K/U
31. Explain why Bunsen burners always have air vents. (7.3) K/U
32. Both baking soda and sodium hydroxide effectively neutralize acid spills. What are the advantages of using baking soda? (7.3) K/U

33. Why are expensive reactants usually the limiting reagents in industrial chemical processes? (7.3) K/U
34. Chemical engineers must decide how much of various reactants to keep on hand at a chemical plant. How could the concept of limiting reagents guide the engineers? (7.4) K/U
35. Briefly describe how we use mole ratios to decide which of two reactants is limiting when they are not present in stoichiometric amounts. (7.4) K/U
36. Could the actual yield of a reaction ever be greater than the theoretical yield? Explain your thinking. (7.5) K/U T/I
37. Why is the percentage yield always less than 100 % when cisplatin, a cancer drug, is synthesized? What problem does this create for manufacturers of the drug? (7.5) K/U

Analysis and Application

38. Dinitrogen pentoxide decomposes to produce nitrogen dioxide and oxygen:

    $2\ N_2O_5(g) \rightarrow O_2(g) + 4\ NO_2\ (g)$

    Copy and complete each of the following sentences in your notebook and fill in the blanks. (7.1) K/U T/I

    (a) 0.50 mol of dinitrogen pentoxide will produce _____ mol of oxygen and _____ mol of nitrogen dioxide.

    (b) _____ mol of dinitrogen pentoxide will produce _____ mol of oxygen and 1.60 mol of nitrogen dioxide.

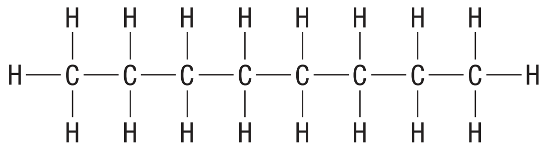
39. Decane, $C_{10}H_{22}(l)$, undergoes complete combustion if there is sufficient oxygen available:

    $2\ C_{10}H_{22}(l) + 31\ O_2(g) \rightarrow 20\ CO_2(g) + 22\ H_2O(g)$ (7.1) K/U T/I

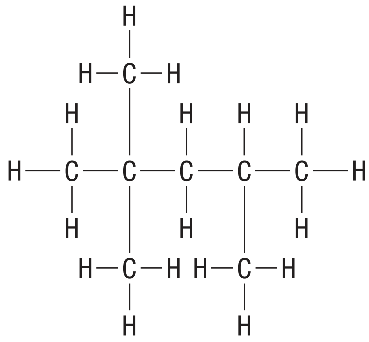
    (a) What amount of oxygen is required for the complete combustion of 5.50 mol of decane?

    (b) A sample of decane is burned, producing 12 mol of carbon dioxide. What amount of water is also produced?

40. When copper and sulfur react, two products are possible: copper(I) sulfide, Cu_2S, and copper(II) sulfide, CuS. A scientist hypothesizes that holding a loop of thin copper wire in the vapour from boiling sulfur will produce pure Cu_2S when the copper and sulfur react. Suggest what measurements, related to mole ratios, the scientist should make to test this hypothesis. (7.1) T/I
41. The structures of two liquid substances commonly found in gasoline are shown in **Figure 1.** Both substances have the chemical formula C_8H_{18}. (7.1, 7.2) K/U T/I C



octane



2,2,4–trimethylpentane

Figure 1

    (a) Write the chemical equation for the complete combustion of these substances.

    (b) Assuming a density of 0.70 g/cm^3 for octane, what mass of carbon dioxide will be formed when 10.0 L of octane is burned in excess oxygen?

42. The combustion reaction between hydrogen and oxygen is used to power rocket engines, including the main engines on the space shuttle. The equation for the reaction is

    $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$ (7.1, 7.3) K/U T/I A

    (a) What is the stoichiometric mole ratio of oxygen to hydrogen in this reaction?

    (b) A typical actual fuel mixture for the reaction in rocket engines is 4 g O_2 : 1 g H_2. Convert this to a mole ratio.

    (c) Based on your answers to (a) and (b), state which reactant is in excess in a fuel mixture with a ratio of 4 g O_2 : 1 g H_2.

43. Think of an example involving common objects that illustrates the meaning of the coefficients in a balanced chemical equation (similar to the examples with s'mores in this chapter). Prepare a table with headings, values, and drawings to illustrate your example. Your table should have three rows: one for individual entities, one for dozens of entities, and one for amount (in moles) of entities. (7.1, 7.2, 7.3, 7.4) T/I C A
44. When copper combines with oxygen, copper(II) oxide is formed. (7.1, 7.3, 7.4) K/U T/I C

    (a) Write the balanced equation for this reaction.

    (b) If 7.0 mol of copper reacts with 4.0 mol of oxygen, what amount of copper(II) oxide is produced? What amount of the excess reactant remains?

45. Glucose, $C_6H_{12}O_6(s)$, can be decomposed into carbon and water:

$C_6H_{12}O_6(s) \rightarrow 6\ C(s) + 6\ H_2O(g)$ (7.2) K/U T/I

(a) Calculate the masses of carbon and water that would be produced when 36.0 g of glucose decomposes completely into carbon and water.

(b) Calculate the mass of glucose that must have fully decomposed if 21.6 g of carbon and 32.4 g of water are produced.

46. When heated, mercury(II) oxide decomposes:

$2\ HgO(s) \rightarrow 2\ Hg(l) + O_2(g)$

This reaction led to the discovery of oxygen in the eighteenth century. What mass of each element is produced when 0.500 mol of mercury(II) oxide is fully decomposed? (7.2) K/U T/I

47. List an object, appliance, or activity in your home that involves stoichiometry and describe how stoichiometry applies. (7.2) K/U A

48. Aluminum foil reacts with aqueous copper(II) chloride:

$2\ Al(s) + 3\ CuCl_2(aq) \rightarrow 2\ AlCl_3(aq) + 3\ Cu(s)$

blue solution (under $CuCl_2(aq)$) colourless solution (under $AlCl_3(aq)$)

Predict what you would expect to observe in each of the following situations. (7.2, 7.3) T/I

(a) A 0.1 mol piece of aluminum foil is placed in a solution containing 0.3 mol of copper(II) chloride, $CuCl_2$.

(b) A 0.1 mol piece of aluminum foil is placed in a solution containing 0.1 mol of copper(II) chloride.

49. Sulfur dioxide can be "scrubbed" out of gaseous emissions from a fossil fuel power plant:

$SO_2(g) + CaCO_3(s) \rightarrow CaSO_3(s) + CO_2(g)$

To make the scrubbing as effective as possible, operators would have to provide an excess of which reactant? Why? (7.2, 7.3) K/U T/I

50. A metallurgist completely dissolves a 10.00 g piece of silver-containing alloy in excess nitric acid and then adds sodium chloride solution to turn the dissolved silver into a precipitate of silver chloride:

$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

She filters out the silver chloride, obtaining 1.20 g. (7.2, 7.3) T/I A

(a) Calculate the mass of silver present in the original sample.

(b) What percentage of the alloy, by mass, is silver?

(c) Why is it important to add excess sodium chloride solution in the precipitation step?

51. Canada consumes about 80 billion m^3 of natural gas annually. Natural gas is principally composed of methane, CH_4, and has a density of about 0.70 kg/m^3. (7.2, 7.3) T/I C A

(a) Write the balanced chemical equation for the complete combustion of methane.

(b) Assuming that all of the natural gas is burned completely, calculate the mass of carbon dioxide in tonnes that is discharged into the atmosphere each year.

(c) Predict the environmental effect of the discharge of the carbon dioxide into the atmosphere.

52. If you cook a meal in a pan over the yellow-orange flame from a campfire, the bottom of the pan becomes blackened. Why does this happen? (7.3) T/I A

53. The concept of limiting reagents can be applied to a natural gas home furnace to save money, conserve fossil fuels, and also protect the environment. Explain how. (7.3) T/I A

54. Preserved fruit sometimes contains a residue of sulfur dioxide, SO_2. You could analyze the residue by soaking the fruit in water to convert the sulfur dioxide to sulfite ions, $SO_3^{2-}(aq)$, adding hydrogen peroxide to convert the sulfite ions to sulfate ions, $SO_4^{2-}(aq)$, and adding barium chloride, $BaCl_2(aq)$, to precipitate the sulfate ions as white barium sulfate, $BaSO_4(s)$. When you measure the mass of the barium sulfate, you can determine the amount of SO_2 originally present in the fruit. Describe a procedure that will precipitate all of the SO_4^{2-} without using any more $BaCl_2(aq)$ than is necessary, thus providing a stoichiometric amount. Specify in your procedure which reagents should be used in excess. (7.3) T/I A

55. **Table 1** shows the results of an investigation in which different masses of zinc were placed in equal volumes of hydrochloric acid. The volume of the hydrogen gas produced in each trial was measured. (7.3) T/I C

(a) Write the balanced chemical equation for the reaction between zinc and hydrochloric acid.

(b) State the limiting reagent in each trial and explain your reasoning.

Table 1 Reaction of Different Masses of Zinc with Hydrochloric Acid

Trial	1	2	3	4	5
volume of HCl(aq) (mL)	50	50	50	50	50
mass of Zn(s)	1.00	1.50	2.00	2.50	3.00
volume of $H_2(g)$ collected (mL)	374	561	611	611	611

56. Sulfur dioxide, $SO_2(g)$, reacts with hydrogen sulfide, $H_2S(g)$, to form sulfur, $S_8(s)$, and water:

$8\ SO_2(g) + 16\ H_2S(g) \rightarrow 3\ S_8(s) + 16\ H_2O(g)$

Copy **Table 2** into your notebook and fill in the missing quantities. (7.4) K/U T/I

Table 2 Amounts in the Reaction of Sulfur Dioxide and Hydrogen Sulfide

Initial amount of $SO_2(g)$ (mol)	Initial amount of $H_2S(g)$ (mol)	Amount of $S_8(s)$ produced (mol)	Amount of excess reactant remaining (mol)
10.0	30.0		
5.0	8.0		
		6.0	2.0 H_2S

57. Arsine, $AsH_3(g)$, was once a product of a test used in forensic analysis. Tissues from a possible poisoning victim were treated with zinc and sulfuric acid. If arsenic was present (usually as diarsenic trioxide, $As_2O_3(s)$), arsine would form:

$As_2O_3(s) + 6\ Zn(s) + 6\ H_2SO_4\ (aq) \rightarrow$
$2\ AsH_3(g) + 6\ ZnSO_4(aq) + 3\ H_2O(l)$ (7.4) T/I

(a) In a lab exercise, a forensic technician reacts 19.8 g of As_2O_3 with 32.7 g of Zn in the presence of excess H_2SO_4. What mass of AsH_3 will be produced?

(b) What mass of the excess reactant (other than sulfuric acid) will remain?

58. The first balloon flight using hydrogen as the buoyant gas took place in 1783. The hydrogen was obtained by reacting iron with sulfuric acid. (7.4) T/I

(a) Write the equation for the single replacement reaction of iron with sulfuric acid. (Hint: The product involves Fe^{2+} ions.)

(b) If 40.0 g of iron is reacted with 100.0 g of sulfuric acid, what mass of hydrogen is made?

(c) What mass of the excess reagent remains?

59. One of the reactions in the industrial production of nitric acid involves the production of nitric oxide:

$4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$ (7.4) T/I

(a) If 4500 kg of ammonia, $NH_3(g)$, react with 7500 kg of O_2, what mass of NO will form?

(b) What mass of the excess reagent will remain?

60. Phosphine, $PH_3(g)$, is an important pesticide. It is released when aluminum phosphide pellets are spread over an area to be treated and the aluminum phosphide reacts with water vapour in the air:

$AlP(s) + 3\ H_2O(l) \rightarrow PH_3(g) + Al(OH)_3(s)$

If 100.0 g of aluminum phosphide pellets are placed in a well-sealed closet and the air in the closet contains 50.0 g of water vapour, what is the theoretical yield, in grams, of phosphine? (7.4, 7.5) T/I

61. Gold can be extracted from ore using sodium cyanide:

$4\ Au(s) + 8\ NaCN(aq) + O_2(g) + 2\ H_2O(l) \rightarrow$
$4\ NaAu(CN)_2(aq) + 4\ NaOH(aq)$

This step of the cyanide process has a 95 % yield. Suppose a gold ore body contains 3.00 kg of gold for every 1.00×10^6 kg of ore. What mass of $NaAu(CN)_2$ will be formed when 1.00×10^6 kg of ore is treated? Give your answer in kilograms, to three significant figures. (7.4, 7.5) T/I

62. Methylbenzene (toluene), $C_6H_5CH_3(l)$, is a common paint thinner. It can be prepared from benzene, $C_6H_6(l)$:

$CH_3Cl(g) + C_6H_6(l) \rightarrow C_6H_5CH_3(l) + HCl(g)$

In an investigation, 25.0 g of benzene is combined with 20.0 g of chloromethane, CH_3Cl, in the presence of an appropriate catalyst. (7.4, 7.5) T/I

(a) Calculate the theoretical yield of toluene.

(b) If 22.0 g of $C_6H_5CH_3$ is actually obtained, what is the percentage yield?

(c) Suggest reasons for the discrepancy between the actual yield and the theoretical yield.

63. Aspirin, $C_9H_8O_4(s)$, can be synthesized from salicylic acid, $C_7H_6O_3(s)$, and acetic anhydride, $C_4H_6O_3(l)$:

$C_7H_6O_3(s) + C_4H_6O_3(l) \rightarrow C_9H_8O_4(s) + HC_2H_3O_2(l)$

In an experiment, a chemistry student obtains 12.2 g of aspirin using this reaction. (7.5) T/I

(a) The student determined the percentage yield to be 72 %. What was the theoretical yield (in grams) for the synthesis?

(b) Considering the theoretical yield, what mass of salicylic acid must have been present initially, assuming acetic anhydride was in excess?

Evaluation

64. To learn about the metabolic rate of crickets, a researcher plans to seal several of the insects in a chamber with a raised mesh floor, beneath which lies a cup containing sodium hydroxide pellets, NaOH(s). The researcher intends to measure the mass gained by the cup of sodium hydroxide pellets as a direct determination of the mass of CO_2 given off by the crickets. The crickets will not be harmed because they cannot come into contact with the sodium hydroxide and they will be removed before the concentration of oxygen in the chamber is significantly lowered. (7.2) T/I A

(a) Evaluate this experimental design.

(b) On what balanced chemical equation has the researcher based the experimental design?

65. In a car design assignment, a college student specifies equal masses of sodium azide in the frontal airbags for front seat occupants. Evaluate this design. (7.2) T/I A

66. There are about 11 ammonia plants in Canada, producing between 4 million and 5 million tonnes of ammonia each year. The diagram in **Figure 2** summarizes the industrial process commonly used to make ammonia.

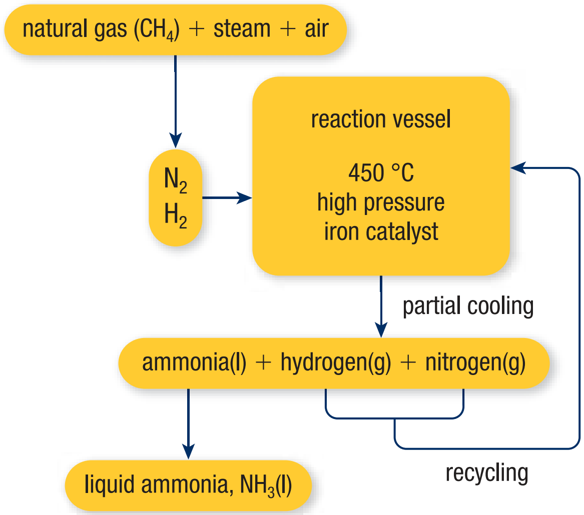


Figure 2

Which part of this process uses principles of green chemistry? How is this beneficial to the company and/or the environment? (7.3) T/I A

67. Table salt is usually "iodized" by the manufacturer. The iodine is generally present as potassium iodide, KI(s). Dissolving a table salt sample in water and then treating it with aqueous potassium iodate, KIO_3(aq), and excess sulfuric acid, H_2SO_4(aq), results in two reactions:

$$2\ KIO_3(aq) + 10\ KI(aq) + 12\ HCl(aq) \rightarrow 6\ I_2(aq) + 6\ H_2O(l) + 12\ KCl(aq)$$

$$3\ KIO_3(aq) + 6\ I_2(aq) + 18\ HCl(aq) + 12\ KCl(aq) \rightarrow 15\ KICl_2(aq) + 9\ H_2O(l)$$

The overall reaction is:

$$KIO_3(aq) + 2\ KI(aq) + 6\ HCl(aq) \rightarrow 3\ KICl_2(aq) + 3\ H_2O(l)$$

As KIO_3(aq) is added in measured increments, the first reaction produces iodine, which turns the solution purple. The colour then fades as the second reaction uses up the iodine. Based on these reactions, describe the measurements and calculations necessary to determine the mass of KI in the salt sample. (7.3, 7.4) T/I

68. A student filters out a precipitate and allows it and the filter paper to dry overnight. He then measures the mass of the filter paper and precipitate. When the student subtracts the original mass of the filter paper from the combined mass of the paper and precipitate, he is perplexed to discover that the amount of precipitate appears to be greater than the theoretical yield. Suggest two reasons this could have happened. (7.5) T/I A

Reflect on Your Learning

69. How did your appreciation of the usefulness of mathematical ratios grow as you studied this chapter? Offer an example of how you might use ratios in your everyday life. A

70. (a) What questions do you still have about limiting reagents and percentage yield?
    (b) Where can you find answers to these questions? K/U A

71. What concept or skill in this chapter did you find to be most difficult? Describe strategies you could employ to improve your understanding. T/I

Research

GO TO NELSON SCIENCE

72. The chemical reaction in sodium azide-propelled airbags produces pure sodium. If the sodium is not immediately converted to a safe compound, it would pose a danger. C A
    (a) Research the chemical reactions that convert the sodium to a safe form. Provide the balanced chemical equations for the reactions that directly produce and remove the sodium.
    (b) In the case of an accident, whom do the sodium-removing reactions protect?
    (c) Do further research and find out about other types of airbag-inflation systems that do not use sodium azide. Summarize what you have discovered by preparing a poster that displays all relevant chemical equations and describes inflation systems through diagrams.

73. Compounds that do not obey the law of definite proportions are called non-stoichiometric compounds. Research these substances, focussing on one particular compound. Write a report of your findings that includes the structure and uses of your chosen compound. A

74. Propane and butane are very common fuels. Find out how they are similar and how they are different. Explore the pros and cons of each one. Summarize your findings in a graphic organizer such as a Venn diagram or t-chart. T/I C A

75. Research how carburetors and fuel injection systems create a proper air/fuel mixture in the internal combustion engines of cars and trucks. Prepare a poster that presents what you have learned in the form of diagrams and flow charts. C A

76. Partially hydrogenated vegetable oils are ingredients in many food products. "Trans fats" are unwanted by-products of the hydrogenation reaction. Research partially hydrogenated vegetable oils to find out how they are made and why trans fats are a problem. Also research what consumers can do to avoid trans fats without sacrificing tasty foods. Prepare an oral report, supported by visual aids, on your discoveries. C A

Fizz Check

Alka-Seltzer is a common remedy for the discomfort caused by excess stomach acid. Each tablet consists of three principal solid ingredients:

- citric acid, $H_3C_6H_5O_7$ (to help the tablet break up)
- acetylsalicylic acid (ASA), $HC_9H_7O_4$, the same pain reliever as in Aspirin (for pain relief)
- sodium hydrogen carbonate (to neutralize stomach acid)

In order to be an effective antacid, an Alka-Seltzer tablet must first break up and dissolve. When it is dropped into water, the three components begin to dissolve (**Figure 1**). Both acids acidify the water and react with sodium hydrogen carbonate, resulting in the following neutralization reactions:

$$\underset{\text{citric acid}}{H_3C_6H_5O_7(aq)} + 3\ NaHCO_3(aq) \rightarrow 3\ H_2O(l) + 3\ CO_2(g) + Na_3C_6H_5O_7(aq)$$

$$\underset{\text{ASA}}{HC_9H_7O_4(aq)} + NaHCO_3(aq) \rightarrow H_2O(l) + CO_2(g) + NaC_9H_7O_4(aq)$$

Figure 1 Bubbles of carbon dioxide indicate that neutralization reactions are proceeding.

The bubbles of carbon dioxide agitate the water, helping the tablet to break up and dissolve.

A large excess of sodium hydrogen carbonate must be present in the tablet in order to react with its own acids as well as to neutralize stomach acid. Conversely, ASA and citric acids must be limiting reagents so that they do not consume all the sodium hydrogen carbonate. The success of the product therefore depends on each tablet containing the correct masses of ingredients.

For this task, consider yourself a quality control chemist at the company that manufactures Alka-Seltzer. You are going to plan and conduct an investigation to determine whether the mass of sodium hydrogen carbonate in a batch of Alka-Seltzer tablets matches the mass given on the product label.

You will add an Alka-Seltzer tablet to different mixtures of water and vinegar. You will know that the reactions are complete when fizzing completely stops. The total mass of the flask and its contents should be less than before because of the loss of carbon dioxide during the reaction. You will use the mass of carbon dioxide produced to predict the mass of sodium hydrogen carbonate that reacts (from each of the three neutralization reactions). This mass represents all the sodium hydrogen carbonate in the tablet provided that vinegar is the excess reagent. To find out how much vinegar is "excess," you will conduct several trials of the investigation, each with a different volume of vinegar. You will draw a graph of mass of sodium hydrogen carbonate against volume of vinegar to help determine in which trials vinegar is the excess reagent.

Purpose

To compare the actual mass of sodium hydrogen carbonate in an Alka-Seltzer tablet with the mass given on the product label

Variables

Identify all major variables that will be manipulated, measured, and/or controlled in this investigation. What is the responding variable?

Equipment and Materials

- chemical safety goggles
- lab apron
- balance
- seven 250 mL Erlenmeyer flasks or plastic cups
- 50 mL graduated cylinder
- dropper
- vinegar (dilute ethanoic acid, $HC_2H_3O_2(aq)$)
- 7 Alka-Seltzer tablets

Procedure

SKILLS HANDBOOK A1.2, A2.4, A3

Design and write a procedure to determine the maximum mass loss in each trial outlined in **Table 1**. Include any necessary safety precautions. Obtain your teacher's approval for your procedure before running your investigation.

Table 1 Contents of Flasks

Trial	Volume of water (mL)	Volume of ethanoic acid (mL)	Number of Alka-Seltzer tablets
1	30	0	1
2	25	5	1
3	20	10	1
4	15	15	1
5	10	20	1
6	5	25	1
7	0	30	1

Analyze and Evaluate

SKILLS HANDBOOK A2.4, A6

(a) Identify the major variables that you measured and/or controlled in this investigation. K/U

(b) Determine the mass of sodium hydrogen carbonate that reacted in each trial of the investigation. T/I

(c) Prepare a graph of mass of $NaHCO_3$ in the tablet versus volume of vinegar used. C

(d) Use the concepts of limiting and excess reagents to explain why the graph levels off as the volume of vinegar increases. K/U T/I

(e) Suggest the minimum volume of vinegar that could be used for this investigation. Explain your answer using the term "stoichiometric amount." K/U T/I

(f) Use your data to answer the question: What is the actual mass of sodium hydrogen carbonate in an Alka-Seltzer tablet? T/I

(g) What mass of sodium hydrogen carbonate is required to neutralize the citric acid and ASA in the tablet? Justify your answer. T/I

(h) After neutralizing the citric acid and ASA in a tablet, what mass of $NaHCO_3$ remains in excess to neutralize stomach acid? T/I

(i) Some carbon dioxide dissolves in water. What effect does this have on your calculated value of sodium hydrogen carbonate per tablet? T/I A

(j) Evaluate the design of the investigation and your skill at conducting it. Are there any parts of the investigation that might cast doubt on your results? Note any likely sources of error. T/I

(k) Suggest improvements to the procedure. T/I

Apply and Extend

(l) Research the solubility of carbon dioxide (in g/L) at the temperature at which you conducted the investigation. Use this data to correct for dissolved carbon dioxide and recalculate the initial mass of sodium hydrogen carbonate per tablet. T/I A

GO TO NELSON SCIENCE

ASSESSMENT CHECKLIST

Your completed Unit Task will be assessed according to the following criteria:

Knowledge/Understanding

- ✓ Identify the manipulated and controlled variables.
- ✓ Distinguish between the limiting and excess reagents in this investigation.

Thinking/Investigation

- ✓ Plan a safe and effective procedure to determine the mass of sodium hydrogen carbonate per tablet.
- ✓ Correctly identify, and effectively manipulate, variables in a controlled study.
- ✓ Conduct the procedures safely and effectively.
- ✓ Record all observations carefully and in an organized manner.
- ✓ Analyze the results.
- ✓ Evaluate the procedure.

Communication

- ✓ Use units and significant digits appropriately in calculations.
- ✓ Graph the data.
- ✓ Prepare a suitable lab report that includes the complete procedures, summary of the observations, any necessary analysis, and evaluation.

Application

- ✓ Analyze the effect of dissolved carbon dioxide on the outcome of the investigation.

For each question, select the best answer from the four alternatives.

1. Which of the following represents a qualitative observation? (6.1) K/U
   (a) development of colour at the end of an investigation.
   (b) height of column of mercury in a thermometer
   (c) waist measurement
   (d) marking of water level on a rain gauge
2. Health Canada's recommended daily intake of sodium is 1500 mg. Which is the closest value to the mass of sodium that an average Canadian actually consumes? (6.2) K/U
   (a) 1000 mg
   (b) 1500 mg
   (c) 2000 mg
   (d) 3500 mg
3. Which of the following is equivalent to 6.02×10^{23} particles? (6.3) K/U
   (a) 1 mol
   (b) 1 mass
   (c) 1 counting unit
   (d) 1 g
4. A carbon atom has a diameter of 1.4×10^{-10} m. If the atoms in 1 mol of carbon were lined up next to one another to form a long chain, what would the length of the chain be? (6.3) K/U T/I
   (a) 7.4×10^{13} m
   (b) 8.4×10^{13} m
   (c) 8.4×10^{14} m
   (d) 7.4×103^{3} m
5. A compound has a molar mass of 14 g/mol. What amount is in 77 g of the compound? (6.4) T/I
   (a) 2.2 mol
   (b) 5.5 mol
   (c) 7.7 mol
   (d) 14 mo
6. Nitrogen gas, N_2, makes up more than 78 % of air. What is the molar mass of nitrogen gas? (6.4) K/U
   (a) $14.01 \dfrac{\text{g}}{\text{mol}}$
   (b) $28.02 \dfrac{\text{g}}{\text{mol}}$
   (c) 6.02×10^{23} entities
   (d) 1.20×10^{24} entities
7. How many atoms there in 5 molecules of hydrogen peroxide, H_2O_2? (6.5) K/U
   (a) 1
   (b) 5
   (c) 10
   (d) 20
8. Propane, C_3H_8, is a hydrocarbon gas. It is abundant in Canada and is a common fuel used for heating and cooking. How many atoms are in 3 mol of propane? (6.5) K/U
   (a) 11
   (b) 33
   (c) 6.02×10^{23}
   (d) 6.62×10^{24}
9. The mass of a sample of a compound is 2.0 g. What is the possible mass of each of the elements in the compound? (6.6) K/U
   (a) 0.25 g carbon, 0.9 g hydrogen, 1.2 g oxygen
   (b) 0.80 g carbon, 0.17 g hydrogen, 1.03 g oxygen
   (c) 1.0 g carbon, 0.5 g hydrogen, 5.0 g oxygen
   (d) 0.95 g carbon, 0.20 g hydrogen, 0.39 g oxygen
10. Which of the following is represented by subscripts in the empirical formula? (6.7) K/U
    (a) simplest ratio of atoms in each molecule
    (b) molar mass of elements
    (c) total number of atoms in each molecule
    (d) total mass of all atoms
11. A mass spectrum of material taken from the surface of a banknote could be identified as cocaine if
    (a) the number of peaks is equal to the molecular mass of cocaine.
    (b) the number of peaks equals the number of atoms per cocaine molecule.
    (c) the lowest peak occurs at the molecular mass of cocaine.
    (d) the number and height of the peaks are characteristic of cocaine. (6.8) K/U
12. What amount of propane, $C_3H_8(g)$, would be required to generate 12 mol of water in a reaction represented by the following equation?
    $C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(l)$ (7.1) T/I
    (a) 1 mol
    (b) 2 mol
    (c) 3 mol
    (d) 4 mol

13. A laboratory analysis requires a sodium chloride solution containing 1.4×10^2 g of salt. What amount of salt is in this solution? (7.2) K/U
    (a) 0.70 mol
    (b) 1.4 mol
    (c) 2.4 mol
    (d) 5.2 mol
14. We use chemicals in many practical day-to-day situations. Sometimes an excess of a reactant is desirable. In which of the following scenarios should an excess of the chemical reactant be used? (7.3) K/U
    (a) fire extinguisher contents
    (b) absorbent used for spill cleanup
    (c) desiccants to absorb moisture
    (d) all of the above
15. Zinc reacts with hydrochloric acid in a displacement reaction:
    $Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
    You combine 1.5 mol of zinc and 3.4 mol of hydrochloric acid. Which of the following statements accurately describes the situation? (7.4) T/I
    (a) Reactants are present in stoichiometric proportions, so there are no limiting reagents.
    (b) There are insufficient amounts of either reactant for the reaction to proceed.
    (c) Hydrochloric acid is the limiting reagent.
    (d) Zinc is the limiting reagent.
16. What is the percentage yield of a reaction if the actual yield is 15 g and the theoretical yield is 22 g? (7.5) K/U
    (a) 15 %
    (b) 22 %
    (c) 37 %
    (d) 68 %

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

17. Measurement of intake and output of fluids is a quantitative measurement. (6.1) K/U
18. The measurement of the mass of a particular gas in the air is an example of qualitative analysis. (6.1) K/U
19. A mole is defined as the number of atoms equal to the mass of the chemical substance. (6.3) K/U
20. The answer to the calculation shown below should have two significant digits. (6.3) K/U
    $$\left(\frac{9.0\ \text{g}}{18.0\dfrac{\text{g}}{\text{mol}}}\right)\left(6.02 \times 10^{23}\frac{\text{molecules}}{\text{mol}}\right)$$
21. We can use the following equation to calculate the molar mass of $MgCl_2$:
    $M_{MgCl_2} = M_{Mg} + 2M_{Cl}$ (6.4) K/U
22. To find the number of entities when given the amount of a substance, divide the amount by Avogadro's constant. (6.3, 6.5) K/U
23. The subscripts in the formula H_2O represent the ratio in which masses of hydrogen and oxygen atoms combine. (6.5) K/U
24. The law of definite proportions states that a compound always has the same proportion of elements by mass. (6.6) K/U
25. Percentage composition is the percentage by number of each element in a compound. (6.6) K/U
26. If molecules of ethanoic acid contain 2 carbon atoms, 4 hydrogen atoms, and 2 oxygen atoms, the molecular formula for this compound could be written as $C_2H_4O_2$. (6.7) K/U
27. A mass spectrometer determines the molecular masses of all vaporized compounds in a sample. (6.6, 6.8) K/U
28. In examining simulated British banknotes for the presence of cocaine, scientists used mass spectrometry to identify substances on the surface of the test strips. (6.8) K/U
29. The molecular formula for glucose is $C_6H_{12}O_6$ and its empirical formula is CH_2O. (6.7, 6.9) K/U
30. A molecular formula represents the simplest ratio of combining atoms in a compound's molecules. (6.9) K/U
31. A mole ratio may represent the amounts of products and/or reactants. (7.1) K/U
32. The amount (n) of a reactant is calculated by dividing the given mass by the molar mass. (7.2) K/U
33. When a reaction has gone to completion, any remaining starting material comes from the limiting reagent. (7.3) K/U
34. The ratio used to convert amount of product to mass has units of mol/g. (7.4) K/U
35. The formula for calculating percentage yield is:
    $$\frac{\text{theoretical yield}}{\text{actual yield}} \times 100\ \%$$ (7.5)

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

Knowledge

For each question, select the best answer from the four alternatives.

1. Which of the following is a quantitative measurement? (6.1) K/U
   (a) colour change of an indicator
   (b) shape of an object
   (c) a count of objects
   (d) the production of gas bubbles
2. You are exploring the effects of salt consumption on human health. Which is a valid research strategy? (6.2) K/U
   (a) Setting up an experiment in which plants are watered with various concentrations of salt solution.
   (b) Collecting data from studies that measured salt intake and health.
   (c) Giving students either a low-salt diet or a high-salt diet, and measuring their blood pressure.
   (d) all of the above
3. What is the correct answer to the following calculation? (6.3) K/U T/I

$$\frac{(6.02 \times 10^{23})(3.0 \times 10^{-3})}{(9.03) \times 10^{-20}}$$

   (a) 5.0×10^{-41}
   (b) 2.0×10^{-40}
   (c) 2.0×10^{40}
   (d) 2.0×10^{0}
4. The exact number of atoms in a standard chemical entity was used to define a mole. Which of the following represents the standard used to define a mole? (6.3) K/U
   (a) 12 g of carbon
   (b) 1 mL of water
   (c) 1 molar mass of O_2
   (d) 2 g of sodium chloride
5. Which of the following samples contains the greatest number of atoms? (6.3, 6.5) T/I
   (a) 1 g of lead
   (b) 5 g of helium
   (c) 10 g of uranium
   (d) 15 g of potassium
6. Which of the following numbers does not have three significant digits? (6.4) K/U
   (a) 0.33
   (b) 1.25
   (c) 5.10
   (d) 8.09
7. Which of the following represents an empirical formula? (6.7) K/U
   (a) CH
   (b) C_2H_2
   (c) C_6H_6
   (d) $C_{12}H_{12}$
8. A chemistry technician is asked to calculate an empirical formula from percentage composition data. The data provided are

   11.19 % H

   88.79 % O

   What is the first calculation in the stepwise process for calculating the empirical formula? (6.7) K/U
   (a) convert grams of each element to moles
   (b) express each element in grams assuming 100 g of sample
   (c) divide the number of moles by the smallest value
   (d) change all numbers to whole numbers
9. British researchers conducted an investigation to determine whether note-counting machines in banks in certain areas of the country were responsible for tainting bank notes with cocaine. Which of the following experimental designs would enable the researchers to answer their question? (6.8) K/U
   (a) Collect used bank notes from around the country and analyze them for cocaine.
   (b) Send clean test-paper strips to banks around the country and have them run through the note-counting machines. Analyze these test strips for cocaine.
   (c) Visit note-counting machines around the country and test them for cocaine.
   (d) Monitor the quantity of cocaine on bank notes over a one-year period.

10. Coefficients are used in chemical equations. What is a function of coefficients in conversions? (7.1) K/U
    (a) calculate densities of solid reactants
    (b) represent molar masses
    (c) serve as a divisor when calculating amounts of reactants
    (d) determine mole ratios
11. You are creating a precipitate of silver chloride, starting with 1.00 mol of silver nitrate and 0.75 mol of potassium chloride:

    $AgNO_3(aq) + KCl(aq) \rightarrow AgCl(s) + KNO_3(aq)$

    What is the maximum amount of silver chloride that can be produced? (7.1) T/I
    (a) 0.50 mol
    (b) 0.75 mol
    (c) 1.00 mol
    (d) 2.00 mol
12. What is an alternative term for the reactant that makes the least amount of product? (7.3) K/U
    (a) stoichiometric amount
    (b) limiting reagent
    (c) molar mass
    (d) excess reagent
13. You are making milkshakes with ingredients in the following ratios:

    1 banana, 1 cup milk, 5 strawberries, 4 ice cubes

    You have the following quantities of ingredients:

    bananas = 3

    milk = 4 cups

    strawberries = 18

    ice cubes = 13

    Which ingredient is the "limiting reagent"? (7.3) K/U
    (a) bananas
    (b) milk
    (c) strawberries
    (d) ice cubes
14. Which of the following represents the quantity that may be used to predict the amount of product formed? (7.4) K/U
    (a) mass of initial reactant
    (b) Avogadro's constant
    (c) molar mass
    (d) limiting reagent
15. Which of the following statements is true related to the product created from most chemical reactions? (7.5) K/U
    (a) theoretical yield is always achieved
    (b) theoretical yield is greater than actual yield
    (c) actual yield is greater than theoretical yield
    (d) percentage yield is never calculated
16. Which of the following values is reported with no units? (7.5) K/U
    (a) percentage yield
    (b) theoretical yield
    (c) actual yield
    (d) molar mass

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

17. Proper placement of a decimal point when prescribing a medication is not vital for accuracy of administration to the patient. (6.1) K/U
18. A mole may be used only to count atoms. (6.3) K/U
19. Avogadro's constant is the number of entities per gram. (6.3) K/U
20. A mole of one element has the same mass as a mole of another element. (6.3) K/U
21. The units used for reporting molar mass are grams. (6.4) K/U
22. 1 mol of H_2 has 1.204×10^{24} molecules. (6.5) K/U
23. The metals combined to make an alloy are always present in the same amounts. (6.6) K/U
24. The proportion of an element in a compound varies with the mass of the sample of the compound. (6.6) K/U
25. The empirical and molecular formulas for CO_2 are identical. (6.7) K/U A
26. The mass of sample required to produce a mass spectrum is usually 1 g or larger. (6.8) K/U
27. To test for the presence of drug contamination by one source, it is necessary to consider other possible sources when designing a field study. (6.8) K/U
28. The molar mass of a compound must be known in order to calculate the molecular formula from the empirical formula. (6.9) K/U
29. Stoichiometry is the area of chemistry that deals with qualitative relationships among reactants and products. (7.2) K/U

30. The reactant that is not entirely consumed in a chemical reaction is the limiting reagent. (7.3) K/U
31. The theoretical yield is the amount of product actually produced. (7.5) K/U

Write a short answer to each question.

32. Describe the concept of molar mass as it relates to a monatomic element. (6.4) K/U
33. Explain why a mass spectrum is a compound's "fingerprint." (6.8) K/U

Understanding

34. Calculate the number of moles and the number of molecules in 400 g of acetic acid, $HC_2H_3O_2(l)$. (6.1, 6.5) T/I
35. A small bar of silver contains 8.0×10^{22} atoms. You are talking to a silversmith and find yourself explaining the concept of moles and atoms. How many moles of silver atoms would you tell the silversmith are contained in this bar? (6.1, 6.5) T/I
36. Calculate the amount and mass of iron(III) oxide, $Fe_2O_3(s)$, that is produced when the following reaction begins with 0.500 mol of pyrite, $FeS_2(s)$:

    $$4\ FeS_2(s) + 11\ O_2(g) \rightarrow 2\ Fe_2O_3(s) + 8\ SO_2(g)$$

    Assume that there is more than enough oxygen present. (6.1, 6.5, 7.1) T/I
37. Canada is home to about 50 major gold mines. In 2008, Canada mined 1.0×10^5 kg of gold compared to about 2.3×10^6 kg mined worldwide, placing Canada in the top 10 gold-producing nations. Pay attention to significant digits as you answer each of the following questions: (6.3) T/I
    (a) What is the average gold production, in grams, for each Canadian mine in 2008? Express your answer using scientific notation.
    (b) What percentage of the world's gold did Canada produce in 2008?
    (c) Suppose the selling price of pure gold is \$43/g. Calculate the value, in dollars, of the gold produced in Canada in 2008. Express your answer in scientific notation.
38. Calculate the mass of a single nitrogen atom. Express your answer in grams, using scientific notation. (6.4) T/I
39. Calculate the molar mass of calcium hydroxide, $Ca(OH)_2(s)$. (6.4) T/I
40. Calculate the mass in grams of 0.473 mol of titanium. (6.4) T/I
41. Calculate the amount of zinc in a 4.20 g sample. (6.4) T/I
42. Distinguish between "mass" and "amount." Describe how these values relate to the empirical formula of a compound. (6.4, 6.7) K/U
43. Compare molar mass to atomic mass. (6.4) K/U
44. How many aluminum atoms are there in 10.00 g of aluminum? (6.5) T/I
45. Imagine a sample of phosphoric acid that has the same mass as the molar mass of the substance. (6.5) T/I
    (a) How many molecules are there in this sample?
    (b) How many atoms are there in the same sample?
46. A certain thermometer contains 0.50 g of mercury. How many atoms are present in this quantity of mercury? (6.5) K/U
47. Ancient alchemists used aqua fortis or "strong water" to dissolve silver. Find the number of molecules in 6.0 mg of this substance, now known as nitric acid, $HNO_3(l)$. (6.5) K/U
48. You are given 100 g of a compound. The compound is composed of 37 % hydrogen and 63 % oxygen. How many grams of each element are present in the 100 g sample? (6.6) T/I
49. A solution containing a total of 20.00 g of potassium nitrate, $KNO_3(aq)$, and potassium chloride, $KCl(aq)$, is combined with silver nitrate solution, $AgNO_3(aq)$. A white precipitate of silver chloride, $AgCl(s)$, forms:

    $$AgNO_3(aq) + KCl(aq) \rightarrow AgCl(s) + KNO_3(aq)$$

    If the mass of the silver chloride precipitate is found to be 8.66 g, what are the percentages by mass of potassium nitrate and potassium chloride in the original 20.00 g ? (7.2) T/I
50. Calculate the percentage by mass of iron and oxygen in iron(III) oxide, $Fe_2O_3(s)$. (6.6) T/I
51. The molecular formula of nicotine is $C_{10}H_{14}N_2$. What is the percentage by mass of nitrogen in this compound? (6.6) T/I
52. Calculate the percentage composition of hydrogen in glucose, $C_6H_{12}O_6(s)$. (6.6) T/I
53. A certain metal alloy is composed of 10 % tin, 16 % antimony, and 74 % lead. If you were to have 500 g of the alloy, how many grams of each of the components would be found in this sample? (6.6) T/I

54. A compound composed of only sulfur and aluminum is analyzed. The mass ratio of aluminum to sulfur is found to be 1.00:1.78. Find the empirical formula of this compound. (6.7) T/I

55. Calculate the empirical formula and molecular formula of a compound that consists of 80 % carbon and 20 % hydrogen, and has a molar mass of 30.1 g/mol. (6.7, 6.9) T/I

56. Explain the relationship between an empirical formula and a molecular formula. (6.7, 6.9) K/U

57. Carbon disulfide, $CS_2(l)$, is made from carbon, C(s), and sulfur dioxide, $SO_2(g)$. The balanced equation for the reaction is given below. In a small-scale industrial test, 500 g of CS_2 is produced and the yield of CS_2 is 86.0 %. What is the mass of the theoretical yield in grams? (6.7, 7.5) T/I

    $3\ C(s) + 2\ SO_2(g) \rightarrow CS_2(l) + 2\ CO_2(g)$

58. Oxygen reacts with hydrogen to produce water. What mole ratio would be used to calculate the amount of water that can be obtained when oxygen reacts with excess hydrogen? (7.1) T/I

59. What amount of ammonia, $NH_3(g)$, can be produced from 15 mol of hydrogen reacting with excess nitrogen:

    $3\ H_2(g) + N_2(g) \rightarrow 2\ NH_3(g)$ (7.1) T/I

60. Dinitrogen pentoxide, $N_2O_5(s)$, is an unusual compound: it is an ionic compound under some conditions and a molecular compound under others. Given the chemical equation below, calculate how many moles of nitric oxide, NO(g), form when 1.75 mol of dinitrogen pentoxide decomposes completely.

    $2\ N_2O_5(s) \rightarrow 4\ NO(g) + 3\ O_2(g)$ (7.1, 7.2) T/I

61. Sodium fluoride, NaF(s), is one of the most commonly used compounds for fluoridating water supplies. When sodium fluoride is decomposed, sodium and fluorine are produced:

    $2\ NaF(s) \rightarrow 2\ Na(s) + F_2(g)$

    Suppose a sample of NaF was decomposed by electricity, yielding 28.8 g of sodium. What is the mass of fluorine, in grams, that was also produced? (7.1, 7.2) T/I

62. Create a list of four practical advantages of applying the concept of limiting reagents. (7.3) K/U A

63. High-purity sodium bromide, NaBr(s), is used in medicine as a sedative. In the laboratory, sodium bromide can be prepared from sodium metal and liquid bromine:

    $2\ Na(s) + Br_2(l) \rightarrow 2\ NaBr(s)$

    Find the limiting reagent when 1.8 mol of Na and 1.4 mol of Br_2 are reacted. (7.3, 7.4) T/I

64. Crystals of magnetite, $Fe_3O_4(s)$, are found in some bacteria and the brains of certain animals. These crystals may help organisms to sense Earth's magnetic field. Magnetite can be synthesized in the laboratory:

    $3\ Fe(s) + 4\ H_2O(g) \rightarrow Fe_3O_4(s) + 4\ H_2(g)$

    What is the limiting reagent when 84.0 g of iron is reacted with 50.0 g of water? (7.3, 7.4) T/I

65. Ethylene glycol, $HOCH_2CH_2OH(l)$, is commonly found in antifreeze. The process for its manufacture uses the reaction between ethylene oxide, $C_2H_4O(g)$, and water:

    $C_2H_4O(g) + H_2O(g) \rightarrow HOCH_2CH_2OH(l)$

    If 880 kg of ethylene oxide is reacted with 400 kg of water, what is the limiting reagent? (7.3, 7.4) T/I

66. The following molecules are combined in a flask: 20 carbon monoxide molecules and 30 oxygen molecules. They react to form carbon dioxide, fully consuming the limiting reagent:

    $2\ CO(g) + O_2(g) \rightarrow 2\ CO_2(g)$

    Describe the quantity of excess reagent that remains. (7.3) K/U

Analysis and Application

67. Suppose you are a doctor and a regular patient of yours is on a trip to a foreign country. You receive a communication from a hospital requesting results of recent blood and other tests for your patient because she is very ill and needs treatment. To ensure that the information is properly interpreted, what must you consider when you send the information to the hospital? Include an example in your answer. (6.1) K/U A

68. Your 22-year-old cousin is visiting you for two weeks and has forgotten to bring his medication for a thyroid condition. He knows that your father also takes thyroid medication and says that he can just share your father's pills until he goes back home. What advice would you give your cousin about his plan? Explain your thinking. (6.1) K/U A

69. Aquarium owners must routinely add different chemicals to maintain the environment in the tank for their fish. Water conditioners, for example, can remove chloramines, ammonia, and toxic metals, all of which can harm fish. (6.1) K/U A
   (a) Why would it be important to pay attention to accuracy in maintaining an aquarium?
   (b) An aquarium owner has been adding three drops of a dechlorinating product to her tank each week. She runs out of the old product and buys a bottle of a new product. Can she safely assume that three drops is still the proper quantity? Why or why not?
70. A friend's grandmother is perplexed about the results of recent medical tests. A urinalysis done by her doctor shows that she has too much sodium in her system even though she has greatly reduced the quantity of salt she adds to her food. (6.2) K/U A
   (a) Explain how her diet could still be the problem.
   (b) State two actions she should take to be sure she is eliminating as much salt as possible from her diet.
71. Explain how chemists use experimental information to establish the chemical formula of a compound. (6.1, 6.7) K/U A
72. A mass spectrum of an unknown compound is shown in **Figure 1**. (6.1, 6.8) K/U T/I

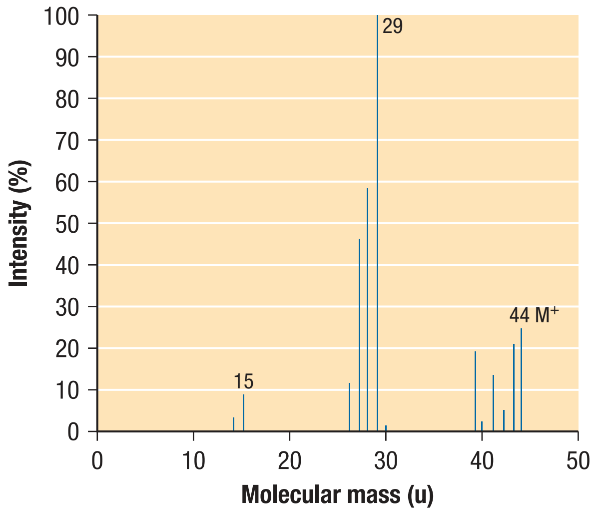


Figure 1

   (a) What qualitative information about the compound can be obtained from this spectrum?
   (b) What quantitative information about the compound can be obtained from this spectrum?
73. Devise a procedure to estimate the number of individual peanuts in a bag of peanuts in the shell. Explain how this estimation procedure could be adapted to estimate the number of molecules in a substance that, like I_2, is made of two-atom molecules. (6.3) K/U T/I A
74. Calcium chloride, $CaCl_2(s)$, is often sprinkled on sidewalks and streets to melt ice and snow. A chemist is researching the de-icing properties of calcium chloride and places an order for 5 mol of this compound. The stockroom sends a 5 kg bag. Explain why this is incorrect. (6.3, 6.4) T/I
75. A teaching assistant asks you to help him devise a series of formulas to instruct students in mathematically converting mass to moles and moles to mass using balanced chemical equations. Write a brief procedure to assist the students in the learning process. (6.4) K/U C A
76. Examine the three formulas and answer the questions that follow:
   (i) CH
   (ii) C_2H_2
   (iii) C_6H_6
   (a) Which formula or formulas represent an empirical formula?
   (b) Calculate the percentage composition of the elements in each of the formulas. Compare the results.
   (c) Calculate the molar mass for each compound. Compare the results.
   (d) For each formula, choose the diagram in **Figure 2** below that represents the percentage composition for the elements in the compound. (6.4, 6.6, 6.7) K/U T/I A

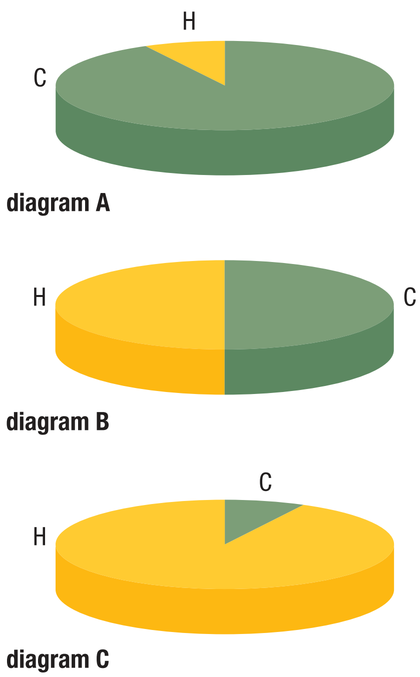


Figure 2

77. Given the information below, determine the molecular formula for the chemical substance described. (6.4, 6.9) K/U T/I

Molar mass = 180.1 g/mol

Composition: C = 54.5 %; H = 9.2 %; O = 36.3 %

78. Magnesium metal is often sold in ribbon form. When held in a burner flame, the ribbon ignites and glows very brightly as it burns. The magnesium in the ribbon is converted to a white ash, which is magnesium oxide. Suppose a certain magnesium ribbon has a "linear density" of 1.3 g/m (each metre of ribbon has a mass of 1.3 g). Calculate the number of atoms in 0.25 m of the ribbon. Show all steps. (6.5) T/I C

79. Compare the percentage by mass of iron, Fe(s), in iron(II) chloride, $FeCl_2(s)$, and iron(III) bromide, $FeBr_3(s)$. Explain how you arrived at your response. (6.6) T/I A

80. Incomplete combustion in a car's engine cylinders produces carbon monoxide. In the catalytic converter of a modern car, this carbon monoxide is converted to carbon dioxide. What happens to the mass percentage of oxygen in the compounds when the carbon monoxide is converted to carbon dioxide? Give numerical evidence to support your answer. (6.6) T/I A

81. The percentage composition of potassium sulfate, $K_2SO_4(s)$, is described as:

28.57 % K

14.28 % S

57.15 % O

You are provided with two containers of K_2SO_4. Container A contains 11 g of the compound. Container B has 23 g of the compound. (6.6) K/U A

(a) How will the difference in mass affect the percentage composition?

(b) Relate the "law of definite proportions" to the percentage composition in each of the two containers.

82. Which would be better to provide to a company that will be manufacturing a compound: the compound's empirical formula or its molecular formula? Explain and provide an example. (6.7) K/U C A

83. A manufacturer produces antacid tablets that neutralize stomach acid:

$$CaCO_3(s) + 2\,HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

The manufacturer would like to advertise the amount of gastric acid that each antacid tablet will neutralize. Suggest the steps that would be required to produce accurate data. (7.2, 7.3, 7.4) K/U T/I A

84. (a) Apply the concept of stoichiometric reactions to theoretical and actual yields.

(b) List three factors that have an impact on the completion of stoichiometric reactions and yields. (7.2, 7.5) K/U C A

85. Draw a diagram of the formation of silver chloride, AgCl(s):

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

Use spheres to represent the ions. Demonstrate silver ions as the limiting reagent and chloride ions as the excess reagent. (7.3) K/U C A

86. Copper wire is used for the following reaction. Explain how you would know if the copper wire was the limiting reagent. (7.3) K/U A

$$Cu(s) + 2\,AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$$

87. A particular chemical reaction is an integral part of a manufacturing process. Each batch of reagents is to contain a designated mass of the limiting reagent and the excess reagent. The supervising chemist would like to order reactants for a future round of manufacturing. The chemist has asked you to calculate how much of the excess reagent will remain after the current manufacturing process is complete. Explain how you would proceed with this calculation. (7.3) K/U T/I A

88. A chemical reaction involves three reactants. The reactants are used in a ratio of 1A:2B:1C. Looking at the diagram in **Figure 3** below, determine which reactant will be the limiting reagent. Explain your answer. (7.3) K/U T/I

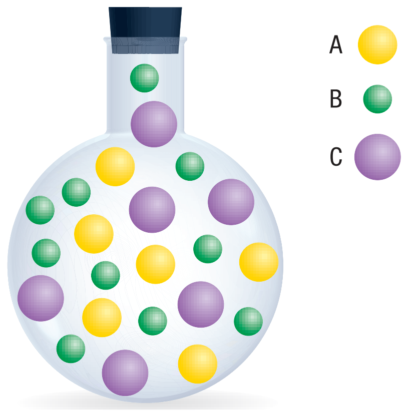


Figure 3

89. Explain the calculation of percentage yield for a reaction. (7.5) K/U

90. Compare and contrast the use of recipes and expected amounts to the concept of chemical reactions and percentage yield. (7.5) K/U A

91. A laboratory assistant is asked to calculate the percentage yield for the compound A_2B_3. The data for the calculation are given below. Explain whether this calculation could be performed with a single step or whether multiple steps would be required. Justify your answer, showing appropriate mathematical equations. State the units that would be used to report the percentage yield. (7.5) K/U T/I C

    Theoretical yield: 14.79 g A_2B_3

    Actual yield: 6733 mg A_2B_3

Evaluation

92. Evaluate the information that you have learned in this unit and relate it to prospective careers in
    (a) forensic science
    (b) the medical assistant/technician field (6.1, 6.2, 6.6, 6.8, 6.9) A

93. Inuit in Canada are increasingly experiencing high blood pressure, and this trend is well correlated to an increase in the quantity of salt in their diet. (6.2) K/U A
    (a) What reasons are there to believe that the increasing high blood pressure among Inuit is due to increased salt consumption?
    (b) Give two reasons why Inuit have begun to consume more salt.
    (c) What role should the government assume in discouraging Inuit from ingesting salt? Give reasons for your position.

94. A garage manager must determine the number of metal ball bearings (BBs) in a stock container by mass measurements. He counts out 5 BBs and finds their total mass to be 1.9 g. He also finds the total mass of the stock container and the BBs to be 239.3 g and the mass of an identical, empty stock container to be 115.8 g. The store manager then uses his data to calculate the number of BBs in the stock container:

    $$\text{Average mass of one BB} = \frac{1.9\text{ g}}{5} = 0.4\text{ g}$$

    $$\text{Mass of BBs in container} = 239.3\text{ g} - 115.8\text{ g} = 123.50\text{ g}$$

    $$\text{Number of BBs in container} = \frac{123.50\text{ g}}{0.4\text{ g/BB}} = 300\text{ BBs}$$

    (6.3) K/U T/I A
    (a) Evaluate the manager's calculations with regard to significant digits. If necessary, rewrite any steps, recalculating with the adjusted quantities. For any step that you have to rewrite, explain why the manager's calculation is in error.
    (b) Should the manager regard the answer to the calculation in (a) as a perfect measure of the number of ball bearings in the container?
    (c) How could the manager have changed his procedure to yield an answer with a greater number of significant figures?
    (d) What is the purpose of paying attention to significant digits in calculations with measured quantities?

95. An analytical chemist sets up the apparatus shown in **Figure 4**. He plans to burn a compound known to contain carbon, hydrogen, oxygen, and chlorine. By measuring the mass of the original sample and the masses of carbon dioxide and water collected in the scrubbers, the chemist believes he will have the necessary data to determine the empirical formula of the compound. Will the chemist's method yield useful results? Why or why not? (6.6) K/U T/I

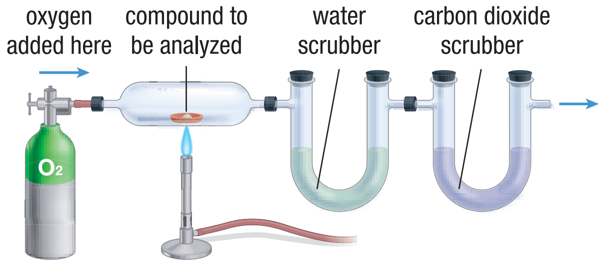


Figure 4

96. A chemist wants to select a substance to serve as an effective carbon dioxide scrubber in a combustion analysis apparatus. He plans to use pellets of sodium hydroxide, NaOH(s). (6.6, 7.2) K/U T/I C
    (a) Use what you have learned about chemical reactions to evaluate the chemist's plan. Is it likely to succeed?
    (b) If you believe the plan will work, write and explain the balanced chemical equation for the reaction that will trap the carbon dioxide. If you believe the plan will not work, offer an alternative that includes a balanced chemical equation.

Reflect on Your Learning

97. Explain an idea that you initially had difficulty with but gained an understanding of in this unit. T/I C

98. What did you learn in this unit that was an unexpected finding? C

99. List some topics contained within this unit about which you would like to have had some additional information. T/I C

100. Health Canada is working very actively to lower the average ingestion of sodium by Canadians. The approach includes recommendations for salt cutbacks by food manufacturers and food preparers, education of citizens, and further research into the issue. Personally, are you worried about the quantity of salt in your diet? Explain. C A

Research

GO TO NELSON SCIENCE

101. Hydrocarbons are the main components of gasoline and occur in many molecular forms. Some are isomers of one another, meaning they have the same molecular formula but different molecular structures. Whether they are straight-chain or branched-chain isomers greatly affects how well hydrocarbons burn in car engines. Research the burning of gasoline in cars in relation to hydrocarbon isomers. In particular, find out about the octane rating of a gasoline blend. Write a report of your findings that includes drawings of relevant molecular structures. T/I C A

102. Spectrophotometers are instruments that perform quantitative measurements using various kinds of "light." Research and explain the general concept employed by a spectrophotometer and give an example of a specific type. Include a schematic diagram in your answer. T/I C A

103. Sifto Canada is a major supplier of salt, producing about 7.5 million tonnes annually. Recently, Sifto introduced low-salt and salt-substitute products into the market. Research salt-substitute and low-salt products. Report your findings in the form of a website or pamphlet in which you report on the pros and cons of such products, including specific data about toxicity and other medical risks in comparison to regular table salt. C A

UNIT 4

Solutions and Solubility

OVERALL EXPECTATIONS

- analyze the origins and effects of water pollution, and a variety of economic, social, and environmental issues related to drinking water
- investigate qualitative and quantitative properties of solutions, and solve related problems
- demonstrate an understanding of qualitative and quantitative properties of solutions

BIG IDEAS

- Properties of solutions can be described qualitatively and quantitatively, and can be predicted.
- Living things depend for their survival on the unique physical and chemical properties of water.
- People have a responsibility to protect the integrity of Earth's water resources.

UNIT TASK **PREVIEW**

Hard water can be a nuisance. It can clog pipes and, when mixed with soap, leave a slimy film on laundry and sinks. In the Unit Task, you will design a process to soften water. You will then test the effectiveness of your process by titrating water samples to measure the concentration of the ions responsible for water hardness.

The Unit Task is described in detail on page 498. As you work through the unit, look for Unit Task Bookmarks to see how information in the section relates to the Unit Task.

FOCUS ON **STSE**

"WATER, WATER, EVERY WHERE"

Day after day, day after day,
We stuck, nor breath nor motion;
As idle as a painted ship
Upon a painted ocean.

Water, water, every where,
And all the boards did shrink;
Water, water, every where,
Nor any drop to drink.

This text is taken from "The Rime of the Ancient Mariner," a poem written in 1797–1798 by Samuel Taylor Coleridge. The poet uses the language of his time for the lament of an old sailor during a long and treacherous voyage.

There is no wind and the sailing ship is drifting helplessly. Supplies are running low. The sailor is thirsty. He recognizes the irony that the ship is surrounded by water but he cannot drink any of it because of its high salt content. Seawater is an aqueous solution of dissolved salts and gases. The two most abundant ions in seawater are sodium and chloride, which give it its salty taste.

Like the ancient mariner's ship, the continents of Earth are surrounded by seawater. Unfortunately we cannot use seawater for most of our needs: we require fresh water. As the human population continues to grow, our demand for water grows even faster. Water will be needed not only to grow more food, but also to produce much-wanted energy and manufactured goods. The challenge we face is how to meet this demand for water.

Like seawater, fresh water is not pure. It also contains dissolved substances—but at much lower concentrations than in seawater. Pure liquid water, in fact, is quite rare. This is due to water's ability to react with or dissolve so many substances, which is why it is called the "universal solvent."

Water's role as a universal solvent has important consumer and environmental implications. Water's excellent ability to dissolve substances makes it difficult to completely remove dissolved contaminants from water.

In this unit you will analyze local and global issues related to our use of water. You will also examine the properties and applications of solutions.

Questions

1. What is the origin of some of the pollutants in our water supply?
2. What is the impact of the pollutants in our water supply?
3. Today, consumers can choose their drinking water from a variety of sources. What are the advantages and disadvantages of these sources?
4. The role of a water treatment plant is to make water drinkable but not to purify it. What do you think is the difference between drinkable water and purified water?
5. Some people believe that there is plenty of fresh water on Earth. They argue that some water shortages result from poor management of water rather than its lack of availability. Give an example of a situation where you think this may apply.

CONCEPTS

- chemical and physical properties
- classification of substances
- the states of matter
- the mole and mole quantities
- differences between ionic and molecular compounds
- intermolecular forces

SKILLS

- name or write the chemical formulas of ionic and molecular compounds
- calculate molar masses
- perform stoichiometric calculations
- use solubility tables to predict the solubility of a compound
- research and collect information
- plan and conduct investigations
- interpret and evaluate experimental data
- communicate scientific information clearly and accurately

Concepts Review

1. Explain the meaning of the following terms, using examples from **Figure 1**: matter, pure substance, element, compound, heterogeneous mixture, homogeneous mixture, aqueous solution. K/U

Figure 1 This graduated cylinder contains mercury (bottom), copper beads (middle), and salt water (top).

2. Develop a concept map that shows the relationship among the terms in Question 1. Also include the following terms: mixture, alloy. K/U C
3. What physical property determines the order in which the substances in Figure 1 appear in the cylinder? Explain. K/U T/I
4. Distinguish between molecular and ionic compounds in terms of their
   (a) physical properties
   (b) chemical formulas K/U
5. An aqueous solution is prepared by dissolving a white solid compound in distilled water. The solution is then tested with a conductivity tester, as shown in **Figure 2.** K/U C

Figure 2

   (a) Is this compound most likely ionic or molecular? Explain your answer.
   (b) Predict whether the melting point of this compound will be greater than or less than the melting point of ice. Explain.
   (c) Sketch a model of the entities present in the resulting solution. Explain how your model accounts for the observation in Figure 2.
6. Why are carbon monoxide molecules polar while carbon dioxide molecules are non-polar? T/I
7. (a) List three attractive forces that occur between molecules and one that occurs between ions.
   (b) Rank these attractive forces in order from weakest to strongest.
   (c) Use the concept of intermolecular forces to rank the following substances in order from lowest to highest melting point. Justify your ranking.
   H_2, H_2O, H_2S T/I

8. **Figure 3** shows a stream of water bending toward a positively charged glass rod. K/U T/I
   (a) Explain why this occurs, referring to the bonding in water in your explanation.
   (b) Predict what you would observe if a negatively charged strip were held next to a stream of water. Explain your prediction.

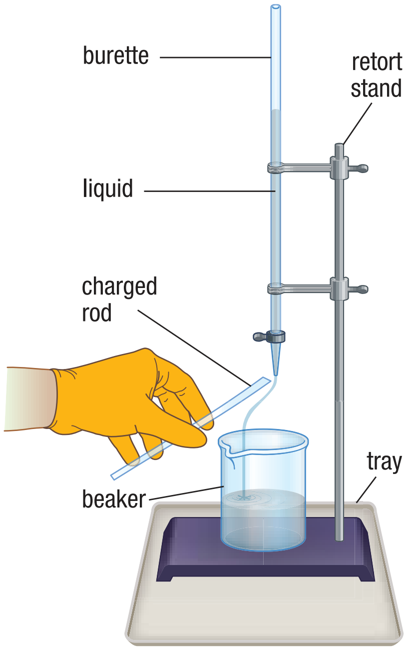


Figure 3 The water is affected by the positively charged glass rod.

Skills Review

9. When solutions of lead(II) nitrate, $Pb(NO_3)_2$, and sodium iodide, NaI, are mixed, a precipitate forms (**Figure 4**). K/U T/I C
   (a) Write a balanced chemical equation for the reaction.
   (b) Use the solubility table in Section 4.6 to identify the precipitate.
   (c) Calculate the mass of sodium iodide required to completely react with 33.12 g of lead(II) nitrate.

Figure 4 Consider what is happening to the ions in solution as this reaction occurs.

10. Name the following compounds: T/I
    (a) K_2S
    (b) $Al(OH)_3$
    (c) NO_2
    (d) P_2O_5
    (e) $PbSO_3$
11. Write the chemical formula for each of the following compounds: T/I C
    (a) potassium nitrite
    (b) iron(III) sulfate
    (c) calcium hydrogen carbonate
    (d) manganese(II) sulfate heptahydrate
    (e) sulfur trioxide
12. Calculate the molar mass of calcium hydroxide, $Ca(OH)_2$. T/I
13. An aqueous solution is tested with pH test paper (**Figure 5**). T/I

Figure 5 pH paper shows the acidity of a solution.

    (a) Estimate the pH of the solution.
    (b) Classify this solution as being acidic, basic, or neutral.
    (c) Which of the following compounds is most likely dissolved in the liquid: $Ca(OH)_2$, KCl, HCl, or $C_6H_{12}O_6$?
14. (a) Write the chemical equation for the reaction between barium hydroxide, $Ba(OH)_2$, and sulfuric acid, $H_2SO_4(aq)$. Your equation should include states.
    (b) Use this equation to show why this reaction is an example of double displacement, precipitation, and neutralization. K/U T/I C

CAREER PATHWAYS **PREVIEW**

Throughout this unit you will see Career Links in the margins. These links mention careers that are relevant to Solutions and Solubility. On the Chapter Summary page at the end of each chapter you will find a Career Pathways feature that shows you the educational requirements of the careers. There are also some career-related questions for you to research.

CHAPTER

8 Water and Solutions

KEY CONCEPTS

After completing this chapter you will be able to

- analyze economic, social, and environmental issues related to drinking water
- explain the challenges that must be overcome in order to have a sustainable supply of fresh water
- distinguish solutions from pure substances or other mixtures
- explain differences in solubility on the basis of intermolecular attractions
- identify the factors that affect solubility
- express the concentration of a solution in a variety of units
- plan and safely perform procedures to make solutions of specified concentrations, both from solids and by dilution

How Can We Ensure a Safe and Sustainable Supply of Drinking Water?

Earth is the ideal distance from the Sun to allow liquid water to exist. It is the only planet in our solar system with water on its surface: not too hot, not too cold, but just right.

Imagine all the water on Earth: salt water, fresh water, water in the atmosphere, and water in the ground. The blue sphere on the opposite page represents all of this water. The vast majority (over 97 %) is salt water. Some of the remaining fraction is part of the gaseous mixture that makes up the atmosphere. Fresh water makes up only a tiny fraction of the blue sphere: about the size of the letter "o." Since we have such a limited quantity of fresh water, it our responsibility to consider the future. We should use water wisely, not polluting or taking too much of this precious resource.

Life as we know it cannot exist without water. Water makes up 55 % to 65 % of your body. It is the key component of fluids that perform essential biological functions. Body fluids transport materials around and out of your body. Fluids also regulate body temperature and facilitate the transmission of nerve impulses. If you lose just 2 % of the water in your body you will probably feel extremely thirsty. A loss of 20 % results in severe dehydration. It could even cause death. That is why it is essential to replenish the water that your body loses each day.

An excellent way to replace your body's lost fluids is to drink clean, fresh water. We call it "fresh water," which might make it sound newly made (like fresh-squeezed orange juice), but most of the water molecules in a glass of water are at least a billion years old. Theories of the origin of water vary. One theory suggests that water formed when certain types of rock in the early Earth decomposed into simpler substances. Another more controversial theory suggests that some water was transported to Earth in comets. A comet is a hard lump of rock and ice, much like a dirty snowball, that orbits the Sun. Biological processes, such as cellular respiration, also produce water. Some of the molecules in your glass of water likely came from this source also.

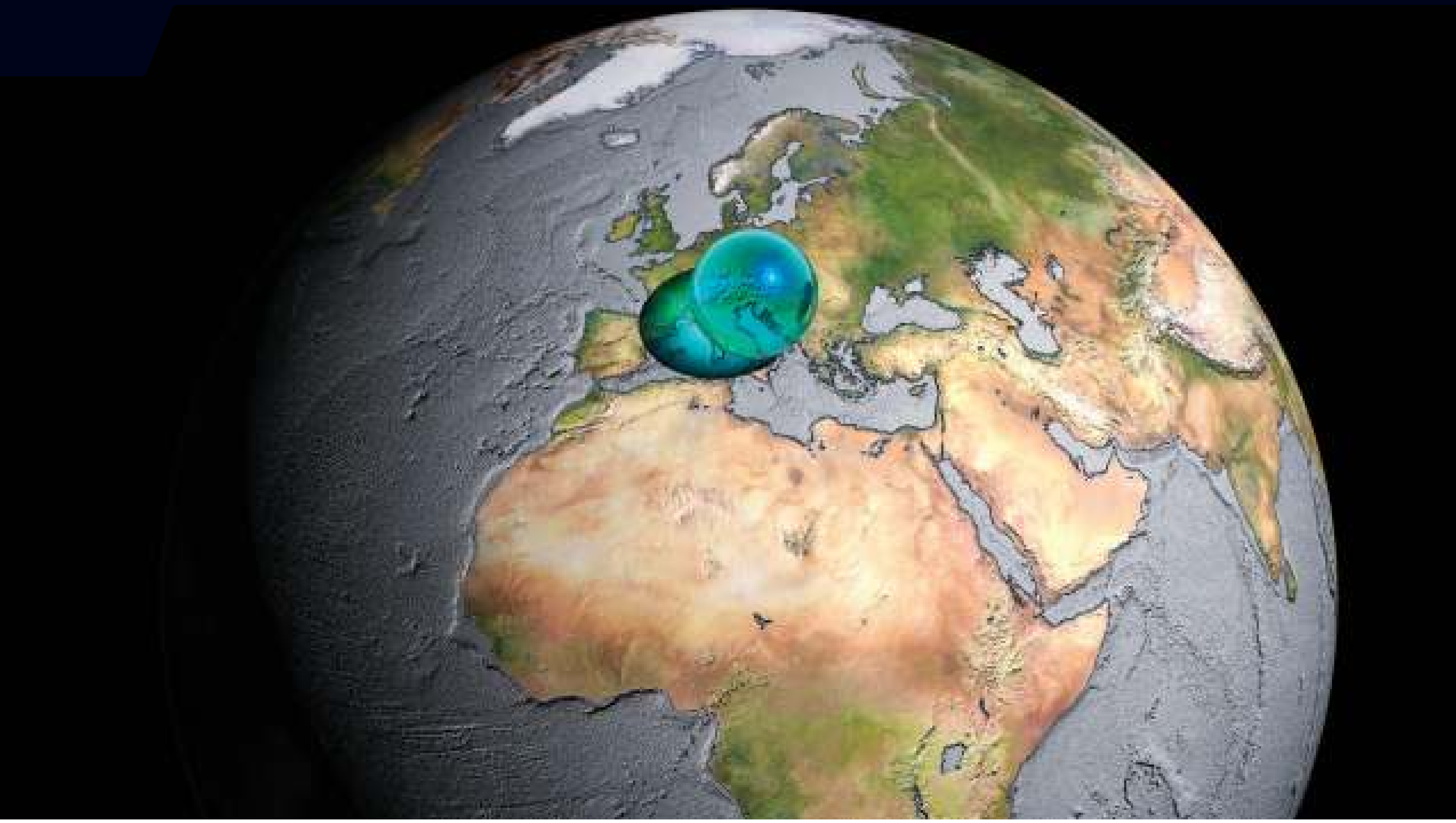
Regardless of its origin, however, there is a limited quantity of water available. As a result, we are just borrowers of Earth's water—all the more reason to take good care of this limited and precious resource.

In this chapter we will explore the many ways we use water, and consider how we could use it more wisely.

STARTING POINTS

Answer the following questions using your current knowledge. You will have a chance to revisit these questions later, applying concepts and skills from the chapter.

1. Why is water such a unique substance?
2. How do human activities contribute to water shortages?
3. How do we treat water to ensure that it is safe to consume?
4. What kinds of substances dissolve in water? What factors determine whether a substance dissolves?
5. List at least five ways in which we can ensure that there is fresh water available for future generations. Include both individual-level and large-scale ideas.



Mini Investigation

Designs and Detergents

Skills: Predicting, Performing, Observing, Analyzing, Communicating

SKILLS HANDBOOK A1.2, A2.4

Detergents are large ionic molecules that can disrupt the hydrogen bonds between water molecules. You might expect detergent to have the same effect on milk. However, milk also contains proteins and fat—long molecules that can twist as they form attractions with detergents. In this activity, food colouring is used to track the movement of fat and proteins.

Equipment and Materials: chemical safety goggles; lab apron; two small plates or Petri dishes; one cotton swab; 4 different colours of food colouring; liquid detergent; milk (2 % or homogenized).

Do not drink or taste anything in the lab.

1. Put on your safety goggles and lab apron.
2. Add water to a plate to a depth of about 1 cm.
3. Add one drop of each type of food colouring to the plate so that the drops are about 4 cm apart.
4. Dip one end of the cotton swab in liquid detergent and keep it there for a few seconds.
5. Predict what will happen when you dip the detergent end of the cotton swab into the middle of the coloured drops.
6. Test your prediction. Dip the swab into the middle and hold it there for about 10 s. Note what happens.
7. Repeat Steps 2 to 6 with a clean plate and milk.

A. Compare the colour changes observed in both plates. T/I

B. What evidence suggests that the molecules within each plate were set in motion by the addition of detergent? T/I

C. What evidence suggests that fat and proteins may be responsible for the observed changes? T/I

D. What role do you think fat and proteins play in producing the observed changes? T/I

E. Suggest how detergents might help to solve an environmental problem. A

F. Mustard powder and egg yolk are two of the many substances that cooks use to help oil and water to mix. These substances are called "emulsifiers." Research other applications of emulsifiers. T/I A

GO TO NELSON SCIENCE

8.1 The Importance of Water

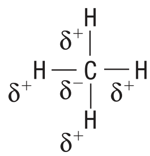
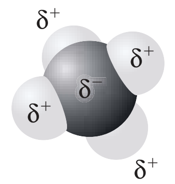
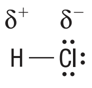
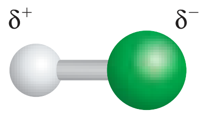
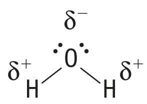
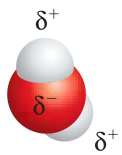
How would you describe water? You might mention how the surface of a lake ripples and reflects sunlight on a summer morning. You could say how cool it feels as you dive in. You could recall the sound of clinking ice cubes as they float in your cold drink. You could discuss water's observed properties in the laboratory: its melting point and boiling point; its appearance; its conductivity; its density as a solid and a liquid; how it interacts with other substances. You could describe its molecular structure: the 2 hydrogen atoms bonded to a single oxygen atom. Water is so familiar to us that we might not realize just how unique a substance it is.

After the oxygen in air, water is the most essential substance we consume. The human body cannot function without it. Water is needed, for example, to transport materials, lubricate tissues, remove waste, and regulate temperature. Water performs these functions because of its unique chemical and physical properties.

Water and Its Properties

Most compounds with molecular masses similar to that of water (methane, for example) are gases at room temperature. Their molecules are small, with relatively few electrons, so their intermolecular forces are weak. Not much energy is required to overcome the forces and allow the molecules to move independently. The fact that water is a liquid at ambient temperatures suggests that its intermolecular forces are unusually strong. Recall that for a compound to exist as a liquid or solid, its intermolecular forces must be strong enough to "stick" the molecules together. The strength or "stickiness" of this attraction increases as the size or polarity of the molecules increases. A methane molecule, for example, consists of a central carbon atom bonded to 4 hydrogen atoms. Since the electronegativity difference between carbon and hydrogen is small, each carbon-to-hydrogen bond is slightly polar. However, the symmetrical arrangement of these bonds around the central carbon atom makes the methane molecule non-polar. In Chapter 3, you learned that the only type of intermolecular attraction between non-polar molecules is the London dispersion force. Since this is the weakest of all intermolecular attractions, small non-polar molecules such as methane are gases at room temperature (**Table 1**).

Table 1 Melting Points of Three Compounds with Small Molecular Masses

Substance	Molecular mass (u)	Melting point (°C)	State at room temperature	Lewis structure	Space-fill model
methane, CH_4	16.05	−182.5	gas		
hydrogen chloride, HCl	36.46	−114.2	gas		
water, H_2O	18.02	0.0	liquid		

Hydrogen chloride is a molecular compound consisting of a hydrogen atom covalently bonded to a chlorine atom. Since chlorine has a greater electronegativity than hydrogen, the chlorine atom attracts the bonding electron pair more strongly. The result is a polar covalent bond and a polar molecule. Recall that the intermolecular attractions between the molecules of a polar compound such as hydrogen chloride are both London dispersion forces and dipole–dipole forces.

In Chapter 3, you also learned that the H–O bond in the water molecule is highly polar. The water molecule is highly polar because of these polar bonds and their asymmetrical arrangement. Hydrogen bonds form between the hydrogen atom of one water molecule and the oxygen atom of a nearby molecule. Hydrogen bonding is a special type of dipole–dipole intermolecular force. Although water molecules also experience London dispersion forces, it is the hydrogen bonding that provides the additional "stickiness" to hold water molecules together as a liquid, up to almost 100 °C.

Figure 1 A dog's upper respiratory tract is moist and lined with many tiny blood vessels. The blood carries thermal energy from around the body to the mouth. Water absorbs a lot of thermal energy from the tissues as it evaporates. Quick, shallow breaths keep air flowing over these moist surfaces and remove the evaporated water.

The Significance of Hydrogen Bonding

Hydrogen bonding accounts for many of the unique physical properties of water. These properties are very significant for life on Earth (**Table 2**).

Table 2 Unusual Properties of Water Resulting from Hydrogen Bonding

Property	Physical or biological significance
high melting and boiling points	• permits water to exist as a liquid at room temperature • keeps body fluids liquid over a large range of temperatures
expansion when cooling from 4 °C to 0 °C	• causes ice to float • causes water to freeze from the top down, allowing life to continue below it
high surface tension	• pulls water into round droplets • allows small insects to "walk on water"
ability to exchange thermal energy with little change in temperature	• enables water to absorb a great deal of thermal energy for a small increase in temperature and to release a great deal of energy for a small decrease in temperature • has a moderating effect on temperature changes in organisms and the environment (**Figure 1**)
inability to mix with non-polar compounds	• enables organisms to retain water because of a waterproof coating (e.g., wax on leaves) • allows organisms to store non-polar substances (e.g., fats, oils)

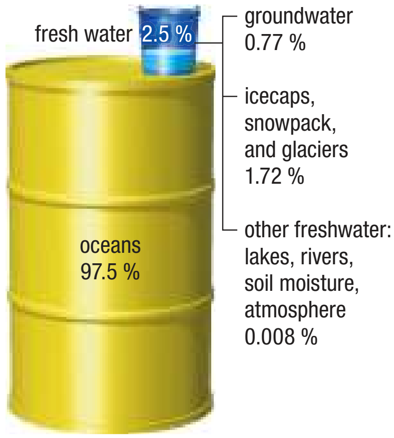


Figure 2 The oceans contain the vast majority of Earth's water. Outside of the oceans, all life depends on the remaining fraction.

Water in Our Environment

About 97.5 % of Earth's water is too salty for consumption. The remaining 2.5 % is fresh water, but much of it is trapped in ice and snow (**Figure 2**). What remains is the tiny fraction of Earth's water that all land plants and animals, and all freshwater organisms, depend on for their survival.

The most accessible water is surface water. Surface water flows in our lakes, rivers, and streams. These sources are fed either by precipitation or by underground springs. Most Canadians—particularly those near the Great Lakes—rely on surface water as our primary source of drinking water.

The supply of water on Earth is finite. This means that any precipitation that falls must ultimately be returned to the sky in a continuous cycle. This is done through the natural processes that occur as part of the water cycle.

WEB **LINK**

To see an animation showing how water flows from one part of the water cycle to another,

GO TO NELSON SCIENCE

water cycle the flow of water on, above, and below the surface of Earth

The **water cycle** describes the movement of water on, above, and below Earth's surface (**Figure 3**). The Sun provides the energy required to keep the water cycle functioning.

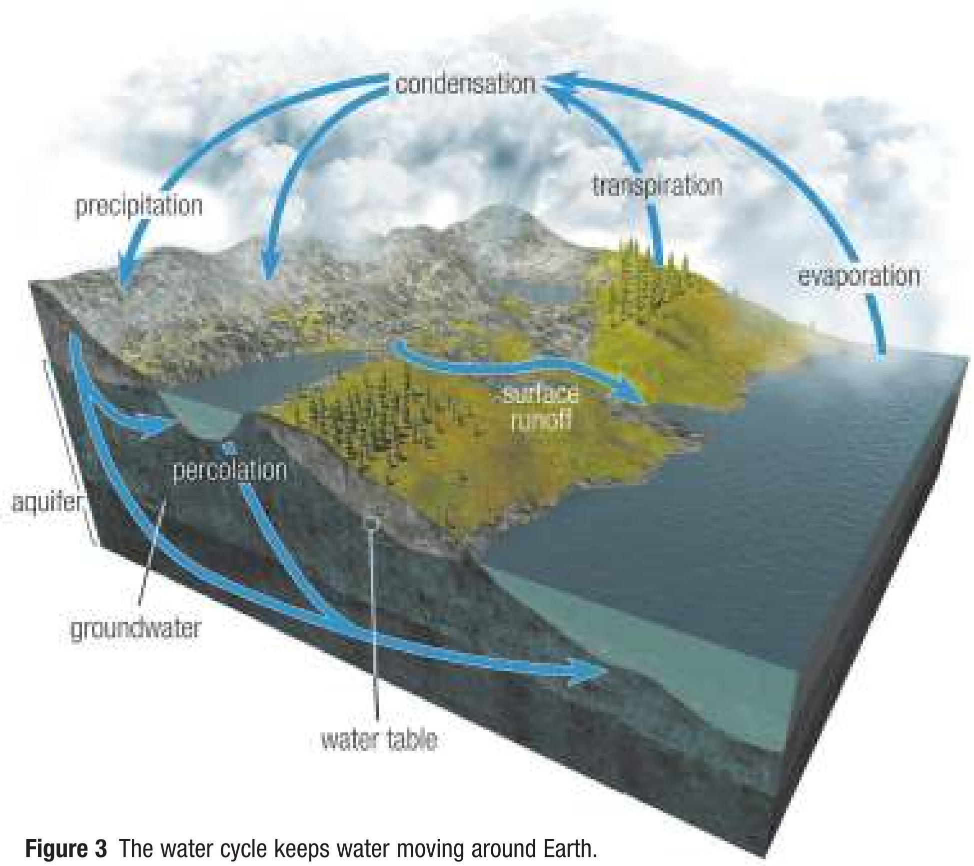


Figure 3 The water cycle keeps water moving around Earth.

transpiration the evaporation of water from the leaves of a plant

aquifer a layer of underground rock that holds a considerable quantity of water; an important source of fresh water

Thermal energy from the Sun first evaporates surface water and sends it into the atmosphere as water vapour. This accounts for as much as 90 % of the water entering the atmosphere. Much of the remainder results from transpiration. **Transpiration** is the process in which water evaporates from tiny pores in plant leaves.

Most precipitation that reaches the ground soaks directly into the soil to become groundwater. As water percolates through the soil it eventually reaches the point where the soil and underlying rock is waterlogged, or saturated with water. The upper limit of this saturated ground is called the water table. The flow of groundwater is influenced by the type of rock it encounters. Highly fractured rock has lots of cracks and holes through which water can slowly move. Saturated rock layers are called aquifers. An **aquifer** is a layer of rock or sediment that holds water and allows groundwater to flow through it. Most rural communities rely on wells drilled into aquifers for their water. Aquifers are an extremely important source of fresh water worldwide. Water in aquifers is about 60 times the volume of fresh water on the surface of Earth. However, aquifers can be drained if we consistently remove more water from them than the natural rate of replenishment.

CAREER LINK

Hydrologists study where and how water moves around underground. This career can involve a lot of fieldwork. If you want to find out more about becoming a hydrologist,

GO TO NELSON SCIENCE

The rate of flow of water from one part of the cycle to another varies considerably. The water in rivers and the atmosphere is continuously replenished. Groundwater can take several years to replace. Replacing the water stored in polar ice caps and glaciers takes much longer. It is important to note that, at any given time, there is far more stored water on Earth than is actually moving through the water cycle.

Water Treatment the Natural Way

potable water water that is suitable for drinking

Several natural processes purify water as it passes through the water cycle. These processes provide a continuous supply of potable water, as long as they are not overburdened. **Potable water** is water that is suitable for human consumption. **Table 3** shows three of the most important processes.

Table 3 Natural Water Purification Processes

Process	Description
evaporation and condensation	As water evaporates, any solutes are left behind. The evaporated water, when it condenses, is very pure.
filtration	As water passes through sand and gravel in the ground, suspended matter is filtered out.
bacterial action	Soil bacteria decompose organic contaminants into simpler, less toxic substances as water percolates through the soil.

Evaporation and condensation are the most significant of these processes. Evaporation from oceans leaves dissolved substances behind, making the remaining water more salty. Similarly, dissolved solutes in plants are left in the leaves after some of the water evaporates during transpiration. In Chapter 5 you learned that certain plants are adept at removing toxins dissolved in groundwater.

The Global Water Crisis

On July 28, 2010, the United Nations General Assembly declared that access to potable water and sanitation is a basic human right. The assembly urged its members and international organizations to support poorer countries in their efforts to provide clean, affordable water and sanitation for all their citizens. The challenges to making this resolution a reality are immense. In 2010, almost 900 million people lacked safe drinking water and another 2.6 billion lacked basic sanitation (**Figure 4**). Furthermore, an estimated 1.5 million children die each year from diseases related to poor water quality and sanitation.

According to Maude Barlow, chair of the Council of Canadians, three of the most important challenges in providing potable water are population growth, increasing global demand for water, and pollution and exploitation of surface water (**Figure 5**, on the next page).

Figure 4 Open sewers like this one in Haiti have resulted in epidemics of water-borne diseases such as cholera.

WEB LINK

The global water crisis is affecting many countries around the world in many different ways. To learn more,

GO TO NELSON SCIENCE

Research This

What Is Your Water Footprint?

Skills: Researching, Predicting, Analyzing, Identifying Alternatives — SKILLS HANDBOOK A5.1

Your water footprint is an indicator of how much water you use. Being aware of your water footprint is an important first step in reducing your consumption of water. Even small reductions in each person's water footprint can result in significant reductions in a population's demand for water.

1. How much water do you think you are responsible for using every year? Predict your annual water footprint.
2. Find a website that tracks water use and calculates the annual water footprint for an individual.
3. Use the website to determine your water consumption over a year.

A. What is your annual water footprint? T/I

B. How did your calculated value compare with your predicted value? T/I

C. Based on this analysis, identify two reasonable lifestyle changes that you could make to significantly reduce your water footprint. A

D. If you live close to a large freshwater source such as the Great Lakes, do you think there is any need to limit your use of water? Is there any other environmental effect of consuming large quantities of water? Explain. A

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Population Growth

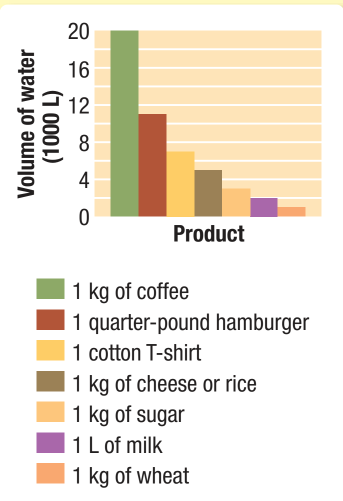
- The world's population may reach 8 billion in your lifetime.
- 80 % of the population will be in developing countries.
- Most people will live in cities.
- Water distribution and sanitation systems will be aging or inadequate.
- More people than ever before will be sharing finite water resources.

Individual Demand

- Individuals, households, and cities account for about 10 % of the demand for water.
- The "per capita" demand for water is greatest in developed countries.

Agricultural Demand

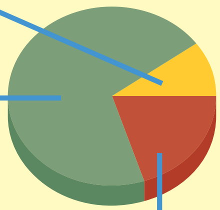
- Agricultural crops use a lot of water, mostly for irrigation. Some agricultural demands are unsustainable. In some agricultural areas, two-thirds of the crops are irrigated with groundwater from rapidly depleting aquifers.
- In some regions, farmers have switched from growing native, water-efficient food crops to growing more financially rewarding crops that require more water. Raising non-native livestock also uses more water.



Pressures on the Water Supply

Increasing Global Demand for Water

- World population tripled in the 20th century, but global demand for water increased sixfold.
- Water is used by individuals, by industry, and by the agricultural sector.



Industrial Demand

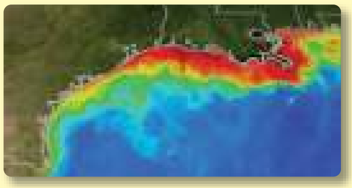
- Industries account for 20–25 % of the total water demand.
- Water is used in almost all manufacturing and power generation processes.
- Increased demand for manufactured goods results in increased demand for water.
- Water used by industry is unavailable for other uses.
- Ideally, industries capture, clean, and reuse water.
- In developing countries, much of the industrial demand for water is to produce goods for export.

Pollution and Exploitation of Surface Water

Water supplies may be contaminated by

- inappropriate use or disposal of industrial and agricultural chemicals
- water passing through mine tailings, becoming acidified, and leaching metals out of rock
- damming rivers for flood control or hydroelectric plants
- diverting water from rivers for shipping or irrigation
- excess fertilizer use for agriculture

Excess fertilizer washes from farm fields into local bodies of water where it stimulates the growth of algae and water plants. When these organisms die their cells decompose, depleting the water of much of its oxygen. The result can be an aquatic "dead zone" where no aquatic life can survive. There is a dead zone at the mouth of the Mississippi River (red area on map).



Producing 1 L of oil from the Alberta oil sands uses about 9 L of water. Oil sands are a mixture of clay, sand, and bitumen. Extraction companies inject large quantities of steam into the ground to soften the bitumen. Much of this water comes from underground aquifers that, once depleted, take decades to replenish. Contaminated wastewater from this process is stored in large lagoons that must be carefully monitored to avoid groundwater contamination.

Figure 5 The endless demand for fresh water is threatening this precious resource.

8.1 Summary

- Hydrogen bonds are a special form of dipole–dipole forces. A hydrogen bond forms between a highly electronegative oxygen atom in one water molecule and a hydrogen atom in another water molecule.
- The water cycle, powered by energy from the Sun, moves water through its various states above, on, and below the surface of Earth. The water cycle cleans the water by removing harmful contaminants.
- An aquifer is a porous underground rock formation that is saturated with water. Aquifers are important sources of fresh water.
- Three factors that contribute to the global water crisis are population growth, increasing demand for water, and pollution of water sources.

8.1 Questions

1. Water, H_2O, and hydrogen sulfide, H_2S, are polar molecules with similar structures. Since oxygen and sulfur are in the same chemical family, you would expect these compounds to have similar properties. Explain why water is a liquid at room temperature while hydrogen sulfide is a gas. K/U
2. What property of water does **Figure 6** illustrate? T/I

Figure 6

3. The high boiling point of water makes it an ideal substance to swim in. Few other widely available substances are liquid at a comfortable temperature! Refer back to Table 2 and, for each of the five properties listed, give an example of why that property is beneficial and unusual. A
4. (a) What is an aquifer?
   (b) What limits the rate at which we can sustainably pump water from an aquifer?
   (c) Why do you think it is important to consider water flow patterns before constructing a mine or an oil refinery in an area? K/U T/I A
5. One part of the water cycle involves evaporation. How does this process result in clean water? K/U
6. Inefficient irrigation methods can waste a great deal of water. Which do you think wastes less water: an in-ground irrigation system or above-ground sprays (**Figure 7**)? Why? T/I A

Figure 7 A spray irrigation system

7. Bottled-water companies are often criticized because of the large volume of water they drain from local freshwater supplies. Some companies respond by saying that soft-drink manufacturers use far more water than the bottled-water industry. Evaluate the criticism and the defence, with reasons. T/I A
8. What effect does changing to a vegetarian diet have on a person's overall water footprint? T/I A
9. What effect do you think global warming might have on the availability of fresh water? Why? T/I A
10. What are five simple changes you can make to reduce your household's overall water consumption? Are you willing to make these changes? Why or why not? T/I A
11. In many cities the infrastructure that distributes clean water to consumers is aging and beginning to fail. In Ontario, local governments control these systems. In other countries, private companies operate the water systems. What do you think are advantages and disadvantages of the privatization of water treatment and distribution? T/I A
12. Historically, settlements and even cities have survived in areas where there is very little precipitation and almost no surface water. Cairo is one example. Research a desert population. Find out how people obtained water in the past, and how they obtain water now. What are the likely environmental effects of this changing practice? T/I A

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8.2 Solutions and Their Characteristics

Figure 1 Water pipes or connectors made of lead can be a source of lead contamination in the drinking water of older homes. The pipes connect the home plumbing system to the public water main under the street.

CAREER **LINK**

A water quality testing technician can work in many environments, from rural outdoor situations to inside large manufacturing facilities. To find out more about this career,

GO TO NELSON SCIENCE

In 2010, the Ontario city of Hamilton issued notices to many of its residents warning of possible lead contamination in their drinking water. Homes built before 1955 were particularly at risk. When these homes were constructed, tradespeople sometimes used lead pipes to link the plumbing in homes to the main water supply under the street (**Figure 1**). Long-term or chronic exposure to lead is a concern because it impairs brain and nervous system development, particularly in children.

How does lead get into the water? If you recall, lead is just above hydrogen on the activity series of metals. This means that lead in the pipes can react with the water, especially if it is slightly acidic. This reaction releases toxic lead ions into the water. The longer the water remains in contact with lead, the greater its concentration of lead ions. As a precaution, Hamilton city officials advised residents to run their taps for five minutes before using the water for drinking or cooking. However, a better way to solve the problem is to remove the source of the contamination: replace the lead pipes.

Tap water is a mixture of a variety of dissolved substances. Some minerals and gases occur naturally in water. Others are intentionally added to improve the quality of drinking water. These include chlorine for disinfection and fluoride compounds to help prevent dental decay. Municipalities conduct frequent tests to ensure that the water coming out of your tap is a clear, colourless solution that is safe to drink.

What Is a Solution?

A solution is a homogeneous mixture of two or more substances. It is a homogeneous mixture because there is only one phase and the components are uniformly mixed, giving a uniform appearance. ("Phase" means "visible part.") As a result, samples taken from two different locations in the solution have exactly the same composition. That is why the first few millilitres of clear apple juice taste just as sweet as the last.

Solutions can be solids, liquids, or gases. Liquid and gaseous solutions are transparent because the entities they contain are too small to block light as it passes through. Solutions may be coloured or colourless depending on the substances they contain.

Let's consider a specific solution: the glucose solution used for intravenous (IV) drips. Glucose (sometimes called dextrose) is a form of sugar. To prepare this solution, a specific mass of glucose, $C_6H_{12}O_6$, is first dissolved in water. After the glucose and water are thoroughly mixed, the grains of glucose are no longer visible (**Figure 2**). At the molecular level, the molecules within each grain have separated and are dispersed evenly throughout the flask. Each molecule is in direct contact with water molecules. Since the mass of these molecules is so small, they will never settle to the bottom of the flask. The diameter of these molecules is also extremely small (in the order of 10^{-9} m), so light can pass directly through the mixture. The result is a clear homogeneous mixture: a solution. This liquid looks just like pure water, but it is different at the molecular level.

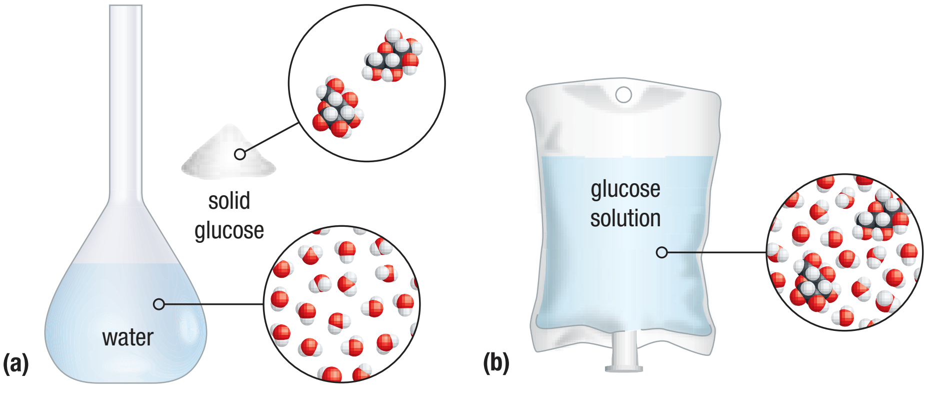

Figure 2 (a) Solid glucose and liquid water are mixed to make a glucose solution for an IV drip. (b) Glucose molecules are evenly distributed in the transparent solution.

Homogeneous and Heterogeneous Mixtures

Solutions are homogeneous mixtures because they are uniform and have only one phase. A **heterogeneous mixture** has two or more phases. Oil and vinegar is an example of a heterogeneous mixture. Even though they are both liquids, oil and vinegar (which is mostly water) separate into two distinct layers or phases. Other mixtures, such as blood and milk, at first appear to be homogeneous, but a closer look reveals that they are not. If you looked at a blood sample under the microscope you would see that it is made up of many different cells suspended in a liquid (**Figure 3**). Its composition is hardly uniform. Blood is a heterogeneous mixture. In fact, all liquid and gaseous mixtures that are translucent (semi-transparent) or opaque (not transparent) are heterogeneous mixtures.

heterogeneous mixture a mixture that contains two or more phases

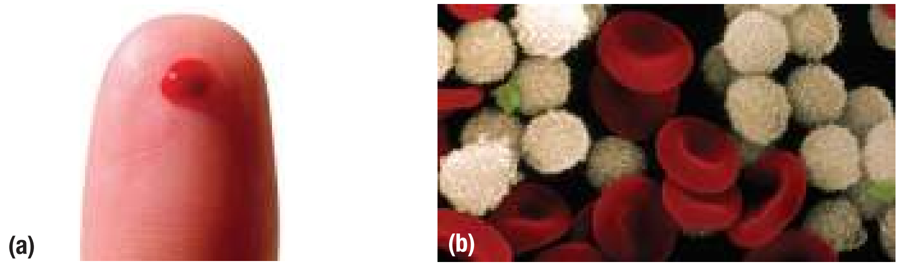

Figure 3 To the unaided eye, a blood sample appears to be homogeneous. Under the microscope, however, blood is a heterogeneous mixture of various types of cells. Blood cells are about 1000 times larger than sugar molecules—large enough to block light from passing through.

Air is an example of a mixture that, depending on its composition, can be either homogeneous or heterogeneous. Dry, clean air is a homogeneous mixture or solution of mostly nitrogen, oxygen, argon, and carbon dioxide. Air is transparent because the gas entities it contains are too small to interfere with the passage of light. The difference in entity size can sometimes be used to distinguish solutions from other mixtures that appear to be homogeneous. The air in your classroom may also appear to be homogeneous. However, when light from a projector passes through it, tiny suspended dust particles in the air deflect the light, making the path of the light beam visible. This is proof that the classroom air is actually a heterogeneous mixture. The same effect can also be seen in liquids (**Figure 4**). We can use this evidence to distinguish true solutions from mixtures that appear to be solutions.

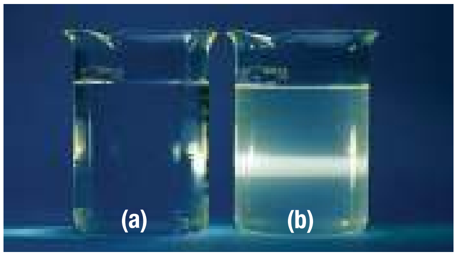

Figure 4 (a) Light passes through the solution but (b) is scattered by tiny suspended entities in a heterogeneous mixture.

Components of a Solution

A solution is a homogeneous mixture of a solute (the substance that is in the lesser quantity) in a solvent (the substance that is in the greater quantity). For the glucose solution in Figure 2, glucose is the solute while water is the solvent. As another example, white rum is a 40 % by volume solution of ethanol (ethyl alcohol). This means that 40 mL of every 100 mL of rum is ethanol. Most of the remainder is water. Therefore, ethanol is the solute and water is the solvent. Alcoholic beverages sold in Canada contain a maximum of 40 % ethanol. In other countries, however, rum may consist of as much as 75 % ethanol. In this case, water is the solute and ethanol is the solvent. Note that a solution may contain more than one solute.

concentration the ratio of the quantity of solute to the quantity of solution or solvent; usually quantity of solute per unit volume of solution

concentrated solution a solution with a relatively large quantity of solute dissolved per unit volume of solution

dilute solution a solution with a relatively small quantity of solute dissolved per unit volume of solution

Perhaps the most important characteristic of solutions is that their composition can change. The term **concentration** is used to describe the ratio of the quantity of solute to the quantity of solution or solvent. A solution that has a relatively high quantity of solute compared to the volume of solution is a **concentrated solution**. A solution with a relatively low quantity of solute compared to the volume of solution is a **dilute solution** (**Figure 5**).

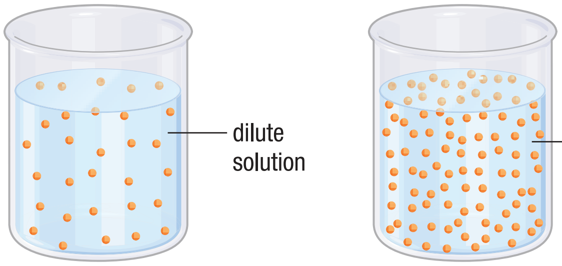

Figure 5 Models of a dilute and a concentrated solution. The concentrated solution contains more dissolved solute per unit volume of solution.

The composition of urine—a solution produced by many animals—varies a great deal. Urine consists of water and dissolved solutes excreted from the body by the kidneys. Changes in the composition of urine can provide important clues that help doctors diagnose medical conditions (**Figure 6**). If you drink plenty of water, normally functioning kidneys produce dilute urine that is pale yellow in colour. If you are dehydrated, the solutes in urine are more concentrated, resulting in a more intense yellow or orange sample. This can also happen if you eat certain foods. People who regularly consume large doses of water-soluble vitamins produce urine that has a larger than normal concentration of these vitamins. Some nutritionists question the value of consuming large quantities of vitamin supplements, suggesting that the body takes only what it needs. The unneeded water-soluble substances are excreted in the urine. Some diseases can also result in the production of strongly coloured urine, or urine with unusual odours.

Figure 6 (a) Urine from a healthy, well-hydrated person is usually pale yellow due to urochrome, a compound produced from the breakdown of old red blood cells. (b) If you are dehydrated, urine is darker in colour because the concentration of urochrome is greater. (c) Orange urine may be a symptom of jaundice, a condition that imparts a yellow colour to the eyes and skin. Jaundice is caused by excess bilirubin, a pigment produced by the gallbladder.

Types of Solutions

There are examples of solutions in different states all around you and within you. You already know that air is a gaseous solution consisting mostly of several gases dissolved in nitrogen. You also know that urine is an aqueous solution of many molecular and ionic compounds.

We have considered several examples of aqueous solutions, but there are many liquid solutions in which water is not the solvent. Gasoline is a solution of mostly liquid hydrocarbons and other solutes. A typical gas station may sell three different grades of gasoline. Each grade has a slightly different composition of solutes, which affects the way the gasoline burns as well as other characteristics. Automobile manufacturers determine the grade that is right for your car.

A solution of two or more metals is called an **alloy**. The gold used for jewellery is alloyed with other metals like silver and copper to stiffen it and produce changes in colour (**Figure 7**). The purity of gold is stated in karats: 24 k gold is pure gold; 14 k gold contains only 58 % gold.

alloy a solution of two or more metals

CAREER LINK

Jewellers are skilled at selecting the appropriate precious metal alloy for each piece they make. They must have a good understanding of the properties of the various alloys. If you are interested in this career,

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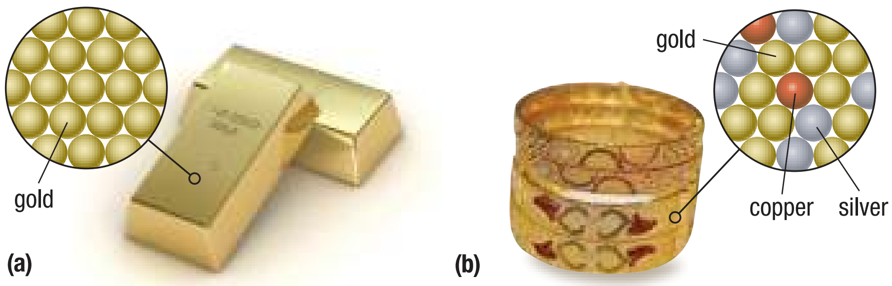


Figure 7 (a) 24 karat gold is pure gold. (b) Most gold jewellery is an alloy of gold and other metals such as copper and silver.

Alloys of mercury are unique because mercury is the only liquid metal at room temperature. As a result, an alloy of mercury is called an **amalgam** to distinguish it from the alloys of other metals. Silver-coloured dental fillings are amalgams of mercury and silver. Mercury also forms an amalgam with gold. If mercury is accidentally spilled on a gold ring it is absorbed by the ring, forming an amalgam. This discolours the gold in the place where the mercury was absorbed.

amalgam an alloy (solution) of mercury with other metals

Solutes and solvents can be solid, liquid, or gas. The different possible combinations of solute and solvent result in a variety of solutions. **Table 1** lists examples of some common solute/solvent combinations. Note that metals only form alloys when they are heated until molten (liquid).

Table 1 Classifying Solutions According to Their Solutes and Solvents

Examples	Original state of solute	State of solvent
air (oxygen, argon, carbon dioxide, and other gases in nitrogen)	gas	gas
carbonated beverages (carbon dioxide and flavour compounds in water)	gas	liquid
humidity (water molecules in air)	liquid or solid	gas
alcoholic beverages (ethanol in water)	liquid	liquid
silver-coloured dental fillings (mercury amalgams)	solid	liquid
air fresheners (vapours from scented solids in air)	solid	gas
clear apple juice (flavour compounds in water)	solid	liquid
brass (an alloy of copper and zinc)	liquid	liquid

Aqueous Solutions

The most familiar types of solutions are aqueous solutions. An **aqueous solution** contains water as the solvent. ("Aqua" is the Latin word for water.) Most of the solutions used in your investigations are aqueous solutions, as are many common consumer products: pop, vinegar, and clear shampoo, for example. All aqueous solutions are transparent. They can be coloured or colourless. Since they are so common and have such important applications, much of this unit will focus on aqueous solutions.

aqueous solution a solution in which water is the solvent

Mini Investigation

When Is a Solution No Longer a Solution?

Skills: Performing, Observing, Analyzing

SKILLS HANDBOOK A1.1, A1.2, A2.4

The addition of acid to thiosulfate ions in water causes these polyatomic ions to decompose, producing sulfur atoms. Over time these atoms clump together to form visible sulfur particles. In this activity, you will use this reaction to distinguish between solutions and other mixtures that appear to be homogeneous.

Equipment and Materials: chemical safety goggles; lab apron; test tube and stopper; scoopula; 10 mL graduated cylinder; test-tube rack; flashlight (optional); 10 mL distilled water; 2–3 crystals of sodium thiosulfate, $Na_2S_2O_3(s)$; dropper bottle containing dilute hydrochloric acid, HCl(aq) (0.1 mol/L).

The hydrochloric acid used in this activity is an irritant. Wash any spills on skin or clothing immediately with plenty of cool water. Report any spills to your teacher.

1. Put on your safety goggles and lab apron.
2. Add 10 mL distilled water to the test tube and place it in the test-tube rack.
3. Put 2–3 crystals of sodium thiosulfate in the water. Stopper the test tube and, with your thumb over the stopper, shake the mixture to dissolve the crystals. Record your observations.
4. Remove the stopper and add 3 drops of hydrochloric acid to the test tube. Replace the stopper.
5. Look for changes in the appearance of the mixture over the next 2 min. Record your observations.
6. (Optional) Use the flashlight to shine a beam of light through the test tube. Record your observations.
7. Dispose of the contents of the test tube as directed by your teacher.

A. Compare the appearance of the mixture in Step 3 with the mixture produced in Step 5. What can you learn, from your observations, about what type of mixture is present? Explain your answer. T/I

B. Compare the sizes of the entities present in both mixtures. What evidence supports your answer? T/I

C. Is the final mixture a solution? Why or why not? T/I

8.2 Summary

- Only one phase is visible in a homogeneous mixture.
- More than one phase is visible in heterogeneous mixtures. Sometimes magnification is required to identify heterogeneous mixtures.
- A solution is a homogeneous mixture of one or more solutes dissolved in a solvent. Solutions can be solids, liquids, or gases.
- The solvent is the component of the solution that is present in the greatest quantity.
- The concentration of a solution is a measure of the quantity of solute per unit volume of solution.
- A concentrated solution contains a relatively large quantity of solute per unit volume of solvent. A dilute solution contains a relatively small quantity of solute per unit volume of solvent.
- Solutions of metals are called alloys.

8.2 Questions

1. Classify the following as solutions, heterogeneous mixtures, or pure substances: K/U T/I
   (a) chlorinated pool water
   (b) sodium chloride
   (c) a jar of jelly beans
   (d) latex paint
   (e) propanone (acetone)
   (f) compressed air in a scuba tank
   (g) chlorine bleach
   (h) a silver-coloured dental filling
   (i) pure liquid honey
   (j) 24 karat gold
2. Use a familiar consumer product as an example to illustrate the difference between the terms "solute" and "solvent." K/U A
3. Identify the solute(s) and solvent in each of the following solutions: K/U T/I
   (a) salt water
   (b) carbonated water
   (c) a 3 % aqueous solution of hydrogen peroxide
4. **Figure 8** shows shafts of light passing through mist. Classify the air in this photo as being a solution or heterogeneous mixture. Justify your choice. K/U T/I

Figure 8 Suspended water droplets and dust particles scatter the light, making shafts of sunlight visible.

5. You are given two clear, colourless liquids. You are told that one of them is an aqueous solution containing a solid solute. The other one is a pure substance. T/I
   (a) Predict what you would observe when the liquids are slowly evaporated.
   (b) Suggest another test that you could perform on the liquids to identify which is the aqueous solution.
6. Why are heterogeneous liquid mixtures translucent or opaque while liquid solutions are transparent? Illustrate your answer with sketches. K/U C
7. Distinguish between an amalgam and an alloy. K/U
8. Water and ethanol are both clear, colourless liquids. When 50 mL of ethanol is combined with 50 mL of water, the final volume of the resulting solution is 96 mL. Assume that no evaporation or spillage occurred. Consider what is occurring at the molecular level when these liquids combine. Suggest an explanation to account for the lower-than-expected volume. T/I
9. What effect do the following changes have on the concentration of a solution? Justify your answer in each case. K/U T/I
   (a) adding more solute
   (b) adding more solvent
   (c) heating the solution to evaporate some of the solvent
   (d) dividing the solution between two different containers
10. Mercury is a highly toxic substance and yet many dentists routinely use mercury amalgams in dental fillings (**Figure 9**). Get the facts! Seek out a few reliable sources and evaluate the dental use of mercury amalgams. Prepare a one-paragraph summary of your evaluation. T/I A

Figure 9 In June 2010, the U.S. Food and Drug Administration (FDA) reopened the debate regarding the safety of mercury amalgam fillings. Are they safe?

11. Gasoline is a common non-aqueous solution. Oxygenates are common additives that are dissolved in gasoline. Research how these compounds make gasoline more green. Give an example of a common oxygenate used in Canadian gasoline. T/I A

GO TO NELSON SCIENCE

8.3 The Dissolving Process

Figure 1 Arctic Fulmar

Toxic bird droppings. That is what researchers discovered as they studied the Arctic Fulmar, a small seabird that nests on Devon Island in the High Arctic (**Figure 1**). The area near the Fulmar nesting areas was rich in nitrogen—as you would expect in an area with bird excrement everywhere. The researchers did not, however, expect their samples to contain toxic industrial chemicals including mercury (from burning coal), PCBs (used in electrical equipment), and DDT (a pesticide). How did these pollutants find their way to the Arctic? The most logical source of the contamination is Fulmar excrement. This led researchers to conclude that the Fulmars' diet of plankton and small fish is tainted with pollutants. These organisms likely absorbed the pollutants from seawater. Researchers also tested large Arctic marine mammals: seals, polar bears, and other animals also contained dissolved pollutants. The highest concentrations of pollutants occurred in the fat of these animals. This evidence suggests that the pollutants are more soluble in fat than in the aqueous fluids in these animals. Moreover, the highest concentration of dissolved pollutants was found in the carnivores at the top of the Arctic food chain (**Figure 2**). This discovery of pollutants in marine mammals is of particular concern because these animals are an important part of the traditional diet of the Aboriginal peoples of the Arctic.

Figure 2 Traces of industrial pollutants have been detected in the fatty tissues of Arctic organisms. Carnivores higher on the food chain showed the greatest concentration of these toxins.

The Dissolving of Ionic Compounds

To understand why a substance may be more soluble in one solvent than another, we should examine the dissolving process at the molecular level. First, we will review the changes that occur as sodium chloride, NaCl, dissolves. In order to dissolve, the ionic bonds within the sodium chloride crystal must be broken. Recall that ionic bonds are so strong that not even a Bunsen burner flame can melt sodium chloride. And yet, sodium chloride readily dissolves in water at room temperature. Furthermore, the resulting solution conducts electricity—proof that ions are released when the sodium chloride dissolves. Why can water break apart a sodium chloride crystal and a Bunsen burner flame cannot?

WEB LINK

To see an animation of the dissolving process,

GO TO NELSON SCIENCE

Recall that water is highly polar. As the molecules approach the crystal, they reorient themselves because the negative (oxygen) end of each molecule is attracted to a nearby positively charged sodium ion. Likewise, the positive (hydrogen) end is attracted to the chloride ion. What happens next is like a tug of war: water–ion attractions are pulling ions away from the crystal while ionic bonds are holding the ions together. If water wins the struggle, as is usually the case with sodium chloride, the crystal dissolves (**Figure 3**). If not, the crystal remains as an undissolved solid.

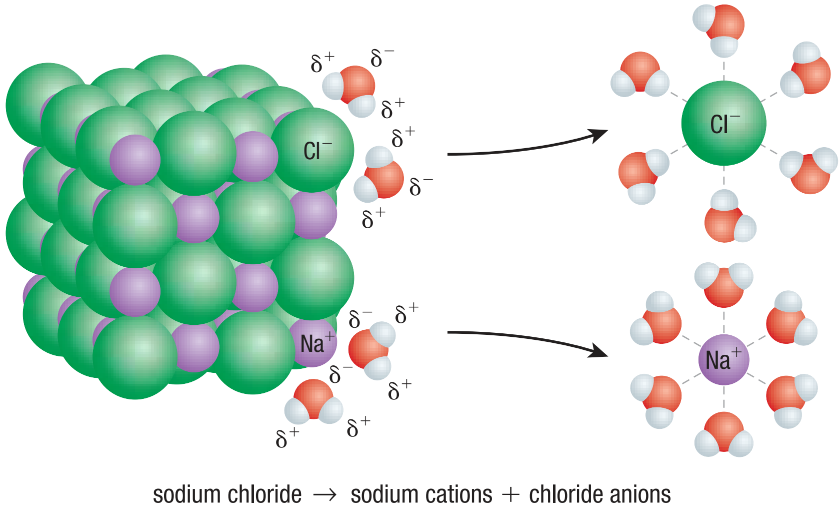


Figure 3 When a sodium chloride crystal is placed in water, water molecules are attracted to the sodium cations and chloride anions. These attractions pull the ions away from the crystal. Once in solution, each ion is surrounded by a layer of water molecules in a process called hydration.

As the ions leave the crystal, they become surrounded by a sphere of water molecules. This process is called **hydration**. Hydration helps to stabilize the ions in the solution, preventing them from attracting each other.

The process in which ions separate from ionic crystals, becoming individual ions, is called **dissociation**. The following equations represent the dissociation of sodium chloride and sodium carbonate:

$NaCl(s) \rightarrow Na^{+}(aq) + Cl^{-}(aq)$

$Na_2CO_3(s) \rightarrow 2\ Na^{+}(aq) + CO_3^{2-}\ (aq)$

The symbol (aq) in these equations indicates that the ions are hydrated, or dissolved in water. Note that water is not included in the chemical equations because it does not undergo a chemical change during the reaction. However, it is necessary for dissociation to occur. Also note that polyatomic ions, such as carbonate, CO_3^{2-}, stay intact rather than breaking apart into their constituent atoms.

hydration the process in which ions are surrounded by water molecules

dissociation the separation of individual ions from an ionic compound as it dissolves in water

Tutorial 1 Writing Dissociation Equations

Here are some points to consider when writing dissociation equations:

- Only ionic compounds undergo dissociation.
- When writing dissociation equations, the solid compound is always written on the left side of the equation and the ions it releases are written on the right.

Sample Problem 1: Writing a Dissociation Equation

Write a chemical equation for the dissociation of potassium phosphate, $K_3PO_4(s)$.

Step 1. Determine the chemical formulas and number of each ion in the compound, keeping each polyatomic ion intact.

$3\ K^{+}$ $\quad PO_4^{3-}$

Step 2. Write the chemical equation (compound on the left; aqueous ions on the right).

$K_3PO_4(s) \rightarrow 3\ K^{+}(aq) + PO_4^{3-}(aq)$

Practice

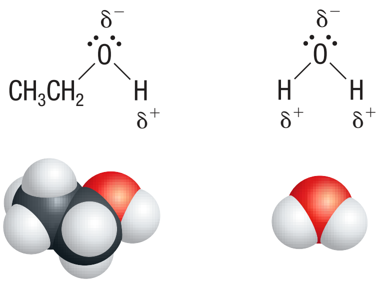
1. Write dissociation equations for the following solid compounds: T/I C
   (a) calcium chloride, $CaCl_2$
   (b) ammonium nitrite, NH_4NO_2
   (c) iron(III) hydroxide, $Fe(OH)_3$
   (d) aluminum sulfate, $Al_2(SO_4)_3$

The Dissolving of Molecular Compounds

miscible able to mix to form a solution; usually describing liquids that mix with each other in all proportions to form a solution

immiscible unable to mix to form a solution; usually describing liquids that do not readily mix

Molecular compounds vary in how easily they dissolve in water. Some, such as glucose, $C_6H_{12}O_6$, and ethanol, C_2H_5OH, mix readily with water to form solutions. Others, like oil, do not. Liquids that mix with each other in all proportions are said to be **miscible**. Liquids that do not mix, like oil and water, are **immiscible**. Why is ethanol miscible in water while most other organic compounds are immiscible? We can explain this difference in miscibility (solubility) by comparing the shapes of ethanol and water molecules (**Figure 4**). Both compounds consist of V-shaped molecules and each of the molecules has a central oxygen atom. Water contains two highly polar O–H bonds while ethanol contains one O–H bond. Therefore, both compounds can form hydrogen bonds with neighbouring molecules. They can also form hydrogen bonds with each other when they are mixed. These attractions between molecules allow water and ethanol to mix to form an aqueous solution (**Figure 5**).



(a) (b)

Figure 4 (a) Ethanol and (b) water are V-shaped polar molecules with a central oxygen atom.

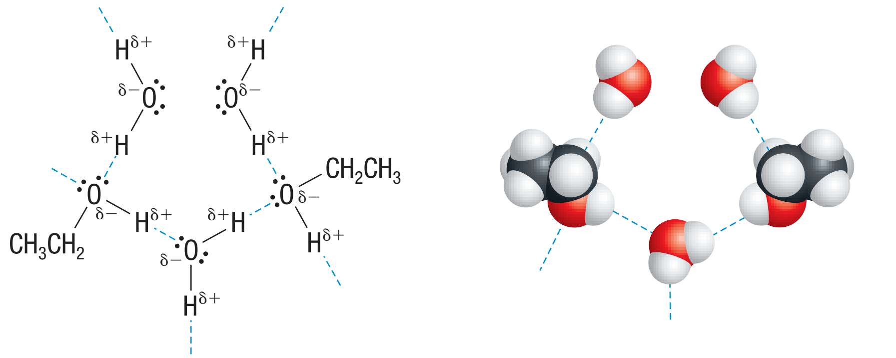


Figure 5 Hydrogen bonding between ethanol and water molecules allows these liquids to form a solution.

The presence of several polar O–H bonds also explains why sugars such as glucose readily dissolve in water (**Figure 6**). The human body relies on this property to function properly. Glucose is the most common energy-rich "fuel" used by body cells. Brain cells rely exclusively on glucose as a fuel. Since these cells are not capable of storing glucose, they rely on the bloodstream to bring them a steady supply. This is why it is essential that glucose is water soluble.

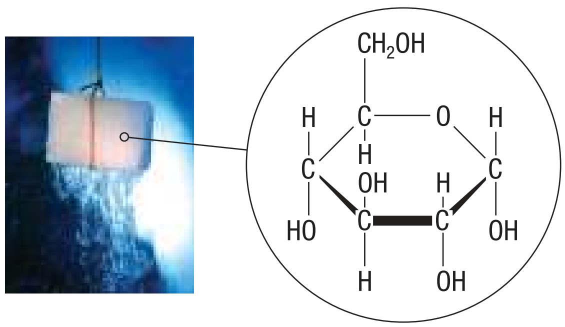


Figure 6 As glucose dissolves, the density of the solution around the cube changes, creating visible ripples in the solution. Glucose is water soluble because of its 5 highly polar O–H bonds. The oxygen atoms, with their partial negative charges, form hydrogen bonds with the hydrogen atoms of water molecules.

Water and Oil Do Not Mix

Although water has been called the universal solvent, there are many substances that it does not mix with. Hydrocarbons, including those found in crude oil, are immiscible with water (**Figure 7(a)**). Crude oil is generally less dense than water, so much of it floats on the surface. This is an advantage during an oil spill. Provided the ocean is calm, some of the oil can be skimmed off the surface (**Figure 7(b)**).

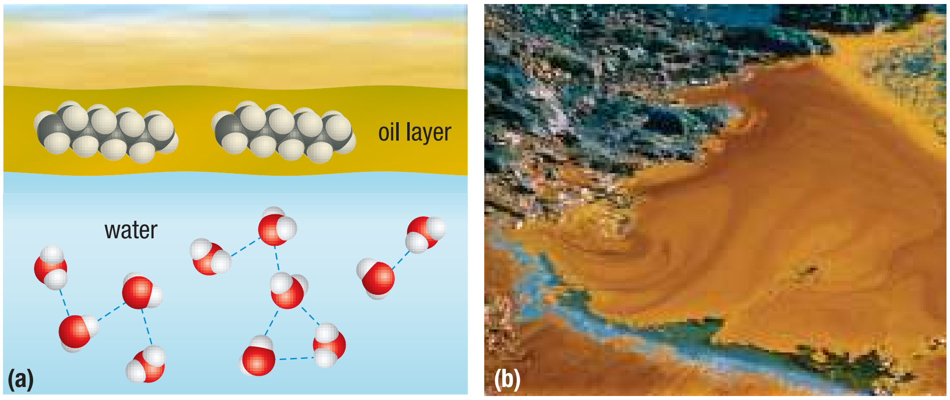


Figure 7 (a) Oil and water are immiscible liquids. The hydrogen bonds between water molecules are too strong to allow oil molecules to penetrate the water layer. (b) When oil is released into the ocean, the less dense parts float on the surface of the water.

WEB LINK

On April 20, 2010, BP's *Deepwater Horizon* oil rig exploded in the Gulf of Mexico. For months, oil leaked into the ocean. To find out more about this oil spill and the clean-up efforts,

GO TO NELSON SCIENCE

To explain why hydrocarbons are immiscible in water, we can examine the structure of a typical hydrocarbon, hexane, C_6H_{14} (**Figure 8**). Its molecules consist of hydrogen atoms connected to a long backbone of carbon atoms.

Recall that the electronegativity of an atom describes its ability to attract electrons within a chemical bond. In Unit 1, you learned that the electronegativity difference between carbon and hydrogen is small. This makes the carbon–hydrogen bond almost non-polar. Recall that molecules that contain only non-polar bonds are always non-polar. (In Chapter 2 we classed them as non-polar molecular compounds.) In addition, many hydrocarbon molecules are symmetrical in shape. This is why hydrocarbons such as hexane are non-polar. The only attractive forces between hydrocarbon molecules are London dispersion forces—the weakest of all intermolecular attractions. However, since hydrocarbon molecules tend to be quite large and regular in shape, these London forces hold the non-polar molecules together quite strongly. However, there are almost no attractive forces between hydrocarbon molecules and water molecules. As a result, water and hexane do not mix.

Some substances are soluble in hydrocarbons. Solid iodine, I_2, for example, readily dissolves in hydrocarbon solvents to give a purple solution (**Figure 9(a)**). This is possible because both solute and solvent consist of non-polar molecules. London dispersion forces loosely attract the iodine molecules to the hydrocarbon, allowing iodine to dissolve.

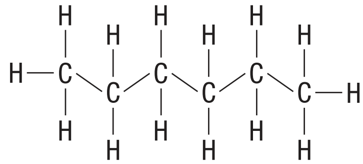


Figure 8 Hexane is one of the hydrocarbons in oil. Like all hydrocarbons, hexane is non-polar.

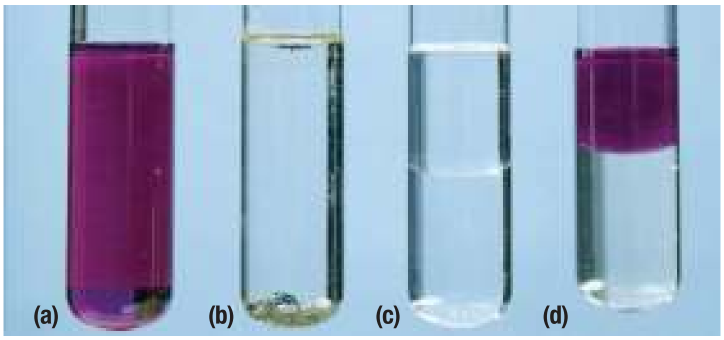


Figure 9 Iodine is soluble in some liquids but not in others. (a) Iodine dissolves readily in hexane to form a purple solution. (b) Iodine does not dissolve readily in water, a polar solvent. (c) Water and hexane form an immiscible mixture. (d) Iodine dissolves much more readily in hexane than it does in water.

CAREER LINK

Toxicology is study of toxins, or poisons, and how they affect the body. To find out about working as a toxicologist,

GO TO NELSON SCIENCE

Now we can explain why industrial pollutants such as DDT are fat soluble. Both DDT and fat are non-polar covalent compounds. Once a non-polar substance, such as a pollutant, dissolves in fatty tissue, it tends to stay there and accumulate. Over the years the compound builds up in the fat of Arctic animals. If DDT were water soluble, it would be flushed out of the body. The same is true for fat-soluble vitamins: they remain stored in our fatty tissue until our bodies break them down.

Like Dissolves Like

In the previous examples, we saw that solutes dissolve when the attractive forces between solute and solvent are stronger than the attractive forces within either the solute or the solvent (**Figure 10**). Ionic compounds and polar covalent compounds dissolve in polar solvents due to the strength of these solute–solvent attractions. Non-polar solutes do not dissolve in polar solvents because the solute–solvent attractions are weak compared to the attractions between solvent molecules. Non-polar solutes do, however, dissolve in non-polar solvents. These solubility generalizations can be summarized by the statement "like dissolves like." In other words, solutes dissolve in solvents of similar polarity.

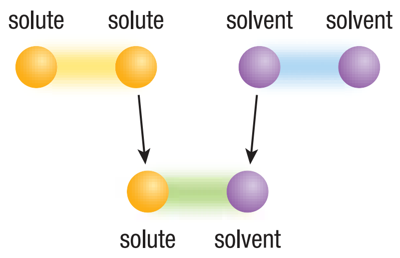


Figure 10 Solute–solvent attractions must overcome the attractions within both solute and solvent in order for the solute to dissolve.

Seeking relief from the heat of eating chili peppers is an interesting application of "like dissolves like". Capsaicin is the compound responsible for the "burning sensation" of a chili pepper. Capsaicin stimulates the cells in the tongue responsible for sensing heat so that the brain interprets that the tongue is burning. If you want quick relief from chili "burn," try drinking milk rather than cold water. The fat in milk and capsaicin are both nonpolar. Fat dissolves capsaicin off the tongue.

Surfactants

surfactant a compound that can reduce the surface tension of a solvent; surfactants have both a hydrophobic part and a hydrophilic part

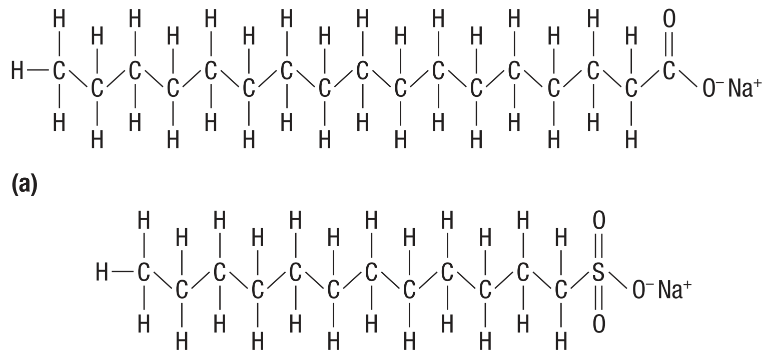
We can "encourage" oil and water to mix with the help of a class of chemicals called "surface acting agents" or surfactants. A **surfactant** is a substance that acts on the surface of a liquid, reducing surface tension. Soaps and detergents are the most familiar surfactants. They reduce the surface tension of water by breaking down the hydrogen bonding network at the surface. This allows a polar solvent such as water to mix with a non-polar liquid like oil. Millions of litres of surfactants were used to disperse the 2010 oil spill in the Gulf of Mexico.

Soaps and detergents are cleaning agents. Soaps are made by reacting animal fats or vegetable oils with a concentrated base. Detergents are made from petrochemicals. The most common types of soap and detergent are very long molecules. A typical soap or detergent is a compound that dissociates in water to form two ions. The cation is usually a sodium ion. The anion is a long structure with a charged "head" and a long, non-polar hydrocarbon "tail" (**Figure 11**).

CAREER LINK

Surfactant scientist are interested in the role of surfactants in many different situations, from oil spills to washing machines. To learn more about this career and the training required,

GO TO NELSON SCIENCE



(a)

(b)

Figure 11 (a) Sodium stearate is used to make bars of soap. (b) Sodium lauryl sulfate is a typical detergent used in toothpaste and shampoo. Both dissolved anions have similar structures: a long hydrocarbon chain with a charged end.

Once a detergent dissolves, the negatively charged end of its anion attracts the polar water molecules while the non-polar hydrocarbon end does not. As a result, we say that the charged end of the anion is hydrophilic or "water-loving." The hydrocarbon end is hydrophobic or "water-fearing." An oily stain on your shirt contains non-polar molecules. When laundry is being washed, the hydrophobic end of the detergent ion attaches to grease. Agitation of the washing machine jostles water molecules back and forth. As they move about, water molecules pull detergent ions along with them. This helps to detach the oil stain from the fabric (**Figure 12**). Once in the water, the oil particle is coated by a layer of detergent ions. This prevents oil and grease from reattaching to the fabric.

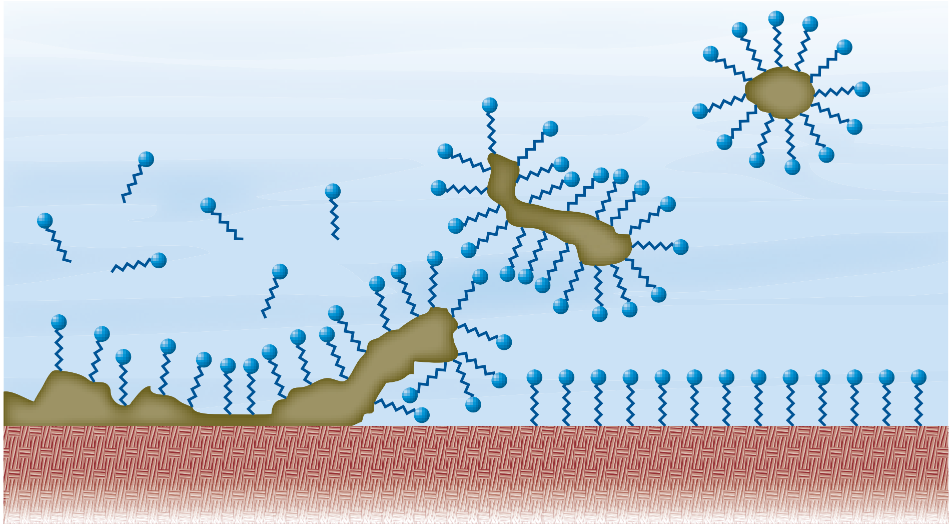


Figure 12 The hydrophobic tails of the detergent ions bond to the grease globule while the hydrophilic ends are attracted to the surrounding water molecules. As the grease is dislodged, it becomes coated with more detergent ions.

Mini **Investigation**

Exploring Selective Solubility

Skills: Predicting, Performing, Observing, Analyzing, Communicating

SKILLS HANDBOOK A1.2, A2.4

At the beginning of this section you read about researchers finding certain compounds stored in the fatty tissue of Arctic mammals. This implies that these compounds are more soluble in fat than in water. In this activity you will use three familiar substances to model this concept. Specifically, you will compare the solubility of iodine, I_2, in two different solvents. Figure 9 shows the purple colour of dissolved iodine, $I_2(s)$. Pure solid iodine, however, is too dangerous to use. Instead, you will use a mixture called Lugol's solution: an aqueous solution of iodine, water, and potassium iodide, KI. The iodide ion, I^-, reacts with iodine, I_2, to form a brown-coloured ion, I_3^-, which readily dissolves in water:

$I_2(s)$	+	$I^-(aq)$	$\rightarrow$	$I_3^-(aq)$
(purple)		(colourless)		brown

Equipment and Materials: chemical safety goggles; lab apron; test tube with stopper; distilled water; mineral oil; dropper bottles of Lugol's solution.

Lugol's solution will stain skin. Any spills on the skin, in the eyes, or on clothing should be washed immediately with cool water. Report any spills to your teacher.

1. Put on your chemical safety goggles and lab apron.
2. Add distilled water to the test tube until it is about one-quarter full.
3. Add an equal volume of mineral oil to the test tube. Allow the two liquids to settle for a few seconds. Record your observations.
4. Add 5 drops of Lugol's solution to the test tube.
5. Stopper and shake the test tube for about 10 s and then allow its contents to settle. Compare the colour of the two layers.
6. Dispose of the contents of the test tube as directed by your teacher.

A. What observation suggests that dissolved iodine was produced during the activity? T/I

B. What evidence suggests that most of the compounds in mineral oil are non-polar? T/I

C. Use the concept of molecular polarity to explain the observed differences in the solubility of iodine. T/I A

Research This

The Search for Greener Solvents

Skills: Researching, Analyzing, Evaluating, Communicating, Defining the Issue, Defending a Decision SKILLS HANDBOOK 5.1

Solvents are the workhorses of many important chemical processes. They dissolve reactants, transport materials from one location to another, and facilitate the separation of one chemical from another. For example, soybean oil is traditionally extracted from soybeans by soaking the beans in the solvent hexane. Since both soybean oil and hexane are non-polar, soybean oil dissolves in hexane. Hexane can be recovered by heating the mixture to boil off the hexane and then condensing the hexane vapour in a process called distillation (**Figure 13**). The liquid remaining in the heating vessel is pure soybean oil. The disadvantage of distillation is that it requires a great deal of energy.

Advances in green chemistry may soon provide a greener way to perform this extraction. Dr. Philip Jessop at Queen's University in Kingston, Ontario is developing solvents that have switchable solubilities. The hydrophobic form of the solvent can be used to extract the oil from the soybeans. The solvent can then be made hydrophilic by bubbling carbon dioxide into it. This causes the solvent and soybean oil to separate into two layers. The oil layer is then drained away. Warming the solvent removes the carbon dioxide, converting the solvent back to its hydrophobic form so that it can be used again.

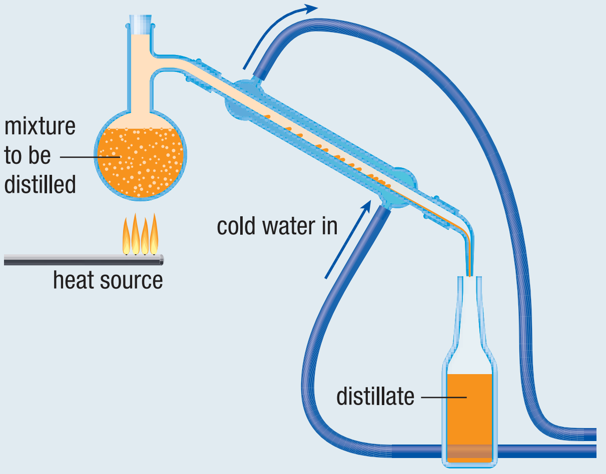


Figure 13 Distillation apparatus

A. Research another application of green solvents. Identify the advantages of using these solvents over traditional solvents used for this process. T/I A

B. Summarize your findings in a brochure that could be used to interest investors in the solvent and its applications. T/I C

GO TO NELSON SCIENCE

8.3 Summary

- For a substance to dissolve, solute–solvent attractions must overcome the attractions within both solute and solvent.
- Ionic compounds dissociate as they dissolve, releasing their ions into solution.
- The symbol (aq) after a chemical formula means that the dissolved entity is surrounded by a sphere of water molecules.
- Solutes dissolve in solvents of similar polarities. ("Like dissolves like.") Ionic and polar covalent compounds are soluble in polar solvents. Non-polar covalent compounds are soluble in non-polar solvents.
- Miscible liquids mix to form a solution.
- Immiscible liquids do not mix.
- Surfactants are compounds that reduce the surface tension of a solvent. Surfactant ions have a hydrophilic end that is attracted to water and a hydrophobic end that is repelled by water.

8.3 Questions

1. The opening paragraph stated that the pesticide DDT was detected in the fatty tissues of Arctic mammals. K/U T/I
   (a) Based on this evidence, is DDT a polar or non-polar substance?
   (b) Predict the relative solubilities of DDT in water and hexane, C_6H_{14}. Justify your prediction.
2. What attractions must be overcome before a solute can dissolve in a solvent? K/U
3. Why is water capable of dissolving sodium chloride but hexane, C_6H_{14}, is not? K/U
4. Water is sometimes called "the universal solvent." Do you think that this is a valid description? Explain your answer. K/U A
5. Why is hydration a necessary part of dissolving? K/U
6. Write a chemical equation for the dissociation of each of the following compounds: T/I
   (a) calcium nitrate, $Ca(NO_3)_2$
   (b) potassium perchlorate, $KClO_4$
   (c) ammonium carbonate, $(NH_4)_2CO_3$
   (d) iron(III) sulfite, $Fe_2(SO_3)_3$
7. Distinguish between the terms "miscible" and "immiscible" using two household examples. K/U
8. Explain what is meant by the expression "like dissolves like." K/U
9. Predict which compound in each of the following pairs is more soluble in water. Why? K/U T/I
   (a) CH_3OCH_3 or CH_3OH
   (b) Na_2CO_3 or CO_2
10. Predict which of these liquids is more miscible with water: methanol, CH_3OH, or dichloromethane, CH_2Cl_2. Why? T/I
11. Iodine, I_2, is a purple solid at room temperature. Corn oil is a yellow liquid. Identical volumes of water, corn oil, and a crystal of iodine are added to the same test tube. The test tube is then stoppered and shaken. Predict the appearance of the test tube after its contents have settled. T/I
12. Lipstick stains on fabric can be difficult to remove (**Figure 14**). Soaking the fabric in water generally does not work. Instead, some people recommend covering the stain with hairspray or petroleum jelly first before washing. Use the concept of polarity to explain why this might work. T/I

Figure 14 Apply your understanding of molecular polarity to help remove this stain.

13. **Table 1** gives the solubilities of several alcohols. The higher the solubility value, the greater the quantity of solute that will dissolve in the solvent.

Table 1 Solubility of Five Alcohol Compounds in Water

Name	Formula	Solubility (g/100 g H_2O at 20 °C)
ethanol	CH_3CH_2OH	miscible
1-propanol	$CH_3CH_2CH_2OH$	miscible
1-butanol	$CH_3CH_2CH_2CH_2OH$	7.9
1-pentanol	$CH_3CH_2CH_2CH_2CH_2OH$	2.7
1-hexanol	$CH_3CH_2CH_2CH_2CH_2CH_2OH$	0.6

   (a) Based on these data, identify one factor that affects the solubility of alcohols in water.
   (b) Predict how the solubility of 1-octanol, $CH_3CH_2CH_2CH_2CH_2CH_2CH_2CH_2OH$, compares with the alcohols in Table 1.
   (c) How could you analyze the data in Table 1 to find the solubility of 1-octanol? Try it. T/I C
14. Describe why the cleaning action of soaps and detergents is an application of "like dissolves like." T/I
15. Vitamin A and vitamin C are both important nutrients. Vitamin A is fat soluble while vitamin C is water soluble. Some people take vitamin supplements to ensure that they are getting enough of both vitamins. Too much vitamin A can be toxic. If you consume it in its natural form, it can also turn your skin orange! Why is it easier to overdose on vitamin A than on vitamin C? A
16. Persistent organic pollutants (POPs) are a class of organic pollutants that adversely affect human health and the environment in many places around the world. Investigate some of the pollutants that are classified as POPs. Research some of their adverse effects. Give three reasons why these compounds are appearing in locations far from where they were originally used. T/I A
17. "We are the land and the land is us. When our land and animals are poisoned, so are we." These are the words of Sheila Watt-Clutier, Canadian member of the Inuit Circumpolar Conference, in 2001. Traces of DDT have been detected in the Arctic, despite the fact that Canada banned its use decades ago. Scientists set about trying to find out how it got there and the potential health effects on the people of Canada's North. What did they discover? How have Inuit responded to the research? Read at least two articles on this topic and write an abstract to summarize each one. T/I C A

GO TO NELSON SCIENCE

Explore an Issue in Solutions

SKILLS MENU

- Defining the Issue
- Researching
- Identifying Alternatives
- Analyzing
- Defending a Decision
- Communicating
- Evaluating

Oil Dispersants—Is the Fix Worse than the Problem?

On April 20, 2010, a massive explosion rocked the *Deepwater Horizon* oil rig that had been drilling for oil beneath the seabed of the Gulf of Mexico. The explosion killed 11 crew members and created a fireball seen 60 km away (**Figure 1**). Two days later the drilling platform sank. Deep below the surface, another disaster was unfolding. The sinking of the platform snapped the pipe that linked the platform to the oil deposit deep below the ocean floor. Oil leaked out of the break at an alarming rate (**Figure 2**). By July 26, 2010, an estimated 4.3 million barrels of oil had escaped—equivalent to about 280 Olympic-sized swimming pools. This was by far the world's worst marine oil spill.

Figure 2 Plumes of oil leaking from the broken wellhead near the ocean floor

Figure 1 Fire crews attempting to extinguish the fire on the *Deepwater Horizon* oil rig, located 80 km off the Louisiana coast in the Gulf of Mexico. The rig had been drilling for oil in the ocean floor, 1500 m below.

Local authorities were concerned that a massive oil slick from the spill could, if it reached land, devastate delicate shoreline ecosystems. To prevent this, crews began deploying large volumes of chemical dispersants, both on the surface and under water (**Figure 3**). The plan was to break up the oil before a large slick could form. Dispersants work much like dishwashing detergents. With the help of wave action, these chemicals break up the oil into smaller droplets that gradually sink to the sea floor. This also makes it easier for naturally occurring micro-organisms in the ocean to decompose the oil into simpler, less toxic substances. The dispersants worked! By August 14, 2010, the environmental impact on the shore was minimal, compared to what it might have been.

Figure 3 Several million litres of chemical dispersants were used on the surface of the Gulf, as well as deep in the ocean, to disperse the oil.

Oil spills are not new, but they are getting larger as our demand for oil increases. Chemical dispersants are not new, either. They have been used successfully to deal with previous oil spills. But never before has such a huge quantity of dispersants—several million litres—been used in one application. Neither have dispersants been used at these depths and pressures. The broken wellhead was 1500 m below the surface. At this depth, pressures are intense. No one really knew how dispersants would respond under these conditions.

The Issue

The dispersants deployed on the Gulf spill may have saved the shoreline from the major catastrophe of a massive oil slick coming ashore. Critics, however, argue that the use of dispersants does not make the oil go away. All it does is push the problem from one ecosystem to another: to the organisms that feed at the ocean floor.

Offshore drilling for oil and natural gas is also a growing industry in many countries, including Canada. Oil companies are already exploring an important oil deposit off the coast of Newfoundland. According to the U.S. Geological Survey (USGS), an estimated 13 % of the world's yet-to-be-discovered oil may lie beneath the Arctic Ocean. Drilling in the frigid Arctic may be next, with all its added challenges.

Your engineering team has been asked to prepare a report for Natural Resources Canada on lessons learned about the large-scale use of chemical dispersants from the Gulf disaster. This report will be used to decide whether such quantities of these chemicals are appropriate for use in Canada.

Goal

To decide whether or not dispersants should be used to treat large-scale oil spills in Canadian waters

Research

SKILLS HANDBOOK A5.1

Work in pairs or in small groups to learn more about the advantages and disadvantages of the use of dispersants in the Gulf of Mexico. Research questions could include the following:

- How do dispersants work?
- Would oil break down in Canadian waters in the same way as in the Gulf of Mexico?
- Is there any evidence of adverse effects of these chemicals on ecosystems?
- What are the unique challenges to drilling for oil and gas in Canada's offshore deposits?

WEB LINK

For suggestions on starting your research into dispersants,

GO TO NELSON SCIENCE

Identify Solutions

As you consider whether or not using dispersants is a good solution, think about these questions:

- Is it appropriate to use chemical dispersants to treat a potential oil spill in Canadian waters?
- If so, what next steps should be taken to ensure these chemicals are readily available where they are needed?
- If not, what other options are there?

Make a Decision

Which solution does your firm recommend? Outline the rationale you used to reach your decision.

Communicate

Complete a report of your findings that can be aired during the science segment of an evening news program on television.

Plan for **Action**

Find out the current status of oil and gas exploration in Canadian waters. What is the position of the current Canadian government on oil and gas exploration in these waters? Do you agree with this position? Write a letter to your Member of Parliament outlining your views on this issue.

GO TO NELSON SCIENCE

Solubility and Saturation

The "phssss" you hear from a freshly opened bottle of pop is escaping carbon dioxide (**Figure 1**). Carbon dioxide is pumped, under pressure, into the pop to give the product its fizziness. Removing the cap releases the pressure and allows some dissolved carbon dioxide to escape. In the next few hours, much of the remaining carbon dioxide also escapes, leaving your pop "flat." Some dissolved carbon dioxide, however, always stays in the drink. A much faster and messier way to produce a flat pop is to add a few mint candies. We will come back to this example shortly.

Figure 1 (a) The fizz from a bottle of pop is caused by the sudden release of carbon dioxide bubbles when the bottle is opened. The release of bubbles proves that the pop contains more dissolved carbon dioxide than it should. (b) Adding mint candies can make the release of gas much more dramatic!

The amount of solute that dissolves in a solvent can have more serious implications. Fish depend on the oxygen dissolved in their lakes and rivers. Without sufficient oxygen, the fish will die. Industrial equipment can be damaged by minerals deposited from water when it is heated. These situations illustrate that several factors determine the quantity of solute that can remain dissolved in a solution.

The Degrees of Saturation

Recall that solubility is the quantity of solute that can dissolve in a given quantity of solvent. Quantitatively, solubility is often expressed as the maximum mass of solute that can be dissolved per 100 g of water at a given temperature. **Table 1** gives the solubilities of some common solutes. According to these data, 204.0 g of sucrose (table sugar) dissolves in 100 mL of water at 20 °C. This is the maximum quantity of sucrose that will dissolve at this temperature.

Table 1 Solubility of Common Solutes in Water at 20 °C

Compound	Solubility (g/100 g H_2O)
calcium carbonate, $CaCO_3$	0.0
calcium chloride, $CaCl_2$	7.5
potassium sulfate, K_2SO_4	12.0
sodium chloride, NaCl	36.0
sucrose, $C_{12}H_{22}O_{11}$	204.0

A solution that contains the maximum quantity of solute at a given temperature is a **saturated solution**. The easiest way to tell whether a solution is saturated is to add more solute. If it is saturated, the added solute remains unchanged: it does not dissolve. A solution that contains less than the maximum quantity of solute is an **unsaturated solution**. More solute added to an unsaturated solution should dissolve. Under special circumstances, you can force a solution to dissolve more solute than it normally would at a given temperature. This solution is a **supersaturated solution**.

saturated solution a solution that contains the maximum quantity of solute at a given temperature and pressure

unsaturated solution a solution in which more solute can dissolve at a given temperature and pressure

supersaturated solution a solution that contains more than the maximum quantity of solute that it should at a given temperature and pressure

Solubility Curves

If you have ever made iced tea you know that sugar is much less soluble in cold water than it is in hot. A convenient way to show the solubility of a substance over a range of temperatures is by using a solubility curve (**Figure 2**).

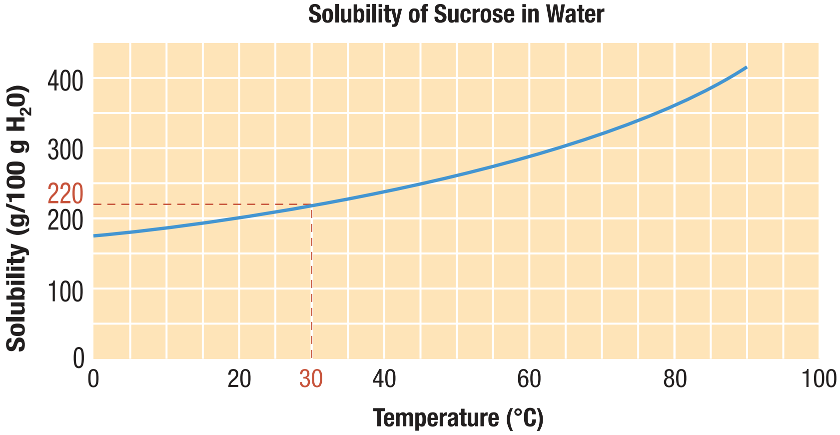


Figure 2 How does the solubility of sucrose (table sugar) change as the temperature increases?

solubility curve a graph of the solubility of a substance over a range of temperatures

A **solubility curve** is a graph of the solubility of a given solute against temperature. Each point along the curve represents the maximum mass of solute that will dissolve at that temperature.

Figure 2 shows that 220 g of sucrose dissolves in 100 g of water at 30 °C. Therefore, a solution at 30 °C that contains only 100 g sucrose/100 g H_2O is unsaturated. This point would be below the solubility curve on the graph. The difference between this point and the curve gives the additional mass of sucrose required to saturate the solution at 30 °C. In this case, an additional 120 g of sucrose is required.

Similarly, a solution containing 300 g sucrose/100 g H_2O at 80 °C is unsaturated. The point on the graph would be below the curve. What do you think will happen if this solution is cooled to 20 °C? According to the graph, when the temperature of the solution reaches 65 °C, the solution should be saturated. Note that the data point (65 °C, 300 g/100 g H_2O) is on the curve. If the solution is cooled further to 20 °C, it becomes supersaturated: the data point (20 °C, 300 g/mL) is now above the curve.

Supersaturated solutions are unstable. Sometimes even a tap on the flask or dropping a single crystal of solute into the solution can initiate crystallization (**Figure 3**). The "seed crystal" provides a surface onto which the excess solute can crystallize.

LEARNING **TIP**

Concentration and Solubility

Concentration and solubility are similar concepts. Concentration describes the quantity of solute per unit volume of solution. Solubility, however, describes the maximum quantity of solute that will dissolve per unit volume at a given temperature. The concentration of glucose (sugar) in a sports drink, for example, may be 6.2 g/mL. But the solubility of glucose in water is 91 g/100 mL at 25 °C. That means that much more glucose could potentially dissolve in a sports drink.

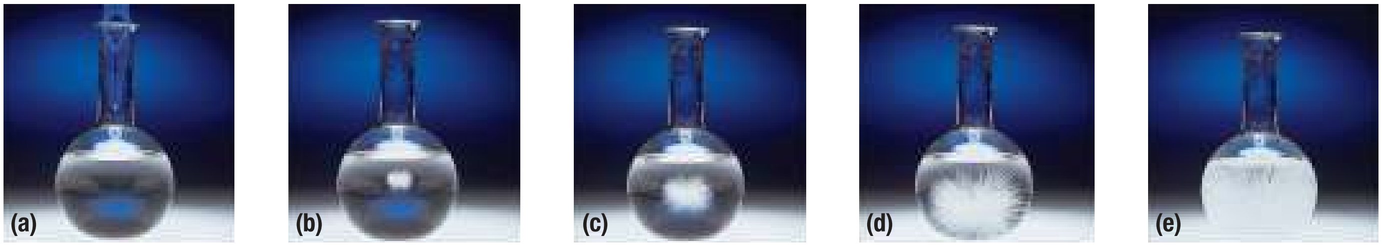


Figure 3 (a) A supersaturated sodium acetate solution is prepared by dissolving sodium acetate in water at a high temperature and then allowing the solution to cool. (b) Adding a seed crystal of sodium acetate causes the excess solute to begin to crystallize. (c)–(e) Crystallization continues until the solution is no longer supersaturated. It is now a saturated solution containing a solid.

A carbonated drink is an example of a solution that is supersaturated with carbon dioxide. Dropping a mint candy into the pop is the trigger that initiates rapid bubble formation—just like adding a seed crystal. The reason this happens is that the surface of the mint is covered with microscopic pits. Each pit is a site where carbon dioxide gas molecules can come out of solution and form bubbles.

Solubility Curves of Ionic Compounds

The solubility of most ionic compounds increases with rising temperature. **Figure 4** shows solubility curves for several common ionic compounds. Note, however, that the solubilities of these compounds are affected by temperature differently. For example, the solubility curve of potassium nitrate, KNO_3, rises more sharply than the solubility curve of potassium chloride, KCl.

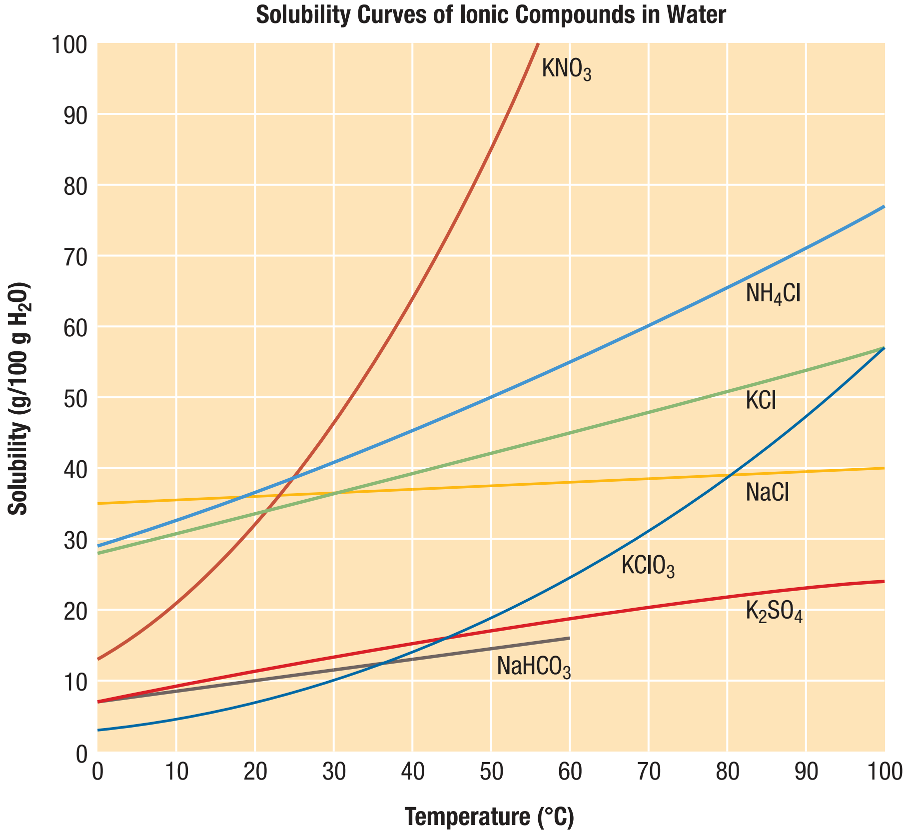
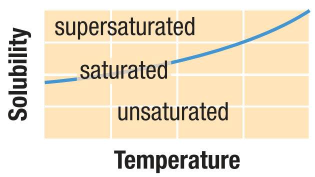


Figure 4 Solubility curves for selected ionic compounds

LEARNING TIP

Interpreting Solubility Curves
Three labels define the different parts of the graph.



Tutorial 1 Interpreting Solubility Curves

The following points summarize the information on a solubility curve:

- Points directly on the curve represent a saturated solution.
- Points below the curve represent an unsaturated solution. The vertical difference between a point below the curve and a point on the curve represents the additional mass of solute/100 g H_2O required to saturate the solution at that temperature.
- Points above the curve represent a supersaturated solution. The vertical difference between a point above the curve and a point on the curve represents the mass of solute/100 g H_2O that will crystallize out of solution at that temperature.

Sample Problem 1: Determining a Solute Mass

SKILLS HANDBOOK A6.4

A potassium sulfate solution, K_2SO_4(aq), containing 11.8 g/100 g H_2O at 20 °C, is warmed to 60 °C. What additional mass of potassium sulfate is required to saturate this solution at 60 °C?

Step 1. Find the data point on the graph. Determine whether the solution is unsaturated, saturated, or supersaturated.

The point (60 °C, 11.8 g/100 g H_2O) is below the solubility curve. The saturation point at 60 °C is 18.4 g/100 g H_2O. Therefore, this solution is unsaturated.

Step 2. Determine the difference between this point and the curve.

18.4 g/100 g H_2O − 11.8 g/100 g H_2O = 6.6 g/100 g H_2O.

Therefore, an additional 6.6 g/100 g H_2O of potassium sulfate is required to make a saturated solution.

Practice

1. (a) Classify a solution that contains 60 g/100 g H_2O of potassium nitrate at 40 °C.
   (b) What mass of solute should crystallize from this solution if it is cooled to 20 °C? T/I [ans: 28 g]
2. What mass of solute is required to saturate a solution containing 10 g/100 g H_2O of sodium hydrogen carbonate, $NaHCO_3$, at 30 °C? T/I [ans: 2 g/100 g H_2O]
3. At what temperature does a solution containing 50 g of ammonium chloride, NH_4Cl, become saturated? T/I [ans: 50 °C]

Solubility Curves of Gases

Unlike with solids, the solubility of gases decreases as the temperature rises (**Figure 5**).

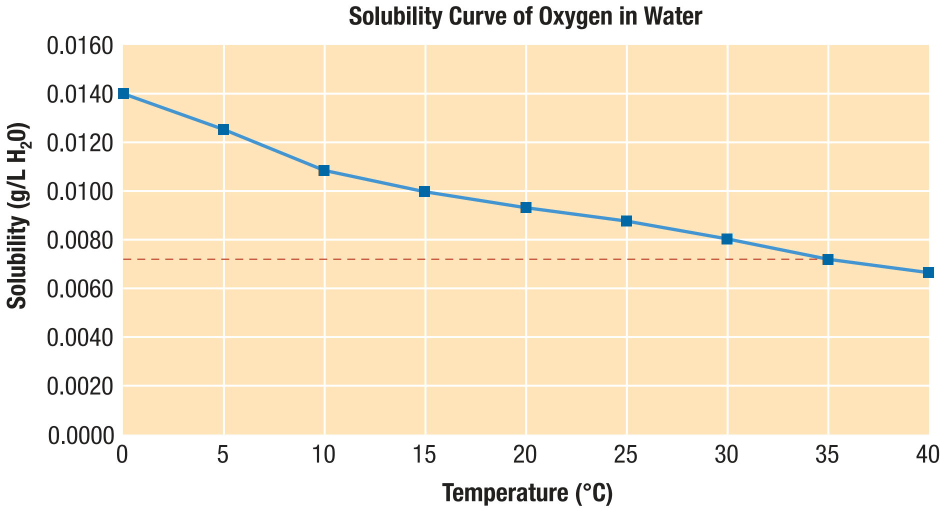


Figure 5 The solubility curve of oxygen. As with other gases, the solubility of oxygen decreases as the temperature rises. Note that water at 0 °C can hold about twice as much oxygen as water at 35 °C.

The trend in solubility for oxygen means that cool water contains more dissolved oxygen than warm water. This has important implications for aquatic animals. The lack of oxygen in warm water can be extremely stressful on fish, equivalent to suffocation. This explains why fish are more likely to be found in deeper, cooler water than near the surface in the summer.

Most aquatic ecosystems can tolerate small increases in temperature over a short period. However, an extended period of hot weather or the release of thermal pollution can reduce oxygen concentrations down to critical levels (**Figure 6**). **Thermal pollution** is excess thermal energy released into water. Thermal pollution can result when warm water used in an industrial process is discharged into the environment. Many manufacturing processes and power plants require large quantities of water for cooling.

THE EFFECT OF PRESSURE ON SOLUBILITY

Pressure is defined as the force applied to a unit area. For example, the force of gravity acting on you exerts a downward force on the ground.

Pressure has little effect on solids and liquids because these states are not very compressible. Gases, however, are significantly affected by changes in pressure. In particular, the solubility of a gas in a liquid increases as the pressure of the gas is increased (**Figure 6**). Soft-drink manufacturers use this property to dissolve far more

thermal pollution an increase in water temperature, usually as a result of warm water being added to an aquatic ecosystem

pressure (*P*) force per unit area; unit is the pascal, unit symbol Pa; 1 Pa = 1 N/m^2

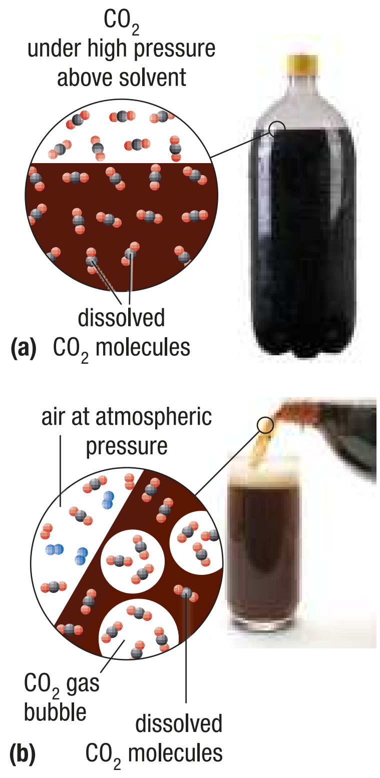


Figure 6 (a) There are no gas bubbles in a sealed bottle of pop because the pressurized gas above the surface keeps carbon dioxide dissolved in the liquid. (b) Opening the bottle reduces the pressure. This decreases the solubility of carbon dioxide, allowing bubbles to form in the liquid.

carbon dioxide into pop than it can hold at normal atmospheric conditions. Chilling the drinks while they are being carbonated also helps more of the gas to dissolve. You will learn more about pressure and gases in Unit 5.

Scuba divers must also be aware of the effects of pressure on dissolved gases. As divers descend, the pressure exerted by the water on their bodies increases. This forces more inhaled gases in the lungs to dissolve in the bloodstream. Later, as the diver ascends, the reduction in pressure causes these gases to come out of solution. The rapid production of gas bubbles results in a condition known as decompression sickness or "the bends." The effect of "the bends" can be quite serious depending on where in the body these bubbles form. Divers can reduce their chances of getting "the bends" by ascending slowly, pausing to allow the gases time to re-dissolve in the blood.

Mini Investigation

Making a Supersaturated Solution

Skills: Performing, Observing, Analyzing, Evaluating, Communicating

SKILLS HANDBOOK A1, A3.1

Recall that hydrates are ionic compounds that contain water molecules within their crystalline structure. For example, sodium thiosulfate pentahydrate, $Na_2S_2O_3 \bullet 5H_2O$, contains 5 water molecules per formula unit of sodium thiosulfate. Heating a hydrate causes these water molecules to evaporate out of the crystal. In this activity, you will use sodium thiosulfate's own water to prepare a supersaturated solution of this compound.

Equipment and Materials: chemical safety goggles; lab apron; large test tube; test-tube rack; scoopula; utility clamp; Bunsen burner clamped to a retort stand; spark lighter; sodium thiosulfate pentahydrate, $Na_2S_2O_3 \bullet 5H_2O(s)$.

This activity involves open flames. Tie back long hair and secure loose clothing. Always be careful when handling materials and equipment that could be hot.

1. Put on your chemical safety goggles and lab apron.
2. Add sodium thiosulfate pentahydrate crystals to a test tube to a depth of about 3 cm.
3. Hold the test tube at an angle with the clamp. Do not direct the opening of the test tube toward anybody.
4. Light the Bunsen burner with the spark lighter. Heat the test tube gently in a Bunsen burner flame, moving the test tube continually.
5. Continue heating until there are no longer any crystals in the test tube.
6. Cool the test tube to room temperature in a stream of cold tap water. Be careful not to agitate the test tube. If your solution crystallizes prematurely, reheat the solution.
7. Place the test tube in a test-tube rack.
8. Add a seed crystal of sodium thiosulfate pentahydrate into the liquid.
9. Carefully touch the sides of the test tube. What do you notice?
10. Return the test tube to your teacher. It can be reused.

A. Write a chemical equation for the reaction that occurred during heating. T/I C

B. Classify the type of reaction that occurred. K/U

C. Use the terms *unsaturated, saturated,* or *supersaturated* to describe the mixture in the tube at the following times: T/I
   (i) prior to the addition of the seed crystal
   (ii) after the addition of the seed crystal

8.5 Summary

- The solubility of a solution is expressed as the mass of solute required to form a saturated solution in 100 g of water at a given temperature.
- Solutions may be unsaturated, saturated, or supersaturated depending on the quantity of solute they hold at a given temperature and pressure.
- A solubility curve shows the solubility of a solute in a specific solvent over a range of temperatures.
- The solubility of solids generally increases as the temperature increases, while the solubility of gases decreases.
- The solubility of a gas increases as the applied pressure increases. Pressure has no significant effect on the solubility of solids and liquids.

8.5 Questions

1. Two identical bottles of carbonated pop are opened. One is placed in the refrigerator while the other is left on the kitchen counter. Which becomes flat first? Why? K/U A
2. Unsaturated, saturated, and supersaturated solutions of an ionic solid look the same. Describe a laboratory test that could be used to distinguish between these solutions. K/U T/I
3. Compare the effect of increasing temperature on the solubilities of sucrose (table sugar) and carbon dioxide. K/U
4. Fresh liquid honey is a saturated sugar solution (**Figure 7**). Over time, crystals of sugar may form in the liquid. K/U A

Figure 7 Grains of crystallized sugar in honey

   (a) What effect would storing honey in the refrigerator have on this process?
   (b) If crystals do form, honey producers recommend placing the jar in a bowl of warm water. Why?
   (c) What is likely to happen if you pour clear liquid honey into a jar that contains a few sugar crystals? Explain.
5. Chemists often use a saturated solution of calcium hydroxide (limewater), $Ca(OH)_2$, to identify carbon dioxide gas. Bubbling carbon dioxide into limewater makes the solution cloudy due to the formation of a calcium carbonate precipitate. **Table 2** provides solubility data for calcium hydroxide. T/I C

Table 2 Solubility of Calcium Hydroxide in Water

Temperature (°C)	Solubility (g/100 g H_2O)
0	0.18
20	0.17
40	0.14
60	0.11
80	0.09

   (a) Plot a solubility curve for calcium hydroxide.
   (b) Compare the trend in solubility of calcium hydroxide with the trends of the other ionic solids given in Figure 4.
   (c) Use your graph to predict the solubility of calcium hydroxide at 100 °C.
   (d) Estimate the solubility of calcium hydroxide at 25 °C and at 50 °C.
   (e) At what temperature is the solubility 0.15 g/100 g H_2O?
   (f) Is a solution containing 0.14 g/100 mL at 0 °C of calcium hydroxide unsaturated or supersaturated? What change is necessary to make the solution saturated?
   (g) If 0.60 g of calcium hydroxide is dissolved in 500 mL of water at 40 °C, will the resulting solution be saturated or unsaturated? Why?
6. Refer to Figure 4 to answer the following questions: T/I
   (a) What mass of potassium chlorate, $KClO_3$, should dissolve in 100 g of water at 60 °C?
   (b) Which substance is least soluble at 20 °C? Which substance is most soluble?
   (c) The solubility of which substance is most affected by temperature?
   (d) What mass of solute can be expected to crystallize when a saturated solution of potassium nitrate, KNO_3, at 50 °C is cooled to 20 °C?
   (e) At what temperature are the solubilities of potassium nitrate, KNO_3, and ammonium chloride, NH_4Cl, equal?
7. How might deforestation at the shore of a river result in thermal pollution? A
8. Why must scuba divers working at great depths ascend to the surface slowly? K/U A
9. Ice cubes made from distilled (or "pure" water) are clearer than ice cubes made with tap water. Distilled water is not completely pure because it still contains dissolved air. Use the trend in the solubility of gases to explain why you can make clearer ice cubes from boiled distilled water than from distilled water at room temperature. K/U T/I A
10. (a) How does the solubility of a gas in water change as temperature changes?
   (b) Name two important dissolved gases in aquatic ecosystems.
   (c) Suggest what effect global warming might have on aquatic organisms. K/U A
11. Gout is a painful form of arthritis caused by the crystallization of uric acid in joints (**Figure 8**). Research the origin of the uric acid, symptoms of this disease, and possible treatment options. A

Figure 8 Gout is affecting the big toe of this person's foot.

GO TO NELSON SCIENCE

8.6 Concentration

Boosting a dead car battery with jumper cables can be hazardous. Conventional car batteries are made up of lead and lead(IV) oxide plates immersed in a concentrated solution of sulfuric acid. Some hydrogen gas is produced during the normal operation of the battery. Hydrogen is highly flammable and can be ignited by a spark. Normally this gas escapes through small vents in the battery. If ventilation is poor, an explosive mixture of hydrogen and air accumulates around the battery. The slightest spark caused by the incorrect use of the jumper cables could initiate an explosion that blows the battery apart, sending concentrated sulfuric acid and casing fragments in all directions. Chemical burns caused by concentrated sulfuric acid are painful and can be disfiguring. Fortunately, the sulfuric acid that you typically use in high school laboratories is much less concentrated than battery acid.

Different industries and organizations measure concentration in different ways. First we will look at the convention used in most laboratories. In Section 8.8 we will explore how concentration is described in other settings.

Amount Concentration

Recall from Section 8.2 that the term "concentration" is used to describe the quantity of solute per unit volume of solution. A concentrated solution contains a relatively large quantity of dissolved solute per unit volume of solution. Conversely, a dilute solution contains a relatively low quantity of solute per unit volume (Section 8.2, Figure 5). For example, the sulfuric acid your school buys is far too concentrated to be used in most investigations. Your teacher will carefully dilute this solution by adding small quantities of it to water. It is then ready for use during your investigations (**Figure 1**).

LEARNING TIP

Diluting Acids
The process of diluting an acid releases so much thermal energy that it can cause water to boil. That is why concentrated acids are always added to water and never the other way around. Water is less dense than acid. If water were added to a concentrated acid, the water would float on top of the acid. The intense heat generated at the boundary between the water and acid causes the water to boil, splattering corrosive acid around. It can even cause the beaker to shatter. Remember: Always Add Acid.

Figure 1 Steam forms when concentrated acid is added to water. When diluting acids, always add acid to water.

amount concentration (*c*) the amount (in moles) of solute dissolved per litre of solution; unit symbol mol/L

Chemists find it most convenient to express solution concentrations in terms of the amount (in moles) of solute per litre of solution. Traditionally, this was called the molar concentration or molarity of the solution, and used the unit symbol "*M*."

According to IUPAC, the governing body that regulates naming conventions in chemistry, the terms "molar concentration" and "molarity" are no longer correct. They have been replaced with the term "amount concentration." The **amount concentration**, *c*, of a solution is determined by dividing the amount (in moles) of solute, *n*, by the volume (in litres) of the solution, *V*:

$$\text{amount concentration} = \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}} \qquad c = \frac{n}{V}$$

The units of amount concentration are mol/L. **Table 1** gives the amount concentration of common stock acid solutions. A **stock solution** is a concentrated solution that is to be diluted to a lower concentration prior to actual use.

Table 1 Amount Concentrations of Common Stock Acid Solutions

Stock acid	Amount concentration (mol/L)
hydrochloric acid, $HCl(aq)$	12
nitric acid, $HNO_3(aq)$	16
sulfuric acid, $H_2SO_4(aq)$	18

stock solution a concentrated solution that is used to prepare dilute solutions for actual use

Tutorial 1 Using the Amount Concentration Formula

When you are given any two values for amount concentration, amount of solute, and volume of solution, use the concentration formula:

$$c = \frac{n}{V}$$

When using this formula always check that:

- the equation is rearranged correctly to solve for the required variables
- amount is expressed in mol
- volume is expressed in L. The most common conversion you will encounter is mL to L and vice versa. Recall that 1 mL $= 1 \times 10^{-3}$ L. For example, 5.0 mL = 0.0050 L or 5.0×10^{-3} L.

After completing a calculation, check that the number of significant digits in the final answer agrees with the least number of significant digits in the calculation. Also check that your final answer is numerically correct and its units are correct.

Sample Problem 1: Calculating Amount Concentration

Sulfuric acid is a solution of hydrogen sulfate, H_2SO_4, in water. When concentrated, it is extremely corrosive (**Figure 2**). A 25.0 mL sample of concentrated sulfuric acid contains 44 g of hydrogen sulfate. Calculate the amount concentration of the solution.

Figure 2 Concentrated sulfuric acid can decompose sucrose (table sugar) into black carbon and water vapour.

Given: volume of solution, $V_{H_2SO_4}$; mass of hydrogen sulfate, $m_{H_2SO_4}$

Required: amount concentration of sulfuric acid, $c_{H_2SO_4}$

Analysis: $c_{H_2SO_4} = \dfrac{n_{H_2SO_4}}{V_{H_2SO_4}}$

Solution:

Step 1. If necessary, convert the volume of the solution to litres. Amount concentration is equal to the amount of solute dissolved *per litre of solution*. Thus, if the volume of the solution is given in units other than litres, you will have to convert the volume into litres using the appropriate conversion factor.

$$V_{H_2SO_4} = 25.0\ \cancel{\text{mL}} \times \frac{1\text{L}}{1000\ \cancel{\text{mL}}}$$

$$V_{H_2SO_4} = 2.50 \times 10^{-2}\ \text{L}$$

Step 2. If necessary, convert the mass of solute to amount of solute. Amount concentration is equal to *the amount of solute* dissolved per litre of solution. Thus, if the mass of the solute is given in the problem statement instead of the amount of solute, then you will have to convert the mass of solute into amount of solute.

$$n_{H_2SO_4} = 44\ \cancel{\text{g}} \times \frac{1\ \text{mol}_{H_2SO_4}}{98.08\ \cancel{\text{g}}}$$

$$n_{H_2SO_4} = 0.4486\ \text{mol [2 extra digits carried]}$$

UNIT TASK **BOOKMARK**

You will use amount concentration values when you perform the analysis involved in the Unit Task on page 498.

Figure 3 The product label on chlorine bleach warns against mixing it with other substances. Toxic fumes may be produced.

Step 3. If necessary, rearrange the concentration equation in the appropriate form. Then substitute in the values, and solve the equation.

$$c_{H_2SO_4} = \frac{n_{H_2SO_4}}{V_{H_2SO_4}}$$

$$= \frac{0.4486 \text{ mol}}{2.50 \times 10^{-2} \text{ L}}$$

$$c_{H_2SO_4} = 18 \frac{\text{mol}}{\text{L}}$$

Statement: The amount concentration of the concentrated sulfuric acid solution is 18 mol/L.

Sample Problem 2: Calculating Volume

Household chlorine bleach is a 0.067 mol/L solution of sodium hypochlorite, $NaClO$ (**Figure 3**). What mass of sodium hypochlorite is required to prepare 225 mL of bleach?

Given: concentration of solution, $c_{NaClO} = 0.067$ mol/L;
volume of solution, $V_{NaClO} = 225.0$ mL

Required: mass of sodium hypochlorite, m_{NaClO}

Analysis: To determine the mass of sodium hypochlorite you first need to find the amount. Use the amount concentration equation to determine the amount of sodium hypochlorite in one litre of solution.

$$c_{NaClO} = \frac{n_{NaClO}}{V_{NaClO}}$$

Solution:

Step 1. Convert the volume of the solution to litres.

$$V_{NaClO} = 225.0 \cancel{\text{mL}} \times \frac{1 \text{ L}}{1000 \cancel{\text{mL}}}$$

$$V_{NaClO} = 0.225 \text{ L}$$

Step 2. Write the appropriate form of the concentration equation, substitute in the values, and solve.

$$n_{NaClO} = c_{NaClO}V_{NaClO}$$

$$= \frac{0.067 \text{ mol}}{1\cancel{\text{L}}} \times 0.225 \cancel{\text{L}}$$

$$n_{NaClO} = 0.01508 \text{ mol } [\text{2 extra digits carried}]$$

Step 3. Convert the amount of solute to mass.

$$m_{NaClO} = 0.01508 \cancel{\text{mol}} \times \frac{74.44 \text{ g}}{1 \cancel{\text{mol}}}$$

$$m_{NaClO} = 1.12 \text{ g}$$

Statement: The mass of sodium hypochlorite required to prepare 225 mL of bleach is 1.12 g.

Practice

SKILLS HANDBOOK A6

1. The food industry uses calcium chloride, $CaCl_2$, to make processed foods, such as pickles, taste saltier without adding sodium chloride. What is the amount concentration of a 750 mL sample of a calcium chloride solution that contains 1.5 mol of calcium chloride? T/I [ans: 2.0 mol/L]
2. Phosphoric acid, H_3PO_4, is the active ingredient in many consumer products designed to remove rust. What is the amount concentration of a solution that contains 63 g of phosphoric acid in 250 mL? T/I [ans: 2.6 mol/L]

3. Fructose, $C_6H_{12}O_6$, is a natural sugar in apple juice. A person with diabetes must be aware of the quantity of sugar she consumes. The amount concentration of fructose in a certain brand of apple juice is 0.67 mol/L. What mass of fructose is present in a 250 mL bottle of apple juice? T/I [ans: 30 g]
4. Soap makers use a 6.0 mol/L solution of sodium hydroxide, NaOH, to make soap. What volume of solution can be prepared using 125 g of sodium hydroxide? T/I [ans: 520 mL]

Preparing Solutions

SKILLS HANDBOOK A3.2, A3.3, A3.7

standard solution a solution for which the precise concentration is known

Many chemical reactions involve aqueous solutions. Often, these reactions are conducted quantitatively. In order to give reliable data, it is important that the solutions be carefully prepared with known concentrations. A solution of a known concentration is called a **standard solution**. In the laboratory, standard solutions are often prepared in volumetric flasks (**Figure 4**). A volumetric flask is a pear-shaped container with an elongated neck. Each size of volumetric flask is designed to contain a certain volume of solution, +/− 0.1 mL at a specific temperature (often 20 °C).

Solutions are prepared by first dissolving a known quantity of solute in about one-quarter of the volume of the solvent, such as water. Then, additional solvent is carefully added until the bottom of the solution's meniscus lines up with the calibration mark on the flask. The final volume of the standard solution prepared in **Figure 5** is 250.0 mL.

Figure 4 Volumetric flasks come in a variety of sizes.

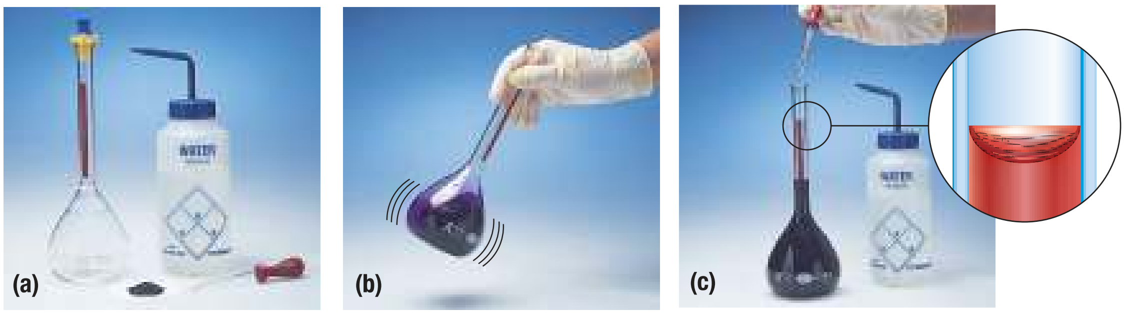


Figure 5 (a) To prepare a 250 mL sample of potassium permanganate solution, you will need a volumetric flask, distilled water, a dropper, and the required mass of potassium permanganate, $KMnO_4$. (b) First dissolve the solid $KMnO_4$ in about 100 mL of distilled water. (c) Use a dropper to add distilled water until the bottom of the meniscus lines up with the calibration mark on the flask.

It may be difficult to transfer the solid directly into the volumetric flask without spillage. To avoid this problem you could first pour the solid into a beaker and add about a quarter of the total volume of water. Pour the solution into the volumetric flask, followed by the water used to thoroughly rinse out the beaker. Finally, add water until the bottom of the solution's meniscus lines up with the calibration mark on the flask.

8.6 Summary

- The concentration of a solution is the quantity of dissolved solute per unit volume of solution.
- Amount concentration is the amount (in moles) of solute dissolved per litre of solution. The units of amount concentration are mol/L.
- Amount concentration is determined using the equation $c = \frac{n}{V}$.
- "Amount concentration" is the preferred IUPAC term for solution concentration (replacing molar concentration and molarity).
- Samples taken from a stock solution are diluted to prepare solutions for use in the laboratory.
- A solution of known concentration is called a standard solution.

Investigation 8.6.1

Preparing a Standard Solution from a Solid (p. 412)
This is your opportunity to learn the skills necessary to make an aqueous solution of known concentration.

8.6 Questions

1. Differentiate between the following pairs of terms: K/U
   (a) concentration and solubility
   (b) a concentrated solution and a saturated solution
2. Briefly describe how to dilute a sample of concentrated acid to about half its original concentration. K/U
3. Silver nitrate, $AgNO_3$, solutions are sometimes used to determine the salt content of food. What is the amount concentration of 500.0 mL of a solution that contains 34.0 g of silver nitrate? T/I
4. A 2.0 mol/L solution of potassium hydroxide, KOH, is used to precipitate toxic metal ions from industrial waste. What mass of potassium hydroxide is required to prepare 1.5 L of this solution? T/I
5. Sodium carbonate, Na_2CO_3, can be used to regulate the acidity of swimming pools. What volume, in millilitres, of a 0.45 mol/L solution contains 35.0 g of sodium carbonate? T/I
6. A solution of sodium oxalate, $Na_2C_2O_4$, is used to analyze for the presence of calcium ions in water. Water quality technicians typically use a 0.100 mol/L solution for this analysis. What volume, in millilitres, can be prepared using 33.5 g of solid sodium oxalate? T/I
7. What mass of copper(II) sulfate pentahydrate is required to prepare 450 mL of a 0.20 mol/L solution? (Be sure to consider the water molecules when you calculate the molar mass of this compound.) T/I
8. Sulfuric acid is one of the most widely used and important industrial chemicals (**Figure 6**). Why do you think it is more economical for companies to purchase concentrated sulfuric acid rather than a safer, more dilute solution? K/U T/I A

Figure 6 The transportation of concentrated sulfuric acid is carefully regulated.

9. A school purchases concentrated solutions of ethanoic (acetic) acid, $HC_2H_3O_2(aq)$, and sulfuric acid, $H_2SO_4(aq)$. The acids are supplied in 1 L containers. The amount concentration of both solutions is 18 mol/L. Which solution has the greatest mass? Why? T/I
10. Cholesterol, $C_{27}H_{46}O$, is a waxy substance found naturally in the body. Under some circumstances cholesterol can clog the arteries. Health Canada recommends that the total cholesterol concentration in the blood not exceed 6.2×10^{-3} mol/L. If you have about 4.7 L of blood, what is the maximum mass of cholesterol that your blood should contain? T/I
11. Consider the design of a 250 mL beaker and a 250 mL volumetric flask (**Figure 7**). Why do you think chemists prefer to use the volumetric flask when preparing solutions of precise concentrations? K/U

Figure 7 Which vessel is better for precisely measuring 250 mL?

12. (a) Why do you think it is important to store standard solutions in sealed containers?
    (b) Predict the effect on the concentration of a sodium chloride solution if it is stored in an open container. Use the equation $c = \frac{n}{V}$ in your answer. T/I A
13. IUPAC makes decisions about the correct terminology that should be used in chemistry. The organization stated that "amount concentration" should be used in place of "molar concentration." What is the benefit of having a body, such as IUPAC, making decisions such as this? What are the drawbacks? Who might be in favour of, and who might oppose, their decisions? T/I A

Preparing Dilutions

Have you ever purchased products, perhaps cleaners or fruit juice, in concentrated form? If so, you probably diluted them before use. This is usually the most economical way to purchase the product because it is in a smaller container (therefore requiring less packaging) and you do not have to pay for transporting the extra water (therefore being cheaper to ship). These factors also make the concentrated product a more environmentally friendly choice over the diluted product.

Solutions that are frequently used in a laboratory are also often purchased in a concentrated form. Not only is this the most economical way to purchase these chemicals, but it also saves valuable storage space in the lab. These solutions are rarely used directly in the laboratory. Instead, they are first diluted with water prior to use in individual investigations.

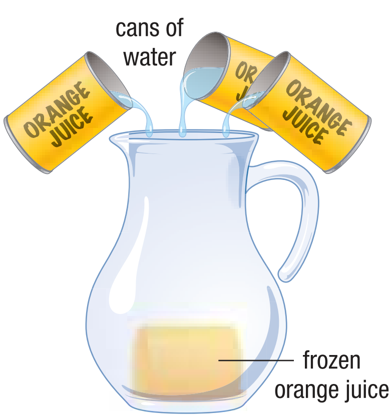


Figure 1 Three cans of water are added to one can of orange juice concentrate. The final volume is four times the original volume. The final concentration is one-fourth the original concentration.

Dilution

SKILLS HANDBOOK A3.3, A3.4

Dilution is the process of reducing the concentration of a solution by adding more solvent. A familiar example of dilution is preparing orange juice from a can of frozen concentrate. This usually involves diluting the concentrate with three cans of water (**Figure 1**). After dilution, the concentrate is diluted to four times its original volume (the original can + 3 cans of water = 4 cans). Therefore, the starting juice concentration is reduced by a factor of 4. Similarly, diluting 1.0 L of a 12 mol/L hydrochloric acid solution to 4.0 L decreases its concentration by a factor of 4. Therefore, the final concentration is 3.0 mol/L. **Figure 2** shows the steps involved in diluting an existing solution of potassium chromate, K_2CrO_4.

dilution the process of reducing the concentration of a solution; usually done by adding more solvent

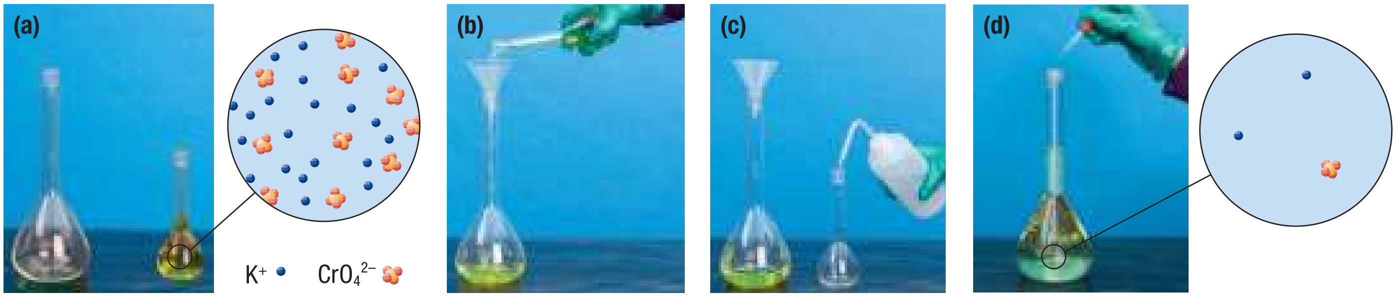


Figure 2 Preparing a dilute potassium chromate solution from concentrate (a) The concentrated potassium chromate solution in the small flask contains a relatively large quantity of dissolved solute per unit volume, giving the solution an intense yellow colour. (b) The concentrated solution is added to the larger volumetric flask. (c) The small flask is rinsed with distilled water and the rinse water is transferred to the larger flask. (d) The concentrated solution is diluted to a larger volume by adding water. The dilute solution contains less solute per unit volume, so the colour is less intense.

Volumetric Glassware

Dilutions often involve transferring precise volumes of a concentrated solution into another container. You have probably used a graduated cylinder to measure and transfer solutions. However, a graduated cylinder is not precise enough for analytical work involving small volumes. Instead, pipettes are used. There are several types of pipettes (**Figure 3**). A volumetric pipette is used to deliver a fixed volume of solution. For example, a 10 mL volumetric pipette delivers 10.00 +/− 0.02 mL of solution. A graduated pipette has volume markings or graduations, much like a graduated cylinder. This type of pipette can deliver a range of volumes, from 0.1 mL to 10.0 mL. Note that the graduated pipette is not as precise as a volumetric pipette.

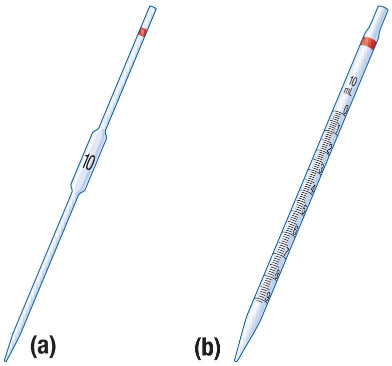


Figure 3 (a) A volumetric pipette can deliver only one specific volume. (b) A graduated pipette can deliver a range of volumes.

Dilution Calculations

Many dilute solutions can be prepared by simply applying some common sense. For example, doubling the volume of the solution by adding water decreases its concentration by a factor of 2. Similarly, increasing the volume by ten times decreases the concentration by a factor of 10.

Investigation 8.7.1

Preparing a Standard Solution by Dilution (p. 413)
In this investigation you will learn the skills necessary to dilute an aqueous solution of known concentration to make a less concentrated solution.

For less obvious situations, we need to develop a mathematical equation that can be used to determine the final concentration of the dilute solution. During dilution, the amount (in moles) of solute present does not change. The concentration changes because the volume of the solution increases (**Figure 4**).

$$\underset{\text{decreases (because volume increases)}}{c\downarrow} = \frac{\overset{\text{remains constant}}{n}}{\underset{\text{increases (because water is added)}}{V\uparrow}}$$

Figure 4 During a dilution, the concentration of a solution decreases due to an increase of its volume. The amount of solute remains unchanged.

The amount of solute in a solution is given by $n = cV$. We can use subscripts to differentiate between the two solutions: subscript "c" represents the initial concentrated solution, and subscript "d" represents the final diluted solution. We can therefore describe the amount of solute in the initial concentrated solution by the equation

$$n = c_c V_c$$

The value of n, the amount of solute, remains unchanged after dilution. The amount of solute in the dilute solution is given by

$$n = c_d V_d$$

Since n remains constant, we can combine the two equations.

The Dilution Equation

$$c_c V_c = c_d V_d$$

LEARNING TIP

Dilution Equation
$c_c V_c = c_d V_d$
c_c is the concentration of the concentrated solution
V_c is the volume of the concentrated solution
c_d is the concentration of the dilute solution
V_d is the volume of the dilute solution

Tutorial 1 Dilution Calculations

Solve dilution problems using the strategy for solving for an unknown variable.

Sample Problem 1: Determining Amount Concentration

What is the final concentration when 250 mL of a 16.0 mol/L nitric acid solution is diluted to 4.5 L?

Given: concentration of initial concentrated solution, c_c = 16.0 mol/L; volume of initial concentrated solution, V_c = 250 mL; volume of final diluted solution, V_d = 4.5 L

Required: concentration of diluted solution, c_d

Analysis:

$c_c V_c = c_d V_d$

Solution:

Step 1. If necessary, convert the volume of solution to litres.

$$V_c = 250\ \cancel{\text{mL}} \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}}$$

$$= 0.250\ \text{L}$$

Step 2. Rearrange the concentration equation in the appropriate form. Then substitute in the values and solve the equation.

$$c_d = \frac{c_c V_c}{V_d}$$

$$= \frac{16.0\ \frac{\text{mol}}{\text{L}} \times (0.250\ \cancel{\text{L}})}{4.5\ \cancel{\text{L}}}$$

$$c_d = 0.89\ \frac{\text{mol}}{\text{L}}$$

Statement: The final concentration of the nitric acid solution is 0.89 mol/L.

Practice

SKILLS HANDBOOK A6

1. What is the final concentration when 4.5 L of 3.0 mol/L car battery acid is diluted to 20.0 L? T/I [ans: 0.68 mol/L]
2. What volume of a 6.0 mol/L sodium hydroxide solution is required to produce 850 mL of a 0.100 mol/L solution? T/I [ans: 14 mL]
3. A teacher needs to prepare 4.5 L of 1.0 mol/L ethanoic acid (acetic acid) for an investigation. What volume of concentrated 17.4 mol/L ethanoic acid is required? T/I [ans: 0.26 L or 260 mL]
4. 250 mL of a 0.100 mol/L iron(III) nitrate solution is prepared using 75 mL of a stock solution. What is the amount concentration of the stock solution? T/I [ans: 0.33 mol/L]

LEARNING TIP

Dilution Factors

A quick way to determine the final concentration of a dilute solution is by multiplying c_c by a dilution factor. The dilution factor is the volume of the concentrated solution divided by the volume of the dilute solution. In Sample Problem 1 the dilution factor is 0.250/4.5 (see Step 2). Chemists routinely use dilution factors when preparing solutions.

8.7 Summary

- Dilutions involve adding more solvent to a solution to decrease the concentration.
- The dilution equation, $c_c V_c = c_d V_d$, is useful when solving dilution problems.
- Volumetric glassware is required to prepare dilutions.

8.7 Questions

1. Describe two applications of the use of dilutions at home. A
2. 25 mL of a 0.6 mol/L cobalt(II) chloride solution is added to each of three beakers. No water is added to the first beaker, 25 mL of water is added to the second beaker, and 50 mL of water is added to the third beaker (**Figure 5**). T/I

Figure 5

(a) What volume of solution is in each of the beakers?
(b) Why is there a difference in the intensity of the colour of the solutions?
(c) Compare the amount of solute in the three beakers.

3. How does the number of entities in a solution change when more water is added to the solution? K/U
4. Describe how to prepare a 0.25 mol/L solution of hydrochloric acid from a 1.0 mol/L solution. T/I
5. What effect do the following changes have on the final concentration of a solution? T/I
   (a) doubling the original volume by adding water
   (b) doubling the original volume by adding more of an identical solution
   (c) increasing the original volume by a factor of 10 by adding water
6. How would you prepare 250 mL of 0.40 mol/L sodium chloride solution, using an available 1.5 mol/L solution? T/I
7. An experiment requires 100 mL each of 0.50 mol/L HCl(aq), 0.25 mol/L HCl(aq), and 0.1 mol/L HCl(aq). How could you prepare these solutions using an available 2.0 mol/L hydrochloric acid solution? T/I
8. Millions of tonnes of concentrated chemicals such as sulfuric acid are transported across North America annually. Why can the transport of chemicals in concentrated form be more environmentally friendly than the transport of the diluted form? What are the environmental risks?
9. Dilution was "the solution to pollution" for much of human history. Research the meaning of this expression. Why is dilution no longer a viable option for disposing of the pollution we generate? A

GO TO NELSON SCIENCE

8.8 Concentrations and Consumer Products

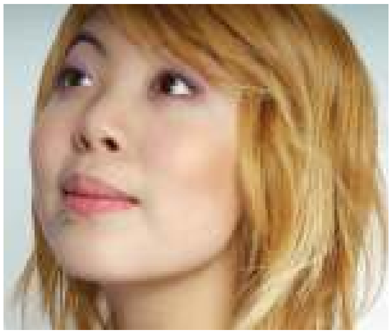


Figure 1 Hydrogen peroxide "lifts" or removes hair colour. An experienced hair stylist can predict which concentration of hydrogen peroxide will lift just the right amount of colour.

Lightening your hair colour using a kit requires skill, patience, and a lot of chemistry. One of the key ingredients in most hair-colouring kits is hydrogen peroxide. This compound works by reacting with the natural pigment in a process that hairstylists call "lifting." The amount of colour "lifted" or removed from your hair depends on the concentration of hydrogen peroxide and the length of time it is in contact with your hair. A concentrated peroxide solution may lift the desired amount of colour in only 25 min. But it can also damage your hair. A less concentrated solution may take longer but could be less damaging (**Figure 1**).

The concentration of consumer products, like hydrogen peroxide, is usually indicated on the product label. Concentration may be expressed in a variety of ways. For example, hairstylists use "20 volume" and "30 volume" hydrogen peroxide to lift hair colour. Stylists know that 30 volume works faster: it is more concentrated. The term "volume" comes from the fact that hydrogen peroxide decomposes to release oxygen gas. A solution of 20 volume hydrogen peroxide, for example, releases 20 times its volume in oxygen gas. The use of "volume of oxygen" to express concentration is unique to hydrogen peroxide. A more common method used to express concentrations of consumer products is percentage concentration.

Percentage Concentration

In Section 8.6 you used an equation to calculate amount concentration. A more general equation for the concentration, c, of a solution is

$$c = \frac{\text{quantity of solute}}{\text{quantity of solution}}$$

The labels of many consumer products, including hydrogen peroxide and vinegar, give concentrations expressed as a percentage concentration (**Figure 2**). As you will see, percentage concentration can involve volumes of solutes and solution, or their masses, or both.

Figure 2 The concentrations of common consumer products are often expressed as percentages.

Percentage Volume/Volume (% V/V)

The label on the bottle of isopropyl alcohol in Figure 2 indicates that the liquid is 50 % isopropyl alcohol by volume. This means that 50 mL out of every 100 mL of solution is the solute: isopropyl alcohol (2-propanol). The percentage volume/volume, % V/V, of the solution is:

$$\begin{aligned} c_{\text{v/v}} &= \frac{50\ \cancel{\text{mL}}}{100\ \cancel{\text{mL}}} \times 100\ \% \\ &= 50\ \% \end{aligned}$$

In general,

$$c_{\text{v/v}} = \frac{V_{\text{solute}}}{V_{\text{solution}}} \times 100\ \%$$

As you might expect, this form of expressing concentration is commonly used when both the solute and solution are liquids.

Percentage Weight/Volume (% W/V)

The concentrations of consumer products may also be given in terms of a percentage weight/volume, % W/V or $c_{\text{W/V}}$. For example, the product label of a "20 volume" hydrogen peroxide bottle may indicate 6 % W/V. This means that 6 g of every 100 mL of the solution is the solute: hydrogen peroxide.

$$\begin{aligned} c_{\text{w/v}} &= \frac{6\ \text{g}}{100\ \text{mL}} \times 100\ \% \\ &= 6\ \% \end{aligned}$$

In general,

$$c_{w/v} = \frac{m_{solute}}{V_{solution}} \times 100\ \%$$

The term "weight" on a product label actually refers to the mass of the solute.

Percentage weight/volume is often used to describe the concentration of household mixtures such as hydrogen peroxide solution and chlorine bleach.

Percentage Weight/Weight (% W/W)

The concentrations of some medications and concentrated acids are often expressed as a percentage weight/weight, % W/W or $c_{W/W}$. For example, benzoyl peroxide is an active ingredient in many acne creams (**Figure 3**). According to the product label, its concentration is 5 %. The concentration of this product is usually expressed as % W/W. This means that about 5 g of every 100 g is benzoyl peroxide.

$$c_{w/w} = \frac{5\ \cancel{g}}{100\ \cancel{g}} \times 100\ \%$$

$$= 5\ \%$$

In general,

$$c_{w/w} = \frac{m_{solute}}{m_{solution}} \times 100\ \%$$

Percentage weight/weight is often used for mixtures of solids in either liquids or other solids.

Tutorial 1 Solving Problems with Percentage Concentrations

Follow these useful tips when solving problems with percentage concentrations.

- All units of measured quantities must be identical (except in the case of $c_{W/V}$).
- If necessary, use conversion factors to convert one unit to a required unit. The following conversions are widely used:

$$\frac{1\ \text{t}}{1000\ \text{kg}};\ \frac{1\ \text{kg}}{1000\ \text{g}};\ \frac{1\ \text{kg}}{1\ 000\ 000\ \text{mg}};\ \frac{1\ \text{L}}{1000\ \text{mL}}$$

Of course, the reciprocal of each can also be used.

Sample Problem 1: Calculating Volume from % V/V

The concentration of ethanol in a 750 mL bottle of wine is 13.5 % V/V. Assuming that wine has the same density as water, calculate the volume of ethanol in the bottle.

Given: $c_{V/V} = 13.5\ \%$; $V_{solution} = 750$ mL

Required: volume of ethanol, $V_{ethanol}$

Analysis:

$$c_{V/V} = \frac{V_{ethanol}}{V_{solution}} \times 100\ \%$$

Solution:

Step 1. If necessary, convert measured quantities to a common unit.

In this case, since the question does not require the answer to be in litres, converting 750 mL to 0.750 L is not necessary.

Step 2. Rearrange the concentration equation and solve for $V_{ethanol}$.

$$c_{V/V} = \frac{V_{ethanol}}{V_{solution}} \times 100\ \%$$

LEARNING **TIP**

Weight and Mass

In everyday language, "weight" and "mass" are often used interchangeably. Remember, however, that they are not the same. Used correctly, "mass" is the quantity of substance and is measured in grams or kilograms. "Weight" is the force of gravity on the object and is measured in newtons. You would weigh much less on the Moon than on Earth, because gravity is not so strong there. Of course, your mass would be unchanged.

Figure 3 This acne medication contains 5 % W/W benzoyl peroxide.

LEARNING **TIP**

Working with Percentages

To determine the percentage of a number, remember that the word "of" means "multiplied by." Simply multiply the number by the percentage written as a decimal. For example,

13.5 % of 750 = (0.135)(750) = 101

$$V_{\text{ethanol}} = \frac{c_{\text{V/V}} V_{\text{solution}}}{100}$$

$$= \frac{13.5 \times 750 \text{ mL}}{100}$$

$$V_{\text{ethanol}} = 101 \text{ mL}$$

Statement: The bottle of wine contains 101 mL ethanol.

Sample Problem 2: Calculating Volume of Solution from Mass of Solute

Glucose, $C_6H_{12}O_6$, is used to prepare intravenous feeding solutions. What volume of a 5.0 % W/V glucose solution can be prepared using 125 g of glucose?

Given: m_{glucose} = 125 g; c_{glucose} = 5.0 % W/V or 5.0 g/100 mL

Required: V_{solution}

Analysis:

$$c_{\text{glucose}} = \frac{m_{\text{glucose}}}{V_{\text{solution}}}$$

Solution:

$$V_{\text{solution}} = \frac{m_{\text{glucose}}}{c_{\text{glucose}}}$$

$$= 125 \text{ g}_{\cancel{\text{glucose}}} \times \frac{100 \text{ mL}_{\text{solution}}}{5.0 \text{ g}_{\cancel{\text{glucose}}}}$$

$$V_{\text{solution}} = 2500 \text{ mL or } 2.5 \text{ L (optional)}$$

Statement: 125 g of glucose can be used to prepare 2.5 L of a 5.0 % W/V solution.

Sample Problem 3: Calculating Dilutions Using Percentage Concentrations

What is the final concentration when 125 mL of a 30.0 % W/V hydrogen peroxide solution is diluted to 2.000 L?

Given: c_c = 30.0 % W/V; V_c = 125 mL; V_d = 2.000 L

Required: concentration of the diluted solution, c_d

Analysis: Convert both volumes to the same unit, either mL or L.

$$125 \text{ mL} = 125 \cancel{\text{mL}} \times \frac{1 \text{ L}}{1000 \cancel{\text{mL}}} = 0.125 \text{ L}$$

Use the dilution equation:

$$c_cV_c = c_dV_d$$

$$\frac{c_cV_c}{V_d} = c_d$$

Solution:

$$c_d = \frac{c_cV_c}{V_d}$$

$$= \frac{30.0 \times 0.125 \cancel{\text{L}}}{100 \times 2.000 \cancel{\text{L}}}$$

$$c_d = 0.0188 \text{ or } 1.88 \text{ \% W/V}$$

Statement: The final concentration of the hydrogen peroxide solution is 1.88 % W/V.

Practice

1. Gasoline contains as much as 10 % V/V ethanol. What is the maximum volume of ethanol in a 40 L fill-up of gasoline? T/I [ans: 4.0 L]
2. A 2.5 L jug of household chlorine bleach contains 125 g of sodium hypochlorite, NaClO. Determine the percentage weight/volume of the bleach. T/I [ans: 5.0 % W/V]
3. Dissolving boric acid in water produces a mild antiseptic that is useful in eye washes. What mass of boric acid is present in 125 g of a 2.5 % W/W solution? T/I [ans: 3.1 g]
4. What volume of a 30.0 % W/V hydrogen peroxide solution is required to prepare 425 mL of a 6.0 % W/V solution? T/I [ans: 85 mL]

Extremely Low Concentrations

Scientists working in the medical or environmental fields often measure concentrations that are extremely low. Expressing these concentrations as amount concentrations or percentage concentrations gives awkward numbers to work with.

For this reason, very dilute concentrations are often expressed in parts per million (ppm, $1{:}10^6$), parts per billion (ppb, $1{:}10^9$), and even parts per trillion (ppt, $1{:}10^{12}$) (**Table 1**). For example, a bottle of water may contain sodium ions at a concentration of 36 ppm (**Figure 4**). This means that there are 36 parts of sodium for every million parts of solution. This concentration is so dilute you would have to drink 56 half-litre bottles of water to consume only 1 g of dissolved sodium ions!

Table 1 Parts Per Million, Billion, Trillion

Part per	Equivalent to
1 ppm	1 drop in a bathtub full of water 30 s out of a year
1 ppb	1 drop in 250 full barrels 3 s out of a century
1 ppt	1 drop in 20 Olympic-sized pools 3 s out of 100 000 years

Masses and Volumes of Very Dilute Solutions

"Parts per" concentrations are actually a special case of % W/W concentrations. Since a very dilute aqueous solution is very similar to pure water, its density is assumed to be 1 g/mL at 20 °C. Therefore, a solute concentration of 1 ppm means that there is 1 g of solute in every 1 000 000 g or 1 000 000 mL of solution. This is a very small mass of solute per gram of solution:

$$c = \frac{1\ \text{g}_{\text{solute}}}{1\ 000\ 000\ \text{g}_{\text{solution}}}$$

$$= 1 \times 10^{-6}\ \text{g}_{\text{solute}}/\text{g}_{\text{solution}}$$

To make this number more convenient to work with, chemists multiply it by 10^6 to define a ppm. The concentration of a solution in ppm is determined using

$$c_{\text{ppm}} = \frac{m_{\text{solute}}}{m_{\text{solution}}} \times 10^6$$

Similarly, concentrations in parts per billion (ppb) and parts per trillion (ppt) can be determined using

$$c_{\text{ppb}} = \frac{m_{\text{solute}}}{m_{\text{solution}}} \times 10^9 \qquad c_{\text{ppt}} = \frac{m_{\text{solute}}}{m_{\text{solution}}} \times 10^{12}$$

When the solvent is water, we can use several units for these concentrations.

$$\begin{aligned}
1\ \text{ppm} &= 1\ \text{g}/10^6\ \text{g} & 1\ \text{ppb} &= 1\ \text{g}/10^9\ \text{g} & 1\ \text{ppt} &= 1\ \text{g}/10^{12}\ \text{g}\\
&= 1\ \text{g}/10^6\ \text{mL} & &= 1\ \text{g}/10^6\ \text{L} & &= 1\ \text{g}/10^9\ \text{L}\\
&= 1\ \text{g}/1000\ \text{L} & &= 1\ \text{mg}/1000\ \text{L} & &= 1\ \text{mg}/10^6\ \text{L}\\
&= 1\ \text{mg/L} & & & &
\end{aligned}$$

When solving problems involving very low concentrations, select the unit that best matches the information given in the problem.

Figure 4 On the labels of bottled water, concentrations are given in parts per million (ppm).

Tutorial 2 Solving Problems with "Parts per" Concentrations

Assuming that 1000 mL or 1 L of a very dilute aqueous solution has a mass of 1000 g or 1 kg simplifies calculations involving these solutions. Here are some basic tips for using the "parts per" formulas:

- All masses must have the same unit.
- The density of water is 1.0 g/mL at 20 °C. Therefore, the mass and volume of a very dilute aqueous solution have the same numerical value. For example, 100 mL of a very dilute aqueous solution has a mass of 100 g.
- 1 t = 1000 kg; 1 kg = 1000 g; 1 kg = 1 000 000 mg; 1 L = 1000 mL

Sample Problem 1: Calculating Concentrations in Parts per Million

A 1.5 L sample of pool water contains 4.5 mg of chlorine. Determine the concentration of chlorine in ppm.

Given: $V_{H_2O} = 1.5\ \text{L};\ m_{Cl} = 4.5\ \text{mg}$

Required: chlorine concentration in parts per million, c_{ppm}

Analysis: $c_{ppm} = \dfrac{m_{solute}}{m_{solution}} \times 10^6$

Assuming that pool water has the same density as water (1.0 g/mL), 1.5 L of pool water has a mass of 1.5 kg. Since 1 kg $= 10^6$ mg,

$$1.5\ \text{kg} = 1.5 \times 10^6\ \text{mg}$$

Solution:

$$c_{ppm} = \frac{m_{Cl}}{m_{H_2O}} \times 10^6$$

$$= \frac{4.5 \times \cancel{10^6\ \text{mg}}}{1.5 \times \cancel{10^6\ \text{mg}}}$$

$$c_{ppm} = 3.0\ \text{ppm}$$

Statement: The chlorine concentration of the pool water is 3.0 ppm.

Practice

SKILLS HANDBOOK A6

1. In 2009, testing of certain brands of baby shampoos found unsafe concentrations of methanal (**Figure 5**). This chemical is thought to form from the breakdown of ingredients in the shampoo. Exposure to 250 ppm of methanal may cause a rash. What mass of methanal is present in 500 g of shampoo if its concentration is 250 ppm? T/I [ans: 0.13 g]
2. The maximum acceptable concentration of fluoride in tap water is 1.5 mg/L. T/I
   (a) Express this concentration in ppm. [ans: 1.5 ppm]
   (b) What volume of tap water contains 1.0 g of fluoride? [ans: 670 L]
3. A 250 mL sample of tap water is found to contain 12 ppb of an antibiotic. Determine the mass of the antibiotic in the sample. T/I [ans: 3.0×10^{-6} g]

LEARNING TIP

Percentages and Exponents
"ppm" is similar to the symbol "%" in the equations involving percentage concentration. You could think of the "× 100" in the above equations as "× 10^2." You could even think of "%" as "pph"—parts per hundred!

Figure 5 Methanal (commonly called formaldehyde) is a contaminant found in some baby shampoos. Methanal can irritate the skin and is a potential carcinogen.

Research This

The BPA Controversy

Skills: Researching, Analyzing, Evaluating, Communicating, Identifying Alternatives, Defending a Decision

SKILLS HANDBOOK A5.1

Bisphenol-A or BPA is used to manufacture hard plastic products including children's toys, rigid plastic bottles, and the inner lining of food cans. Traces of BPA have been detected in some rather unusual places—like seawater and beach sand. There is much debate about the potential health risks of BPA. For example, BPA has been shown to mimic estrogen, a female hormone. Some Canadian scientists suspect that BPA may reduce fertility. In October 2010, health concerns prompted the federal government to declare BPA a toxic substance. This announcement came shortly after the European Food Safety Authority reaffirmed that BPA is safe for use in food containers.

1. Research the current status of our understanding of the sources and potential health risks of BPA.

A. How is the "toxic" designation likely to affect the companies that manufacture this product? T/I

B. List three examples of BPA concentrations detected in common items. T/I

C. Do you think the "toxic" designation is justified? Explain. T/I

D. What alternatives do we have if the use of BPA is banned? T/I A

GO TO NELSON SCIENCE

8.8 Summary

Several different units are used to communicate concentrations of consumer products and substances in the environment (**Table 2**).

Table 2 Measure of Concentration

Name	Abbreviation	Equation	Application
percentage volume/volume	% V/V	$c_{V/V} = \frac{V_{solute}}{V_{solution}} \times 100\ \%$	liquid–liquid mixtures
percentage weight/volume	% W/V	$c_{W/V} = \frac{m_{solute}}{V_{solution}} \times 100\ \%$	solid–liquid mixtures
percentage weight/weight	% W/W	$c_{W/W} = \frac{m_{solute}}{m_{solution}} \times 100\ \%$	solid–liquid or solid–solid mixtures
parts per million	ppm	$c_{ppm} = \frac{m_{solute}}{m_{solution}} \times 10^6\ \text{ppm}$	to express small concentrations (e.g., composition of air)
parts per billion	ppb	$c_{ppb} = \frac{m_{solute}}{m_{solution}} \times 10^9\ \text{ppb}$	to express very small concentrations (e.g., metal contaminants in water)
parts per trillion	ppt	$c_{ppt} = \frac{m_{solute}}{m_{solution}} \times 10^{12}\ \text{ppt}$	to express extremely small concentrations (e.g., traces of medications in water)

8.8 Questions

1. Swimming pool manufacturers recommend maintaining the pool chlorine concentration at 3.0 ppm. What is the mass of chlorine in a pool containing 3.4×10^6 L of water? T/I
2. Formalin, a 37 % W/W solution of methanal in water, is used to preserve biological specimens. What mass of water must be added to 250 g of methanal to make formalin? T/I
3. Rubbing alcohol is a 70 % V/V solution of 2-propanol (isopropyl alcohol) in water. What volume of 2-propanol does 1.5 L of rubbing alcohol contain? T/I
4. A lab technician prepared 250 mL of an aqueous solution containing 45 g glucose, $C_6H_{12}O_6$. Express the concentration of the solution as: T/I
   (a) percentage weight/volume
   (b) amount concentration
5. In 2003, technicians analyzed the wastewater going into a water treatment plant. They detected a 30.08 mg/L concentration of ibuprofen, a pain-relieving medication. T/I
   (a) Express this concentration in ppm.
   (b) What volume of water contains 1.00 g of ibuprofen?
6. Many compounds have been detected in the wastewater from Canadian homes. Triclosan is a common antibacterial agent in hand sanitizers. A 1.5 L sample of water was found to contain 38 ng of Triclosan. (1 ng $= 10^{-9}$ g) What is the concentration of Triclosan in parts per billion? T/I
7. Which is more concentrated, a 1.0 mol/L solution of NaCl or a 1 % W/W solution of NaCl? Explain. T/I
8. Why is it important to standardize the way in which concentrations of industrial chemicals are expressed? A
9. A solution of sodium hydrogen carbonate (also called sodium bicarbonate), $NaHCO_3$, is used as a medium for intravenous injections (**Figure 6**). Use the product label to determine the following quantities: T/I

Figure 6 Sodium hydrogen carbonate solution

   (a) the mass of sodium hydrogen carbonate in the bottle
   (b) the amount of sodium hydrogen carbonate in the bottle
   (c) the concentration of sodium hydrogen carbonate as % W/V
   (d) the concentration of sodium hydrogen carbonate in ppm
   (e) the amount concentration
10. The concentration of gold in the ocean is estimated to be 12 ppt. The volume of the Atlantic Ocean is 354 million cubic kilometres. (1 km^3 $= 10^9$ m^3) T/I
    (a) Given that the average density of the ocean is 1035 kg/m^3, calculate the average mass of the Atlantic Ocean.
    (b) Calculate the mass of gold in the Atlantic Ocean.
11. (a) Research, either at home or on the Internet, 3 everyday products that show the concentration on the container.
    (b) In each case, describe the units of the concentration and explain why those units were chosen. A

GO TO NELSON SCIENCE

Investigation 8.6.1 OBSERVATIONAL STUDY

Preparing a Standard Solution from a Solid

SKILLS MENU

- Questioning
- Researching
- Hypothesizing
- Predicting
- Planning
- Controlling Variables
- Performing
- Observing
- Analyzing
- Evaluating
- Communicating

Preparing a standard solution is an essential skill that every chemist must master. In this activity you will practice the skills required to prepare a specific volume of a standard solution. You will be preparing a solution of copper(II) sulfate. You will start with solid copper(II) sulfate pentahydrate, so your calculations must take into account the mass of water associated with each formula unit of copper(II) sulfate.

Purpose

To prepare 100 mL of a 0.100 mol/L solution of copper(II) sulfate pentahydrate, $CuSO_4{\bullet}5H_2O$

Equipment and Materials

SKILLS HANDBOOK A1.1, A1.2

- chemical safety goggles
- lab apron
- scoopula
- balance
- 150 mL beaker
- 100 mL volumetric flask and stopper
- funnel
- massing paper
- wash bottle of distilled water
- dropper bottle of distilled water
- copper(II) sulfate pentahydrate, $CuSO_4{\bullet}5H_2O(s)$

Copper(II) sulfate is toxic and an irritant. Avoid skin and eye contact. If you spill copper(II) sulfate on your skin, wash the affected area with plenty of cool water. Report any spills to your teacher.

Procedure

SKILLS HANDBOOK A3.2, A3.3, A3.7

1. Calculate the mass of copper(II) sulfate pentahydrate, $CuSO_4{\bullet}5H_2O(s)$, required for the solution. Remember to take into account the mass of the water of hydration.
2. Put on your chemical safety goggles and lab apron.
3. Measure the required mass of solute and transfer it to a clean, dry 150 mL beaker. Rinse the paper to ensure that all of the crystals are washed into the beaker.
4. Dissolve the solute in about 50 mL of distilled water.
5. Ensure that the volumetric flask is clean by rinsing it with three small volumes of distilled water. Discard the rinse water.
6. Transfer the solution to the volumetric flask, using the funnel.
7. Rinse the beaker three times with small volumes of distilled water. Transfer each rinse into the flask through the funnel.
8. Stopper the flask and swirl to ensure that all the solute is dissolved and the solution is uniformly mixed.
9. Add distilled water to the flask until the level of the solution is about 1 cm below the 100.0 mL mark.
10. Use water from the dropper bottle to fill the flask exactly to the 100.0 mL mark. The bottom of the water meniscus should be at the same level as the mark on the flask.
11. Stopper the flask and invert it three times to ensure thorough mixing.
12. Save this solution for the next investigation.

Analyze and Evaluate

(a) What mass of copper(II) sulfate pentahydrate is required to prepare this solution? Show your calculations. T/I

(b) Which of the following pieces of equipment had to be dry initially: beaker, flask, scoopula? Which did not have to be dry? Why? T/I

(c) Why was it necessary to transfer each rinse to the flask? T/I

(d) Why is it necessary to store the solution in a stoppered container? T/I

(e) Why could you not prepare this solution by dissolving the required mass of copper(II) sulfate pentahydrate in 100.0 mL of distilled water? T/I

Apply and Extend

(f) A student accidentally prepared a solution in which the final position of the meniscus is above the volume marking on the flask. Is the final concentration of this solution greater than, less than, or equal to 0.100 mol/L? Why? T/I

(g) Design a lab procedure that could be used to return the solution in (f) to 0.100 mol/L. Describe why your procedure would work. T/I

Investigation 8.7.1 OBSERVATIONAL STUDY

Preparing a Standard Solution by Dilution

SKILLS MENU

• Questioning	• Planning	• Observing
• Researching	• Controlling Variables	• Analyzing
• Hypothesizing	• Performing	• Evaluating
• Predicting		• Communicating

In this activity you will use the solution you prepared in the previous activity to prepare at least one dilute solution of known concentration. Refer to Section 8.7, Figure 2, to review the technique.

Purpose

SKILLS HANDBOOK A1.1, A1.2

To prepare 100 mL of a 0.010 mol/L solution of copper(II) sulfate from a 0.100 mol/L solution

Equipment and Materials

- chemical safety goggles
- lab apron
- variety of pipettes
- pipette pump
- 100 mL volumetric flask and stopper
- funnel
- dropper
- wash bottle of distilled water
- beaker containing a prepared solution of 0.100 mol/L copper(II) sulfate, $CuSO_4$(aq)

Copper(II) sulfate is toxic and an irritant. Avoid skin and eye contact. If you spill copper(II) sulfate on your skin, wash the affected area with lots of cool water. Always use a pipette pump to draw solutions into the pipette. Never pipette orally. Report any spills to your teacher.

Procedure

SKILLS HANDBOOK A3.3, 3.4, 3.7

1. Calculate the volume of 0.100 mol/L copper(II) sulfate, $CuSO_4$(aq), required to prepare 100 mL of a 0.010 mol/L solution. Decide which type of pipette you will need. Check with your teacher to approve your plans before proceeding.
2. Put on your chemical safety goggles and lab apron.
3. If this is your first time using a pipette and pump to transfer solutions from one container to another, practice using tap water. Once you are proficient at this skill, proceed to Step 4.
4. Follow the technique outlined in Section 8.7, Figure 2, to complete the dilution (**Figure 1**).
5. Dispose of the solution as directed by your teacher and wash your hands thoroughly.

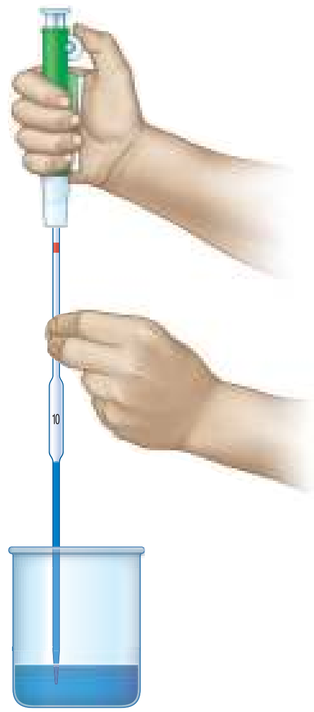


Figure 1 Using a volumetric pipette to transfer a solution

Analyze and Evaluate

(a) What volume of the concentrated solution was required? Show your calculations. T/I

(b) What visual evidence suggests that the final solution is more dilute than the original? T/I

(c) Which of the following pieces of equipment had to be dry initially: pipette, flask, dropper? Which could be wet with distilled water? Why? T/I

(d) Suppose that the distilled water contained a contaminant that forms a precipitate with copper(II) ions. What effect would this have on the final concentration of the solution? Why? T/I

(e) What effect, if any, does the addition of distilled water to the flask have on
   (i) the mass of Cu^{2+}(aq) in the flask
   (ii) the amount of Cu^{2+}(aq) in the flask
   (iii) the concentration of Cu^{2+}(aq) T/I

Apply and Extend

(f) Design a procedure to prepare two other dilute solutions as directed by your teacher. T/I

Summary Questions

1. Create a chapter summary based on the points listed in the margin on page 368. For each point, create three or four sub-points that provide relevant examples, explanatory diagrams, or general equations.

2. Look back at the Starting Points questions on page 368. Answer these questions using what you have learned in this chapter. Compare your latest answers with those that you wrote at the beginning of the chapter. Note how your answers have changed.

Vocabulary

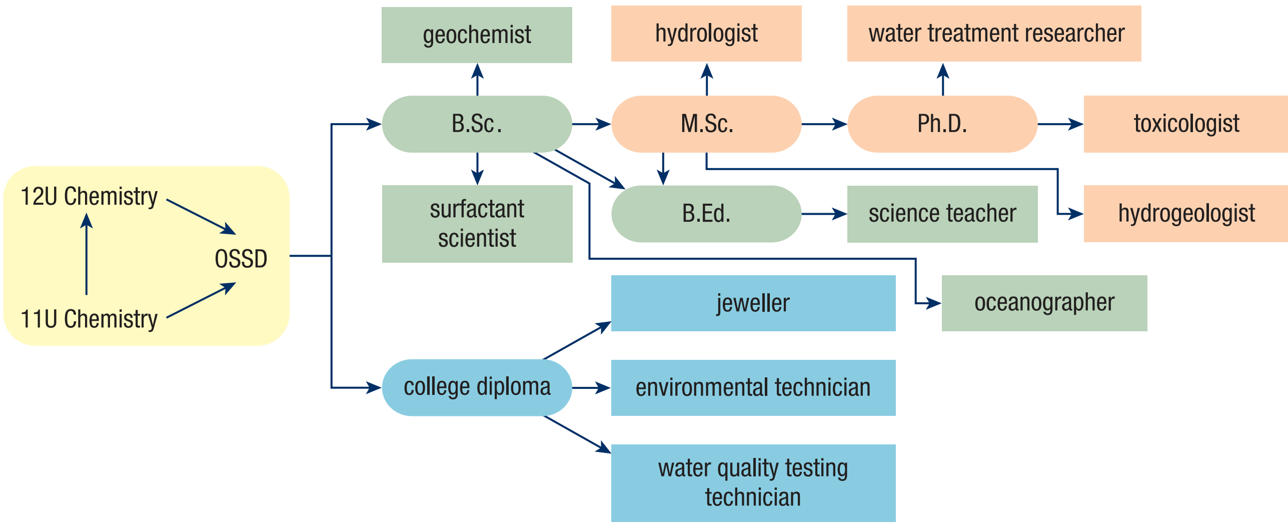
water cycle (p. 372)
transpiration (p. 372)
aquifer (p. 372)
potable water (p. 372)
heterogeneous mixture (p. 377)
concentration (p. 378)
concentrated solution (p. 378)
dilute solution (p. 378)
alloy (p. 379)
amalgam (p. 379)
aqueous solution (p. 379)
hydration (p. 383)
dissociation (p. 383)
miscible (p. 384)
immiscible (p. 384)
surfactant (p. 386)
saturated solution (p. 392)
unsaturated solution (p. 392)
supersaturated solution (p. 392)
solubility curve (p. 393)
thermal pollution (p. 395)
pressure (p. 395)
amount concentration (p. 398)
stock solution (p. 398)
standard solution (p. 401)
dilution (p. 403)

CAREER PATHWAYS

SKILLS HANDBOOK A7

Grade 11 Chemistry can lead to a wide range of careers. Some require a college diploma or a B.Sc. degree. Others require specialized or post-graduate degrees. This graphic organizer shows a few pathways to careers mentioned in this chapter.

1. Select two careers, related to water and solutions, that you find interesting. Research the educational pathways that you would need to follow to pursue these careers. What is involved in the required educational programs? Prepare a brief report of your findings.
2. For one of the two careers that you chose above, describe the career, main duties and responsibilities, working conditions, and setting. Also outline how the career benefits society and the environment.



GO TO NELSON SCIENCE

CHAPTER 8 SELF-QUIZ

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

For each question, select the best answer from the four alternatives.

1. What is the state, at room temperature, of most pure compounds with molecular masses similar to that of water? (8.1) K/U
   (a) solid (c) gas
   (b) liquid (d) solution
2. What name is given to the movement of water on, above, or below Earth's surface? (8.1) K/U
   (a) the water cycle
   (b) high surface tension
   (c) dissolving
   (d) transpiration
3. Alloys and salt water are examples of what type of mixture? (8.2) K/U
   (a) amalgam
   (b) homogeneous
   (c) heterogeneous
   (d) aqueous
4. Which of the following substances does *not* form a solution with water? Assume that only water and the other substance are present. (8.3) K/U
   (a) sugar
   (b) alcohol
   (c) vegetable oil
   (d) table salt
5. Which of the following statements explains why ionic and polar compounds dissolve in water? (8.3) K/U
   (a) Its bonds are polar.
   (b) It has high surface tension.
   (c) It has high specific heat capacity.
   (d) It is immiscible in oil.
6. A student drops a crystal of sodium chloride into an aqueous sodium chloride solution. The crystal dissolves. What term best describes the original solution? (8.5) K/U
   (a) saturated
   (b) unsaturated
   (c) supersaturated
   (d) heterogeneous
7. Which statement expresses the correct way to dilute a concentrated acid? (8.6) K/U
   (a) Add acid to water.
   (b) Add water to acid.
   (c) Pour both water and acid at the same time into another container.
   (d) Either (a) or (b)
8. What happens to the number of moles present during a dilution? (8.7) K/U
   (a) It decreases.
   (b) It increases.
   (c) It stays the same.
   (d) It may increase or decrease, depending on the dilution.
9. In general, the concentration of a solution is a comparison of what two quantities? (8.7, 8.8) K/U
   (a) quantity of solute and quantity of solvent
   (b) quantity of solute and quantity of solution
   (c) quantity of solvent and quantity of solution
   (d) total quantity of solute and solvent, and quantity of solution
10. A glucose solution contains about 30 % glucose and about 70 % water. What is the approximate percentage volume/volume of the solution? (8.8) K/U
    (a) 20 % (c) 70 %
    (b) 30 % (d) 100 %

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

11. Hydrogen bonding accounts for the high melting point of water. (8.1) K/U
12. When a gas is dissolved in a liquid, the gas is always the solvent. (8.2) K/U
13. A solution that contains a relatively large quantity of solute dissolved per unit volume of solvent is a dilute solution. (8.2) K/U
14. Liquids that do not mix with each other are said to be miscible. (8.3) K/U
15. During an oil spill, oil forms a layer on the top of water because water and oil are immiscible. (8.3, 8.4) K/U
16. A supersaturated solution contains more solute than an equal volume of a saturated solution at the same temperature. (8.5) K/U
17. In general, the solubility of an ionic compound decreases as temperature increases. (8.5) K/U
18. A standard solution is one that is to be diluted to a lower concentration before it is used. (8.6) K/U
19. During a dilution, concentration decreases. (8.7) K/U
20. The volume of solution decreases during a dilution. (8.7) K/U
21. Very dilute concentrations might be expressed in ppm. (8.8) K/U

To do an online self-quiz,
GO TO NELSON SCIENCE

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

Knowledge

For each question, select the best answer from the four alternatives.

1. What is the main source of energy for the water cycle? (8.1) K/U
   (a) the Sun
   (b) heat from Earth's interior
   (c) thermal energy from the ocean
   (d) the specific heat capacity of water
2. What is the name given to a layer of rock or sediment that holds water and allows water to flow through it? (8.1) K/U
   (a) water table
   (b) aquifer
   (c) surface water
   (d) gravel
3. A student compares the mass of a solid and the volume of liquid in which the solid dissolves. What concept is the student determining? (8.2) K/U
   (a) solute
   (b) solvent
   (c) solution
   (d) concentration
4. Which of the following substances can "encourage" oil and water to mix? (8.3) K/U
   (a) an ionic compound
   (b) a hydration agent
   (c) a hydrocarbon
   (d) a surfactant
5. What was the biggest environmental concern during the oil spill in the Gulf of Mexico in 2010? (8.4) K/U
   (a) The oil would damage life at the bottom of the ocean.
   (b) The oil spill caused large amounts of non-renewable resources to be lost.
   (c) The oil could reach land and damage ecosystems.
   (d) The spill could affect wave action in the Gulf of Mexico.
6. Examine **Figure 1**. Which compound has the solubility that is *least* affected by an increase in temperature? (8.5) K/U
   (a) NaCl
   (b) KNO_3
   (c) $NaHCO_3$
   (d) NH_4Cl

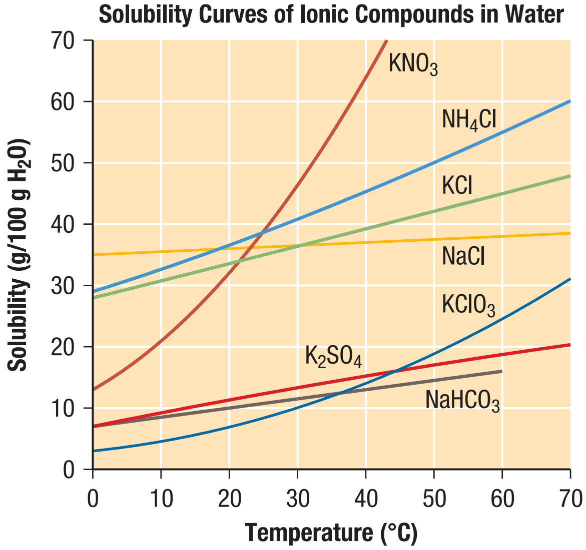


Figure 1

7. What is the correct equation for the amount concentration of a solution? (8.6) K/U
   (a) $V = \frac{c}{n}$ (c) $c = \frac{n}{V}$
   (b) $n = \frac{V}{c}$ (d) $c = \frac{V}{n}$
8. Dilute sodium chloride solutions are commonly used as IV fluids and to treat nasal congestion. During a dilution, what happens to the concentration of a sodium chloride solution? (8.7) K/U
   (a) The concentration increases.
   (b) The concentration decreases.
   (c) The concentration stays the same.
   (d) The concentration depends on the nature of the solute.
9. "Parts per" units of concentration are special cases of what type of concentration? (8.8) K/U
   (a) percentage volume/volume
   (b) percentage weight/volume
   (c) percentage volume/weight
   (d) percentage weight/weight

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

10. Most of the water on Earth is fresh water. (8.1) K/U
11. When a solid dissolves in a liquid, the solid is always the solvent. (8.2) K/U
12. Liquids that are immiscible with each other mix in all proportions. (8.3) K/U

13. In an oil-and-water mixture, dispersants break up oil into small droplets. (8.4) K/U
14. Changes in pressure significantly affect the solubility of a solid in a liquid. (8.5) K/U
15. Molar concentration is the correct IUPAC name for solution concentration. (8.6) K/U
16. A standard solution is a solution of known concentration. (8.6) K/U
17. A volumetric pipette can deliver a range of volumes. (8.7) K/U
18. Aqueous solutions with extremely low concentrations have a density equal to the density of water. (8.8) K/U

Match each description on the left with the appropriate term on the right.

19. (a) a mixture that is *not* uniform in appearance
    (b) a mixture that is uniform at the molecular level
    (c) a substance that dissolves
    (d) the medium that does the dissolving
    (e) quantity of solute compared to quantity of solvent

    (i) solute
    (ii) homogeneous mixture
    (iii) concentration
    (iv) solvent
    (v) heterogeneous mixture (8.2) K/U

20. (a) percentage volume/volume
    (b) parts per billion
    (c) percentage weight/weight
    (d) percentage weight/volume
    (e) parts per million
    (f) parts per trillion

    (i) $\frac{m_{\text{solute}}}{V_{\text{solution}}} \times 100\ \%$
    (ii) $\frac{m_{\text{solute}}}{m_{\text{solution}}} \times 10^{12}$
    (iii) $\frac{V_{\text{solute}}}{V_{\text{solution}}} \times 100\ \%$
    (iv) $\frac{m_{\text{solute}}}{m_{\text{solution}}} \times 10^{6}$
    (v) $\frac{m_{\text{solute}}}{m_{\text{solution}}} \times 100\ \%$
    (vi) $\frac{m_{\text{solute}}}{m_{\text{solution}}} \times 10^{9}$ (8.8) K/U

Write a short answer to each question.

21. Explain why water is a liquid, not a gas, at room temperature. (8.1) K/U
22. According to the United Nations General Assembly, everybody in the world has the right to potable water. What is meant by "potable water"? (8.1) K/U
23. Explain the difference between a homogeneous mixture and a heterogeneous mixture. Use an example of each type of mixture in your explanation. (8.2) K/U
24. At 25 °C, approximately 2000 g of sugar will dissolve in 1000 g of water. Is this a concentrated solution or a dilute solution? Explain. (8.2) K/U
25. (a) What is an amalgam?
    (b) Give an example of where an amalgam might be used. (8.2) K/U
26. What happens during the process of hydration? (8.3) K/U
27. What property of water does a surfactant change so that a non-polar molecule and water will mix? (8.3) K/U
28. How does the combination of dispersants and micro-organisms reduce the damage caused by oil spills? (8.4) K/U
29. Examine **Figure 2**. Suppose a solution contains 150 g of sucrose dissolved in 100 g of water at 5 °C. Is the solution saturated, unsaturated, or supersaturated? Explain your answer. (8.5) K/U

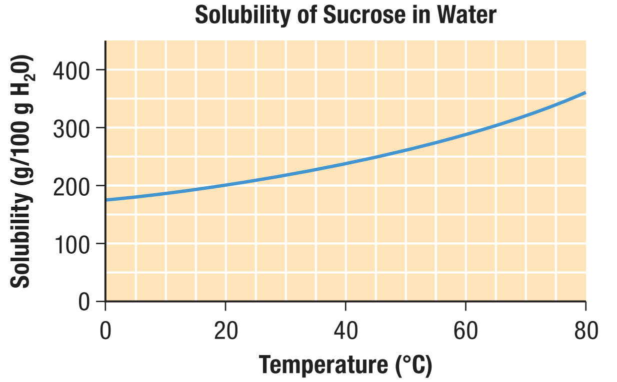


Figure 2

30. (a) In general, how do temperature and pressure affect the solubility of a solid in a liquid?
    (b) How do temperature and pressure affect the solubility of a gas in a liquid? (8.5) K/U
31. Outline the steps required to prepare 1.00 L of a solution with a specified amount concentration. (8.6) K/U
32. Why do students usually *not* use stock solutions of acids in a high school laboratory? (8.6) K/U
33. Rewrite the concentration equation to solve for each of the following variables: (8.6) K/U
    (a) concentration
    (b) volume
    (c) amount of solute (in moles)
34. Compare and contrast a volumetric pipette and a graduated pipette. (8.7) K/U
35. Reorganize the dilution equation to solve for each of the following variables: (8.7) K/U
    (a) concentration of the concentrated solution
    (b) volume of the concentrated solution
    (c) concentration of the dilute solution
    (d) volume of the dilute solution

36. When you clean your floor, you might pour some household floor cleaner into a bucket of water. However, when you prepare a dilute hydrochloric acid solution in a lab, you would carefully measure both the volume of concentrated acid and the total volume of solution. Explain why you treat these two dilutions differently. (8.7) K/U
37. A concentration of a trace mineral was found to be 2.53 ppb in an aqueous solution. (8.8) K/U
    (a) What mass of solution contains 2.53 g of the mineral?
    (b) What volume of solution contains 2.53 g of the mineral?

Understanding

38. Many properties of water are the result of the high polarity of water molecules. Why are water molecules considerably more polar than most other molecules that contain polar bonds? (8.1) K/U T/I
39. (a) Compare and contrast the two types of intermolecular forces that exist between water molecules.
    (b) To which of these forces are most of the physical properties of water attributed? (8.1) K/U T/I
40. Explain the high boiling point of water. (8.1) T/I
41. Examine **Figure 3**. (8.1) T/I

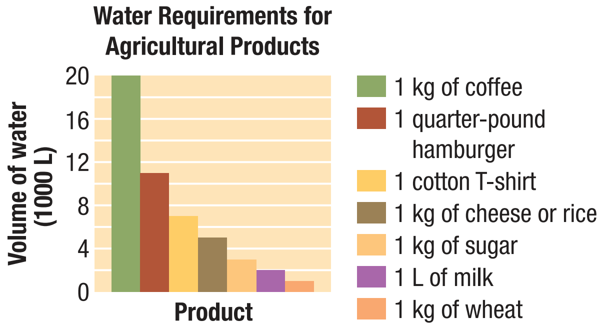


Figure 3

    (a) Compare the volumes of water needed to produce 1 kg of coffee and 1 kg of rice.
    (b) What other product requires about the same volume of water, per kilogram, as 1 kg of sugar and 1 L of milk?
42. Concentrated acetic acid is diluted with water to make a 5 % acetic acid solution, similar to the vinegar used on French fries. Explain how you know which substance is the solvent and which is the solute. (8.2) K/U
43. Classify each of the following mixtures as homogeneous or heterogeneous: (8.2) K/U A
    (a) a tossed salad
    (b) 14-karat gold, an alloy
    (c) clear, dry air
    (d) sweetened iced tea
44. Classify each of the following solutions as either concentrated or dilute: (8.2) K/U A
    (a) Sugar is added to a cup of coffee until no more sugar dissolves.
    (b) A few grains of salt are sprinkled into a bowl of chicken broth.
    (c) Steel is an alloy made of small percentages of various metals and carbon mixed with iron.
45. For each of the following sets of terms, write a sentence that shows how the terms relate to each other: (8.2, 8.3) K/U T/I C
    (a) homogeneous, heterogeneous
    (b) solute, solvent, solution
    (c) dilute, concentrated
    (d) surfactant, detergent
    (e) miscible, immiscible
46. In your own words, describe what happens when an ionic compound dissolves in water. (8.3) K/U T/I
47. For each of the following ionic compounds, write a chemical equation that shows the dissociation of the ions in an aqueous solution: (8.3) K/U C
    (a) $KCl(s)$
    (b) $MgBr_2(s)$
    (c) $Li_2SO_4(s)$
48. Why is water not written as part of a chemical equation that shows the dissociation of an ionic compound in an aqueous solution? (8.3) K/U
49. Ontario, Manitoba, Saskatchewan, and Alberta all have plants that manufacture ethanol, C_2H_5OH. Ethanol is mostly used as a fuel. Explain why ethanol, unlike most molecular compounds, mixes readily with water. (8.3) K/U
50. What is meant by the statement, "Like dissolves like"? (8.3) K/U
51. A student drops a crystal of potassium nitrate into an aqueous potassium nitrate solution. Explain each of the following possible observations, referring to the degree of saturation of the solution: (8.5) K/U T/I
    (a) The crystal quickly dissolves.
    (b) The crystal does not dissolve; it just rests on the bottom of the container, unchanged.
    (c) The crystal does not dissolve. Instead, it rapidly grows larger.
52. (a) What is thermal pollution?
    (b) What effect might thermal pollution have on aquatic organisms? (8.5) K/U
53. If you were operating a pop bottling plant, what temperature and pressure conditions would you use to obtain a high concentration of carbon dioxide gas in the pop? (8.5) K/U A

54. Calcium chloride, $CaCl_2$, is spread on roads in icy conditions because it stops ice from forming as long as temperatures stay above −18 °C. Sometimes a calcium chloride solution is sprayed on the road. If 245 mol of calcium chloride is used to make 1500 L of a de-icing solution, what is the amount concentration of the solution? (8.6) K/U T/I

55. Sodium hydroxide, NaOH(s), is the main ingredient in many drain cleaners. Suppose 125 mL of a drain cleaning solution is prepared by dissolving 8.2 g of sodium hydroxide in water. What is the amount concentration of the solution? (8.6) K/U T/I

56. (a) When using the concentration equation $c = \frac{n}{V}$, does it make any difference whether you use the volume of the final solution or the volume of solvent added, for V? Explain your answer.
    (b) Use your answer from Part (a) to explain why, when preparing 500 mL of a standard solution, you cannot just add 500 mL of water to the needed amount of solute. (8.6) K/U T/I A

57. Potassium nitrate, KNO_3, is a compound commonly used in fireworks, fertilizer, and rocket propellants. A 25 g sample of potassium nitrate is dissolved in enough water to make 250 mL of solution. What is the amount concentration of the resulting solution? (8.6) T/I

58. Potassium permanganate, $KMnO_4$, is sometimes used on farms as a disinfectant spray. Its advantage is that its reaction products are non-toxic. A farmer prepares a potassium permanganate solution to spray around a cattle pen. What mass of potassium permanganate is needed to prepare 14.5 L of solution with an amount concentration of 0.585 mol/L? (8.6) K/U T/I

59. Copy and complete **Table 1** by writing the form of the equation used to determine the unknown quantity. (8.6, 8.7) K/U

Table 1 Unknown and Known Quantities in Equations

Unknown quantity	Given (known) quantity	Equation
amount concentration	volume of solution, amount of solute	(a)
volume of solution	amount concentration, amount of solute	(b)
concentration of dilute solution	concentration of concentrated solution, volume of concentrated solution, volume of dilute solution	(c)

60. (a) Write the dilution equation.
    (b) Does this equation take into account the amount of solute, in moles? Explain your answer. (8.7) K/U T/I

61. Why is the following definition of dilution incorrect? "dilution: the process of reducing the concentration of a solution by adding water" (8.7) K/U

62. Sulfuric acid is an important component in a car battery. Although the amount concentration of this acid can vary, most battery acid has an amount concentration of approximately 4.0 mol/L. A stock solution of sulfuric acid usually has a concentration of 18.0 mol/L. (8.6, 8.7) K/U T/I
    (a) How might you prepare 1.00 L of battery acid from a stock solution of sulfuric acid?
    (b) As a battery is used, the amount concentration of battery acid decreases. The amount concentration of a 100 mL sample of battery acid is determined to be 3.3 mol/L. What volume of fresh battery acid contains the same amount of sulfuric acid as this sample of used battery acid?

63. A student dilutes 245 mL of a 1.50 mol/L solution. The final volume is 568 mL. What is the final amount concentration? (8.7) K/U T/I

64. A nurse prepares a solution of glucose, $C_6H_{12}O_6$, by dissolving 1.2 mol of glucose in enough water to make 1200 mL of solution. (8.8) K/U T/I
    (a) What else does the nurse need to know to calculate the percentage weight/volume of the solution?
    (b) What is the percentage weight/volume of the final solution?

65. Why is percentage concentration always less than 100 %? (8.8) T/I

66. On the label of a bottle of Ontario spring water, the concentration of dissolved minerals is given as 290 ppm. (8.8) K/U T/I
    (a) What is this concentration in ppb?
    (b) What is the concentration in ppt?

Analysis and Application

67. (a) Why do industries use so much of the world's available fresh water?
    (b) Choose a specific industry and suggest ways that it could reduce its consumption of water. (8.1) T/I A

68. Glucose, $C_6H_{12}O_6$, is a molecular compound that is soluble in water **(Figure 4)**. (8.3) K/U T/I A

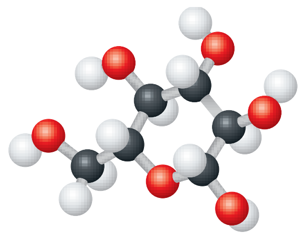


Figure 4 Glucose

    (a) Why is glucose soluble in water when most molecular compounds are not?
    (b) How does the solubility of glucose in water affect human health?

69. Would you expect liquid bromine, Br_2, to be more soluble in water or in a liquid hydrocarbon such as carbon tetrachloride, $CCl_4(l)$? Explain your answer by referring to electronegativity. (8.3) T/I C

70. (a) What is a surfactant?
    (b) Describe two different uses of surfactants.
    (c) Explain how adding a surfactant to your laundry can help remove a grease stain from your favourite shirt. (8.3) K/U A

71. A student prepares a supersaturated solution by dissolving 60 g of ammonium chloride, $NH_4Cl(s)$, in 100 g of water at 70 °C. She then gradually cools the solution to 50 °C. Finally, she adds a small seed crystal to the solution. Refer to Figure 4 in Section 8.5 to predict how much ammonium chloride crystallizes out of solution, leaving a saturated solution. (8.5) T/I A

72. Two aquaria contain the same volume of water and the same type of fish. One aquarium receives direct sunlight and the other is in the shade. Which aquarium probably needs a pump to bubble air through the water? Explain your answer. (8.5) T/I A

73. The water in a full 20.0 L saltwater aquarium has the ideal concentration of magnesium ions: 1.3 g/L. Then some of the water evaporates until the concentration of magnesium ions is 1.5 g/L. What volume of water must be added to the tank to bring the concentration of magnesium ions back to the ideal concentration and the volume back to 20.0 L? (8.7) T/I A

74. A student prepared a solution to use for feeding hummingbirds. She dissolved 88 g of sugar in enough water to make 439 mL of solution. (8.8) K/U T/I A
    (a) What is the percentage weight/volume of the solution?
    (b) Why is the concentration measured in percentage weight/volume instead of percentage volume/volume?

75. Two possible units used to represent ppm are 1 g/10^6 mL and 1 g/10^6 g. How can these two units be equivalent? (8.8) T/I

Evaluation

76. Using chemical processes, such as those used to help clean up the Gulf of Mexico oil spill in 2010, involves both benefits and risks. (8.4) K/U T/I A
    (a) List some benefits and risks associated with this use of dispersants.
    (b) Do you feel that using dispersants was a good decision? Explain your answer.

77. The data in **Table 2** show the solubility of copper(II) nitrate in water at various temperatures. (8.5) K/U T/I A

Table 2 Solubility of Copper(II) Nitrate in Water

Temperature (°C)	Solubility (g/100 mL H_2O)
0	83.5
10	100
20	125
30	156
40	163
60	182

    (a) Draw a solubility curve for copper(II) nitrate, using the data in Table 2. Your graph should be similar to Figure 2, Section 8.5.
    (b) Following the pattern shown by the existing graph, extend the curve to 80 °C.
    (c) From your graph, predict the solubility of copper(II) nitrate in water at 80 °C.
    (d) The actual solubility of copper(II) nitrate in water at 80 °C is 208 g/100 g H_2O. Calculate how close your predicted value is to the actual value by subtracting the smaller value from the larger one.
    (e) Calculate the percent error by dividing the difference in the actual and predicted values by the actual value, then multiplying by 100 %.

78. A student states that a solution containing 100 g of water and 10 g of potassium nitrate, KNO_3, at 30 °C is a dilute solution. She also states that a solution containing 100 g of water and 10 g of potassium chlorate, $KClO_3$, at 30 °C is a saturated solution. Can both of these statements be true? Refer to Figure 4 in Section 8.5 to explain your answer. (8.5) T/I

79. Explain how to prepare 250 mL of a standard solution from a solid. Write your explanation either in paragraph form or in a sequence concept map. In your explanation, include the following points:
    - the general procedure followed
    - how to read the volume in the flask
    - the advantage of using a volumetric flask over a beaker or another type of flask (8.6) K/U C A

80. A bottle of rubbing alcohol contains 70 % alcohol and 30 % water. (8.8) K/U T/I
    (a) Why is the description "70 % alcohol" on the label not technically accurate?
    (b) According to the naming guidelines in Section 8.8, how could this bottle of rubbing alcohol be labelled?

Reflect on Your Learning

81. (a) List three of the most important ways water is naturally purified.
    (b) Determine the original source of water for your home. Which of the ways listed in (a) is most important for purifying the water source for your home? Explain your answer. T/I A
82. The following conditions have been named as the three most important challenges in providing potable water to populations worldwide. For each condition, explain why it might be such a challenge. T/I A
    (a) population growth
    (b) increasing industrialization
    (c) pollution of surface water
83. For each of the following solutions, name:
    - the original state of the solute
    - the original state of the solvent
    - the state of the solution itself K/U T/I

    (a) bronze, an alloy
    (b) mercury amalgams
    (c) the mixture of molecules in a scuba tank
    (d) pop sweetened with sucrose
84. (a) List five aqueous solutions that you have used or observed today.
    (b) What makes these solutions aqueous solutions? K/U C A
85. Water heated by industrial processes or power plants must be cooled before it is released into water in the environment. How might this recaptured heat be made useful, rather than just releasing it into the environment? T/I A
86. Preparing a standard solution from a stock solution has certain similarities to using a recipe to prepare food. Explain why kitchen measuring devices, such as measuring cups, can be used to prepare food, but not to prepare standard solutions. T/I A
87. The human body needs sodium ions for proper functioning of nerves and muscles. However, too much sodium can interfere with water retention and raise blood pressure. How can you use dilution to reduce the amount of sodium you ingest? Give at least one specific example. T/I A
88. (a) The percentage concentrations of some solutions just need to be approximate. Name two solutions for which exact concentrations are not essential.
    (b) The percentage concentration of household chlorine bleach can vary slightly, but it needs to be fairly accurate. What do you think might happen if the percentage concentration were too high? What might happen if it were too low? T/I A

Research

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89. Industries that process oil and gas use huge volumes of water. K/U T/I C A
    (a) Why are large volumes of water needed for oil and gas production?
    (b) What type of groundwater is sometimes used as an alternative to fresh water?
    (c) What are these industries doing to prevent releasing polluted water back into the environment after the water is used?
    (d) What water-free processes are sometimes used so that water-intensive processes are not needed?
90. Solutions that are used in intravenous (IV) drips must have the correct concentration to be compatible with the fluids already in the human body. T/I A
    (a) Consider the glucose solution described in Section 8.2. Such a solution is said to be isotonic. What does it mean for a solution to be isotonic?
    (b) Why is it important for an IV solution to be isotonic?
    (c) Another solution that is commonly used for IVs is saline solution. What is the solute for a saline solution?
    (d) What percentage of an isotonic saline solution is solute?
91. Modern dispersants are effective in dealing with oil spills. However, dispersants were not as effective in dealing with the oil spill off the coast of Alaska in 1989 when an Exxon tanker spilled large volumes of oil into the waters there. T/I A C
    (a) Why were dispersants not very effective in treating the *Exxon Valdez* oil spill?
    (b) What was primarily used instead of dispersants to clean up the oil spill?
92. Canada has six nuclear power plants located in Ontario, Québec, and New Brunswick. Thermal pollution can result when water that has been used in nuclear reactors such as these is released into lakes or rivers. T/I
    (a) Name two ways in which large quantities of water are used in nuclear power plants.
    (b) What are the two main methods of cooling water heated by nuclear power plants before it is either reused or released into the environment?
93. The term "tincture" is used to describe some solutions. T/I A
    (a) What is a tincture?
    (b) If the percentage weight/volume of a tincture of iodine solution is 5 %, what mass of iodine is present in 1.5 L of solution?

CHAPTER 9

Solutions and Their Reactions

KEY CONCEPTS

After completing this chapter you will be able to

- analyze medical and environmental applications of solution chemistry
- predict combinations of ions that produce precipitates and write balanced net ionic equations for these reactions
- describe some of the technologies involved in water treatment
- investigate issues related to water treatment
- perform stoichiometric calculations related to chemical reactions of solutions
- design and conduct qualitative and quantitative analyses of solutions

How Do Reactions in Solution Affect Consumer Products and Water Quality?

Water is an ideal medium in which to conduct chemical reactions. You have already learned that water is the "universal solvent," meaning that it dissolves many substances. It is also relatively cheap and accessible. Once in solution, the entities that make up each substance completely separate. This allows them to mix well, at the molecular level, with other substances, and react quickly.

Some reactions in solution occur very quickly; others take a long time. For example, it takes millions of years for groundwater to sculpt a cave and form stalactites and stalagmites. The chemical reaction involves calcium carbonate and carbon dioxide dissolved in water:

$$CaCO_3(s) + H_2O(l) + CO_2(aq) \rightarrow Ca(HCO_3)_2(aq)$$

As a result, water droplets on the ceiling of the cave contain highly soluble calcium hydrogen carbonate. If some of the water in the droplets evaporates, calcium hydrogen carbonate decomposes to form calcium carbonate:

$$Ca(HCO_3)_2(aq) \rightarrow CaCO_3(s) + H_2O(l) + CO_2(g)$$

Since calcium carbonate is only slightly soluble, it precipitates from the solution. If precipitation happens before the drop falls, it forms the stalactites that grow from the ceiling of the cave. If it happens after the drop falls, calcium carbonate forms the stalagmites growing upward from the floor.

The precipitation of compounds from solution also has important applications in living systems. For example, the growth of coral depends on the precipitation of dissolved calcium and carbonate ions. Also, the growth and health of our bones depend on a steady supply of calcium ions in our diet. Some precipitation reactions, however, result in serious medical conditions. Anyone who has suffered through having a kidney stone, for example, could tell you how painful it is. Kidney stones are composed of precipitated calcium oxalate. The precipitated crystals form in the kidneys and move through the urinary system.

In this chapter you will examine some of the different types of reactions that can occur in solution. You will see how these reactions can be used to produce useful consumer products. You will also see how differences in solubility can be used to remove undesirable substances from solutions.

STARTING **POINTS**

Answer the following questions using your current knowledge. You will have a chance to revisit these questions later, applying concepts and skills from the chapter.

1. What is "safe" drinking water? Who should set the criteria for what makes drinking water safe?
2. How can you identify substances dissolved in a solution?
3. Why are most consumer products tested regularly? Who should be responsible for the cost of this testing?
4. How can you predict the quantity of product expected when two dissolved substances react?

Mini Investigation

Save the Suds

Skills: Performing, Observing, Communicating

SKILLS HANDBOOK A1.2, A2.4

Water is "hard" if it contains an unusually high concentration of calcium and/or magnesium ions. Water with low concentrations of these ions is "soft." In this investigation you will compare the effect of water hardness on soap.

Equipment and Materials: chemical safety goggles; apron; three test tubes and stoppers; test-tube rack; dropper bottles of distilled water, tap water, hard water, and liquid hand soap

1. Put on your safety goggles and lab apron.
2. Place test tubes in the test-tube rack.
3. Add 5 mL of distilled water to the first test tube, 5 mL of tap water to the second, and 5 mL of hard water to the third.
4. Add one drop of liquid hand soap to each of the test tubes.
5. Stopper each test tube and shake them for 1 min.

A. Compare the quantity of foam in the three test tubes. T/I

B. Which sample of water would be the best at cleaning clothes? A

9.1 Reactions of Ions in Solution

Thick, nasty, and chalky. These are just some of the adjectives patients use to describe the liquid they consume prior to an X-ray of the gastrointestinal tract (GI tract). The unappealing liquid is a mixture of mostly barium sulfate and water called a "barium meal." Consuming this mixture coats the lining of the GI tract with barium sulfate. Barium sulfate crystals block X-rays and make the GI tract much clearer in the image (**Figure 1**). This makes it easier to detect abnormalities in the GI tract.

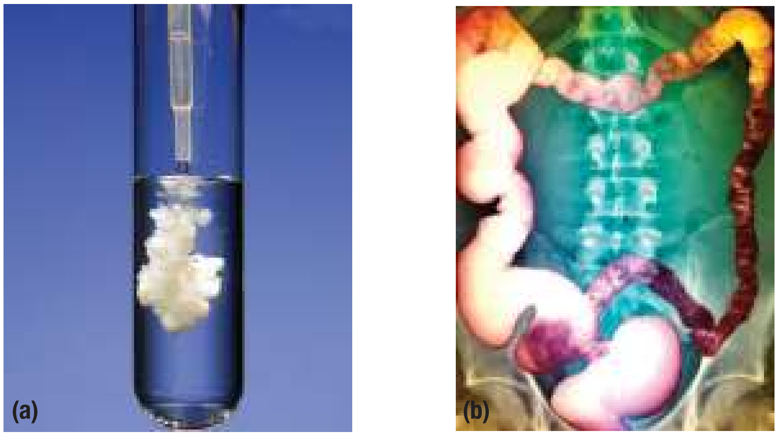

Figure 1 (a) Mixing solutions of barium chloride and sodium sulfate produces a barium sulfate precipitate and aqueous sodium chloride. (b) This barium X-ray shows that the right side of the patient's GI tract is constricted due to muscle spasms caused by irritable bowel syndrome.

CAREER LINK

X-rays of the GI tract and other parts of the body are performed by radiological technicians. They work as part of a healthcare team. To learn more about this rewarding career,

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What makes the consumption of a barium meal somewhat unusual is that dissociated barium ions are a well-known toxin. In fact, some schools in Ontario have restricted the use of barium compounds in science laboratories for that reason. If barium ions are toxic, why are barium meals safe to consume? To answer this question we have to consider the solubility of barium sulfate. According to **Table 1**, barium sulfate is only slightly soluble. This means that only a small percentage of the barium sulfate dissociates into barium ions and sulfate ions:

$$BaSO_4(s) \rightarrow Ba^{2+}(aq) + SO_4^{\ 2-}(aq)$$

Consequently, a barium meal contains too few barium ions to make it toxic. Its taste and texture, however, remain quite "nasty."

Table 1 Solubility of Ionic Compounds at Room Temperature

Solubility	Ion	Exceptions
very soluble (aq) $\geq$0.1 mol/L	$NO_3^{\ -}$	none
	Cl^- and other halides	except with Cu^+, Ag^+, $Hg_2^{\ 2+}$, Pb^{2+}
	$SO_4^{\ 2-}$	except with Ca^{2+}, Ba^{2+}, Sr^{2+}, Hg^{2+}, Pb^{2+}, Ag^+
	$C_2H_3O_2^{\ -}$	Ag^+
	Na^+ and K^+	none
	$NH_4^{\ +}$	none
slightly soluble (s) $<$0.1 mol/L	$CO_3^{\ 2-}$	except with Group 1 ions and $NH_4^{\ +}$
	$PO_4^{\ 3-}$	except with Group 1 ions and $NH_4^{\ +}$
	OH^-	except with Group 1 ions, Ca^{2+}, Ba^{2+}, Sr^{2+}
	S^{2-}	except with Groups 1 and 2 ions and $NH_4^{\ +}$

Equations for the Reaction of Ions

The barium sulfate required for a barium meal can be produced in the laboratory by mixing solutions of a soluble barium compound like barium chloride with sodium sulfate. The chemical equation for this reaction is

$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2\ NaCl(aq)$$

Recall that we can predict whether a compound precipitates or remains dissolved by consulting a solubility table (Table 1). In this case, the precipitate must be barium sulfate, since it is the only compound that is slightly soluble. All the others are very soluble.

The chemical equation for this reaction is sometimes called a **formula equation** since it gives the formula of each reactant and product. However, the formula equation can be somewhat misleading because three of the four compounds are soluble and are dissociated into ions. In other words, sodium chloride does not exist as an intact compound in the solution.

Total Ionic Equations

Another way to describe this reaction is to represent each soluble compound by its dissociated ions. The formula for barium sulfate remains intact, however, since very little of the barium sulfate dissolves: most of the barium sulfate remains in the form of $BaSO_4(s)$. Thus, the reaction of barium chloride with sodium sulfate may also be represented by the equation

$$Ba^{2+}(aq) + 2\ Cl^-(aq) + 2\ Na^+(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s) + 2\ Na^+(aq) + 2\ Cl^-(aq)$$

A chemical equation in which soluble ionic compounds are written in their dissociated form is called a **total ionic equation**. **Figure 2** shows the reaction of the ions to form barium sulfate.

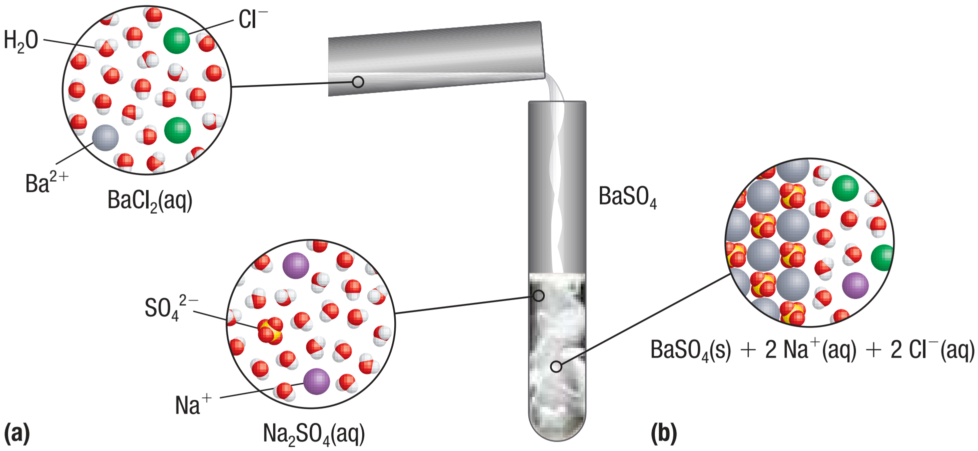

Figure 2 (a) Barium chloride solution is poured into sodium sulfate solution. (b) The barium sulfate forms a precipitate but sodium and chloride ions are left in solution.

Like all chemical equations, a total ionic equation must be balanced. To be balanced, the equation must meet two important criteria:

1. All entities on one side of the equation must be accounted for on the other. That is why the coefficient "2" appears before both Na^+ and Cl^- in the equation.
2. The net electrical charge on both sides of the equation must be the same.

In the barium sulfate reaction, the sum of the charges on each side of the equation equals zero (**Figure 3**).

$$\underbrace{Ba^{2+}(aq) + 2\ Cl^-(aq) + 2\ Na^+(aq) + SO_4^{2-}(aq)}_{\underbrace{1\,(+2) \quad + \quad 2\,(-1) \quad + \quad 2\,(+1) \quad + \quad 1\,(-2)}_{0}} \rightarrow BaSO_4(s) + \underbrace{2\ Na^+(aq) + 2\ Cl^-(aq)}_{\underbrace{2\,(+1) \quad + \quad 2\,(-1)}_{0}}$$

Figure 3 In any balanced chemical equation, the total number of atoms of each element must be the same on both sides of the arrow. The sum of the charges must also be equal on both sides.

LEARNING **TIP**

Precipitation Reactions
Most precipitation reactions involving two ionic compounds as reactants are double displacement reactions. They follow with the general pattern
AB + CD → AD + CB

formula equation a chemical equation in which all compounds are represented by their chemical formulas

total ionic equation a chemical equation in which all highly soluble ionic compounds are written as dissociated ions

Net Ionic Equations

spectator ions ions that are not involved in a chemical reaction

Notice that the sodium and chloride ions in the total ionic equation appear on both sides of the equation as ions. This means that they remain unchanged. Ions that do not participate in the reaction are called **spectator ions**. These ions can be omitted from the total ionic equation to give a simpler equation called a net ionic equation.

Total ionic equation:

$$Ba^{2+}(aq) + \cancel{2Cl^-(aq)} + \cancel{2Na^+(aq)} + SO_4^{2-}(aq) \rightarrow BaSO_4(s) + \cancel{2Na^+(aq)} + \cancel{2Cl^-(aq)}$$

Net ionic equation:

$$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$$

net ionic equation a chemical equation that includes only the entities that react during the reaction

The **net ionic equation** describes the chemical change that occurs in a reaction involving ions. It does not, however, specify which substances were mixed to result in this reaction. Note that no net reaction takes place if all reactants and products are soluble and remain as ions in solution.

Net ionic equations are not limited to double displacement reactions. For example, in Chapter 4 you learned that zinc metal undergoes a single displacement reaction when placed in a solution of copper(II) sulfate (**Figure 4**). Since zinc is above copper in the activity series, it displaces copper from copper(II) sulfate. The chemical equation for this reaction is

$$Zn(s) + CuSO_4(aq) \rightarrow Cu(s) + ZnSO_4(aq)$$

Since both sulfate compounds are very soluble, they are completely dissociated in solution. Therefore, the total ionic equation for this reaction is

$$Zn(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Cu(s) + Zn^{2+}(aq) + SO_4^{2-}(aq)$$

Since sulfate ions are common to both sides, this equation simplifies to the following net ionic equation:

$$Zn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Zn^{2+}(aq)$$

Notice that, in the previous net ionic equation, the charge on each side of the arrow is not zero. However, the charges on each side of the arrow are the same: +2.

Figure 4 In this reaction of zinc metal with copper(II) ions, the solution is lighter blue near the bottom where most of the reddish copper metal has formed. This is evidence that copper ions are being used up.

Tutorial 1 Writing Total Ionic Equations and Net Ionic Equations

Here are some useful tips for writing total ionic equations and net ionic equations. As you get proficient at writing these equations, you will be able to combine some of these steps.

1. Write the balanced chemical equation for the reaction.
2. Use the solubility table (Table 1) to determine whether a compound precipitates.
3. Write the total ionic equation by showing all of the soluble compounds as ions.
4. Cancel spectator ions.
5. Write the net ionic equation.

Sample Problem 1: Writing a Net Ionic Equation from a Formula Equation

Write the total ionic equation and net ionic equation for the reaction that occurs when solutions of lead(II) nitrate and potassium iodide are combined.

Step 1. Write the formula equation for the reaction, omitting the state symbols of the products.

$$Pb(NO_3)_2(aq) + 2\ KI(aq) \rightarrow PbI_2 + 2\ KNO_3$$

Step 2. Use the solubility table to determine whether a compound precipitates.

The only reaction product of low solubility is PbI_2. All others are highly soluble. Therefore the formula equation is

$$Pb(NO_3)_2(aq) + 2\ KI(aq) \rightarrow PbI_2(s) + 2\ KNO_3(aq)$$

Step 3. Write the total ionic equation by showing all of the soluble compounds as ions.

$Pb^{2+}(aq) + 2\ NO_3^-(aq) + 2\ K^+(aq) + 2\ I^-(aq) \rightarrow PbI_2(s) + 2\ K^+(aq) + 2\ NO_3^-(aq)$

Step 4. Cancel spectator ions.

$Pb^{2+}(aq) + \cancel{2\ NO_3^-(aq)} + \cancel{2\ K^+(aq)} + 2\ I^-(aq) \rightarrow PbI_2(s) + \cancel{2\ K^+(aq)} + \cancel{2\ NO_3^-(aq)}$

Step 5. Write the net ionic equation.

$Pb^{2+}(aq) + 2\ I^-(aq) \rightarrow PbI_2(s)$

Sample Problem 2: Single Displacement Reactions

Write the net ionic equation for the reaction that occurs when aluminum metal is placed in a solution of copper(II) chloride, $CuCl_2$.

Step 1. Write the formula equation for the reaction, omitting the state symbols of the products.

Since aluminum is above copper on the activity series, copper metal is displaced from $CuCl_2$. Therefore, the formula equation for this reaction is

$2\ Al(aq) + 3\ CuCl_2(aq) \rightarrow 3\ Cu + 2\ AlCl_3$

Step 2. Use the solubility table to determine whether a compound precipitates.

Both chloride compounds are highly soluble. Metals are solids. Therefore, the formula equation is

$2\ Al(s) + 3\ CuCl_2(aq) \rightarrow 3\ Cu(s) + 2\ AlCl_3(aq)$

Step 3. Write the total ionic equation.

$2\ Al(s) + 3\ Cu^{2+}(aq) + 6\ Cl^-(aq) \rightarrow 3\ Cu(s) + 2\ Al^{3+}(aq) + 6\ Cl^-(aq)$

Step 4. Cancel spectator ions.

$2\ Al(s) + 3\ Cu^{2+}(aq) + 6\ \cancel{Cl^-(aq)} \rightarrow 3\ Cu(s) + 2\ Al^{3+}(aq) + 6\ \cancel{Cl^-(aq)}$

Step 5. Write the net ionic equation.

$2\ Al(s) + 3\ Cu^{2+}(aq) \rightarrow 3\ Cu(s) + 2\ Al^{3+}(aq)$

Practice

1. Write the formula equation, the total ionic equation, and the net ionic equation for each of the following reactions: K/U C
   (a) $AgNO_3(aq) + NaF(aq) \rightarrow$
   (b) $CuCl_2(aq) + KOH(aq) \rightarrow$
   (c) $AgNO_3(aq) + Zn(s) \rightarrow$
2. Water is considered to be "hard" if it has an unusually high concentration of calcium or magnesium ions. Excess calcium ions can be removed by adding a solution of potassium carbonate. Write the net ionic equation for the reaction that occurs. K/U C

9.1 Summary

- The reaction of ions in solution can be represented by formula equations, total ionic equations, and net ionic equations.
- Formula equations list all reactants and products as either elements or compounds.
- Total ionic equations show all highly soluble compounds dissociated in their ions.
- Net ionic equations show only the ions involved in the reaction. The net ionic equation is obtained by removing the spectator ions from the total ionic equation.
- Spectator ions are ions that are not involved in the reaction.

9.1 Questions

1. Write and balance the formula, total ionic, and net ionic equations for the following reactions: K/U C
   (a) $ZnCl_2(aq) + K_2CO_3(aq) \rightarrow$
   (b) $Fe(NO_3)_3(aq) + NaOH(aq) \rightarrow$
   (c) $Cu(NO_3)_2(aq) + Al(s) \rightarrow$
2. The reaction of tin metal with a solution of silver nitrate, $AgNO_3(aq)$, releases $Sn^{2+}(aq)$ ions into the solution. K/U C
   (a) Write the formula equation for this reaction.
   (b) Write the net ionic equation for this reaction.
   (c) Tin ions are toxic. Suggest a soluble compound that could be added to precipitate all the tin ions from the solution.
   (d) Write a net ionic equation for the reaction that occurs in (c).
3. For each of the equations below, identify the product with the lowest solubility, identify the spectator ions, and write the net ionic equation. K/U T/I C
   (a) $FeBr_2(aq) + Na_2S(aq) \rightarrow FeS + 2\ NaBr$
   (b) $Ba(OH)_2(aq) + (NH_4)_2CO_3(aq) \rightarrow BaCO_3 + 2\ NH_4OH$
4. Why do you think the term "spectator" was chosen to describe spectator ions? T/I
5. Chlorine gas, $Cl_2(g)$, bubbled into a potassium iodide solution produces potassium chloride and aqueous iodine. Iodine is a molecular element and does not remain as ions in the solution. Write the net ionic equation for this reaction (**Figure 5**). K/U C

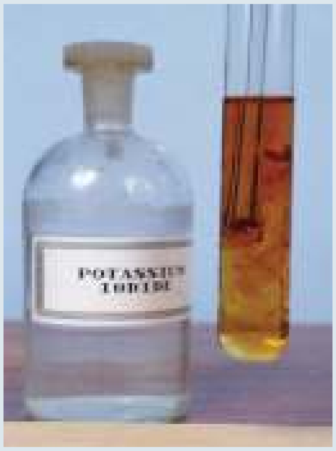

Figure 5 Aqueous iodine is visible as brown swirls in the liquid.

6. Some groundwater naturally contains unusually high concentrations of iron(III) ions. Excess iron is a nuisance because it affects the taste of the water and causes rust stains to form on plumbing fixtures (**Figure 6**). The addition of a soluble hydroxide compound precipitates much of the iron from water. K/U T/I C
   (a) Suggest two compounds that may be used.
   (b) Write the net ionic equation that occurs.

Figure 6 The brown stain inside this toilet tank is caused by a high concentration of iron(III) ions in the water supply.

7. Many municipalities use chlorine gas to disinfect drinking water. Chlorine is made industrially by passing electricity through a sodium chloride solution. The chemical equation for this reaction is

$$2\ NaCl(aq) + 2\ H_2O(l) \xrightarrow{\text{electricity}} Cl_2(g) + H_2(g) + 2\ NaOH(aq)$$

   Note that two of the products do not remain in solution. Write the net ionic equation for this reaction. K/U C
8. One way of removing excess calcium ions from water is to precipitate them as calcium phosphate. K/U T/I C
   (a) Use the solubility table to identify three suitable compounds that could be dissolved in water to make a solution that contains aqueous phosphate ions.
   (b) Write a net ionic equation for the precipitation of calcium ions with phosphate ions.
9. The inside of an aluminum pop can is coated with a plastic liner to separate the aluminum from the contents of the can. It is possible to remove all of the aluminum from the can by reacting it with copper(II) chloride solution. Only the plastic liner remains. K/U T/I C
   (a) Write a balanced chemical equation for this reaction.
   (b) Use the solubility table to suggest a soluble compound that could be used to precipitate the aluminum ions produced in (a).
10. Kidney stones are mostly made up of calcium oxalate, CaC_2O_4. K/U C A
   (a) Write the net ionic equation of the formation of calcium oxalate from its ions.
   (b) Research the symptoms and causes of the formation of kidney stones.
   (c) Identify some of the treatment options available to patients with kidney stones.

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Water Treatment

9.2

The early 1800s saw the introduction of two important innovations in the wealthy homes of London, England: indoor running water and flush toilets. These modern conveniences, however, came with some serious drawbacks. There were no sewage treatment plants at the time. Instead, domestic wastewater flowed directly into the sewer system and ultimately into the River Thames—an important source of drinking water. Meanwhile, a cholera epidemic gripped London. We now know that cholera is transmitted through contact with food or water contaminated by a bacterium found in human feces (**Figure 1**). However, in the mid-1800s, people had no idea that micro-organisms could cause disease. By 1849, the death rates of people who relied on the Thames for water had risen to 200 for every 10 000 people.

In 1852, two water companies were competing to supply water to the citizens on the south side of the Thames. To attract more customers, one company tried to improve the quality of its water. This was done by moving the water intake further upstream, to a location above the sewer discharge. This change resulted in a 90 % reduction in mortality rate compared to the customers of the other company.

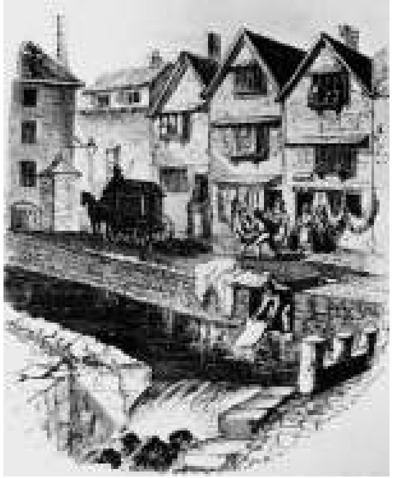


Figure 1 The cholera epidemic in England during the 1830s may have been spread by people washing linen, soiled by cholera victims, in the local drinking-water supply. Cholera is caused by a bacterium that is spread by contaminated water.

Sanitation and Water Quality

Proper sanitation is critical for safe drinking water. Sanitation is the safe disposal of human waste. Unfortunately, there are still many places in the world—including some parts of Canada—where poor sanitation threatens the quality of drinking water. In fact, approximately 2.6 billion people do not have access to basic sanitation. An estimated 1.5 million children die each year as a result of water-borne diseases related to poor sanitation. A cholera epidemic, similar to the one in the 1800s in London, affected Haiti in 2010. Haiti, already one of the poorest countries in the world, was seriously damaged by an earthquake early that year. Almost 2 000 000 people became homeless. The earthquake destroyed much of the water and sewage infrastructure (**Figure 2**).

Improving water quality through better sanitation is an important first step to eliminating water-borne diseases. Improvements in sanitation also have many benefits. According to the United Nations University, every dollar invested in improving water quality and sanitation in developing countries results in a return of between 3 and 30 dollars in productivity. As people become healthier they require less healthcare and can work or attend school regularly.

Improvements in sanitation can also have a direct impact on education. For example, international charities are finding that the construction of school latrines in developing countries significantly improved both children's health and their school attendance (**Figure 3**).

Figure 2 After the 2010 earthquake, many Haitians moved to crowded temporary camps without access to fresh water and effective sanitation. Thousands fell ill with cholera and other water-borne diseases.

Figure 3 Latrines, like these under construction at Madaraka Public School, Kenya, provide students with a safe, private, and dignified place to eliminate bodily wastes.

WEB LINK

The United Nations has declared that clean water and sanitation are basic human rights. To read more about the UN's position on this issue,

GO TO NELSON SCIENCE

Types of Water Contaminants

Whenever a large number of people concentrate in one area, it is inevitable that some of their waste contaminates local sources of water. Industries also add pollutants to water. The challenge is to keep the contamination to a minimum because once the water is contaminated it is difficult and expensive to treat. Water contaminants can be classified as being physical, biological, or chemical. Note that physical and biological contaminants are not in solution. Rather, they form heterogeneous mixtures with water.

PHYSICAL CONTAMINANTS

Physical contaminants can be floating debris like twigs and discarded plastic products. They can also include fine suspended particles like sand, clay, and decaying organic material, which make water cloudy. Fine particles can give water an unpleasant appearance, taste, or odour. Suspended particles are also a nuisance because they can interfere with the disinfection of water during treatment. Oil and petroleum products floating on water also give a foul odour and can contaminate shorelines and wildlife.

BIOLOGICAL CONTAMINANTS

Bacteria, viruses, and protozoa are the major biological contaminants in water. Many of these micro-organisms exist naturally in water. Some are introduced to water through waste materials. For example, *E. coli* are bacteria found in feces. High levels of these organisms make water unfit for bathing or consumption. A common source of *E. coli* contamination of lake water is overloaded or leaking sewers. As a precaution, beaches are closed when the levels of these organisms are too high (**Figure 4**).

Agriculture can be another important source of *E. coli* contamination. During heavy rains, the runoff from piles of animal manure can contaminate nearby bodies of water with *E. coli*. Preventing this form of contamination is a particular challenge for large-scale cattle, pork, or poultry farms (**Figure 5**).

Figure 4 Beaches are closed when water quality technicians find high levels of *E.coli* in the water. An *E. coli* infection can cause earaches, diarrhea, and vomiting.

Figure 5 Large-scale poultry farms generate a great deal of waste.

In addition to being a hazard to human health, fecal matter affects aquatic ecosystems. Fecal matter is rich in plant nutrients, including ammonia, nitrogen, and phosphorus. This can result in a short-term growth in the populations of aquatic plants and algae. When these organisms die, decomposers deplete much of the dissolved oxygen in the water as they decompose the dead plants and algae. The reduction of oxygen can threaten fish populations because fish need oxygen to survive.

In the spring of 2000, Canada's worst-ever *E. coli* outbreak occurred in Walkerton, Ontario. Seven people died and hundreds more became sick. The source of the contamination was cattle manure with a particularly deadly strain of *E. coli*. Farm runoff containing bacteria had seeped into the ground and contaminated an aquifer that was an important source of drinking water for the town. The investigation into the Walkerton tragedy concluded that it was preventable: the levels of chlorine used to disinfect the water had been inadequate.

CHEMICAL CONTAMINANTS

Chemical contaminants are substances that are soluble in water. These can include industrial and consumer chemicals, ions from mines and landfill sites, agricultural chemicals like fertilizers and pesticides, and other water-soluble organic substances. These chemicals all form homogeneous mixtures—solutions—with water.

Leachate is fluid that has passed through solid waste material and has extracted suspended or dissolved substances from mines or dump-sites. Leachate can pollute nearby bodies of water or underground aquifers (**Figure 6**). For example, cadmium is a particularly toxic metal that can cause organ failure and fragile bones over the long term. Cadmium occurs naturally in certain mineral deposits. A significant proportion of cadmium in drinking water comes from the leachate draining from dump sites containing discarded nickel-cadmium batteries. **Table 1** summarizes some of the chemical contaminants in water and their sources.

leachate a liquid contaminated with dissolved or suspended substances picked up as the liquid passed through old mines or dump sites

Figure 6 The red colour of this canal in Manchester, England is caused by leachate containing iron oxides from a closed mine.

Table 1 Sources of Chemical Contaminants in Water

Contaminants	Source	Effect on human health or the environment
heavy metals (such as lead, mercury, chromium, cadmium)	leachate from mines and landfill; leakage from industrial processes such as electroplating	interfere with the development of the brain and nervous system; some are carcinogenic
organic solvents	petroleum-based products; cleaners	poisonous or carcinogenic
nitrates and phosphates	residential and agricultural fertilizers	toxic to many animals, including humans; promotes growth of aquatic plants resulting in algal blooms that deplete the water of oxygen
pesticides	residential and agricultural use of pesticides	toxic; some can accumulate in the food chain
salt	runoff from roads and sea spray	kills freshwater organisms; makes water unfit for drinking
sulfur compounds	naturally occurring minerals; bacteria	no health effect, but an unpleasant smell
synthetic hormones and other medications	household sewage	uncertain

Waste metal leachate from abandoned mines has posed a threat to several Canadian communities. For example, in 2005, scientists completed a study of a Cree community in Northern Québec to determine whether metal leachate from an abandoned mine posed a threat to the community. The researchers collected and analyzed blood, urine, and hair samples for metal contaminants. They also collected and analyzed dietary information from the community. Fortunately, the study found no significant levels of leachate metals in the participants. However, it did detect elevated levels of mercury contamination—likely from eating fish caught in the area. Mercury is a concern because it may impair brain development, particularly in children.

Establishing Water Quality

In Canada, three levels of government work cooperatively to ensure that the water that comes out of your tap is safe to drink. Health Canada sets the guidelines for what should and should not be in drinking water. These guidelines address levels of biological and chemical contaminants. They also deal with aesthetic

WEB LINK

To check out maximum acceptable concentrations of contaminants in water,

GO TO NELSON SCIENCE

Investigation 9.2.1

How Effective Are Water Filters? (p. 450)
In this investigation you will conduct chemical tests to see if a cartridge-type water filter can remove dissolved ions from the water.

characteristics of water like taste, clarity, and odour. Health Canada's guidelines are based on ongoing research and are reviewed regularly to ensure that they are up to date.

Included in Health Canada water quality guidelines are the maximum acceptable concentrations (MAC) of several hundred substances in drinking water (**Table 2**). For example, water whose concentration of cadmium exceeds 0.005 ppm (parts per million) is considered unfit for consumption.

Table 2 Maximum Acceptable Concentrations of Selected Chemicals in Canadian Drinking Water

Substance	Typical source	MAC (ppm)
arsenic	mining waste; industrial effluent	0.010
benzene	industrial effluent; spilled gasoline	0.005
cadmium	leachate from landfill	0.005
cyanide	mining waste; industrial effluent	0.2
lead	leachate from landfill; old plumbing	0.010
mercury	industrial effluent; agricultural runoff	0.001
nitrate	agricultural runoff	45
tetrachloroethene	dry cleaners	0.03
trichloromethane	water chlorination	0.1

Health Canada's guidelines are only recommendations. Provincial and territorial governments use these guidelines to set their own standards for water quality. They are also responsible for the testing program that ensures contaminants in drinking water do not exceed these standards. Municipalities are responsible for the day-to-day operation of the water treatment and distribution networks that deliver potable water to your tap.

The Water Treatment Process

Most places in Canada have safe, high-quality tap water. But not all tap water is created equal. That is why the tap water may taste or feel different from region to region. The composition of tap water depends on where the water comes from (such as a lake or aquifer) and what it dissolved on the way to your tap. It also depends on how the water is treated.

Water treatment can be as simple as adding disinfection tablets to a bottle of lake water on a camping trip, or as complex as the process at a municipal water treatment plant. The ultimate goal of water treatment is not to purify water but to make the water potable, or safe to drink. This means removing or reducing the substances in water so that they no longer pose a threat to public health. Treatment also involves ensuring that the water has an appropriate taste, colour, and odour.

People in rural areas likely get their drinking water directly from a nearby well. The well's owner is responsible for testing the water to ensure that it is potable. The residents of an urban area usually get their water from a municipal or regional water authority. This water is processed in several steps to ensure that it is potable. A large municipality treats and pumps vast quantities of water each day. For example, the City of Toronto produces about 1300 ML per day—more than enough to fill a major sports stadium. Processing and pumping this much water uses a lot of energy. Water treatment is therefore one of the largest energy costs of any municipality.

CAREER LINK

A water treatment engineer is responsible for the quality of the water leaving the water treatment facility. To learn more about this very important and responsible career,

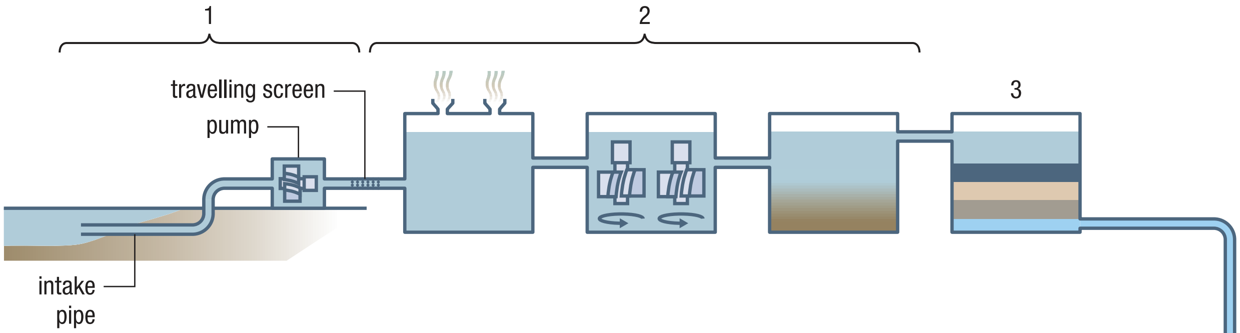
GO TO NELSON SCIENCE

Figure 7 (on the next page) outlines the key steps involved in taking water from its source and transforming it into the potable product that comes out of a tap in the city.

1. Collection: Travelling screens remove debris and other large particles as the water enters the treatment plant.

2. Coagulation, Flocculation, and Sedimentation: Chemicals known as coagulants mix with the water, causing the small particles in the water to clump together. Flocculation, which involves gentle mixing, forms a light, fluffy precipitate called a floc. During sedimentation the floc sinks very slowly, taking suspended particles to the bottom of the tank and clearing the water.

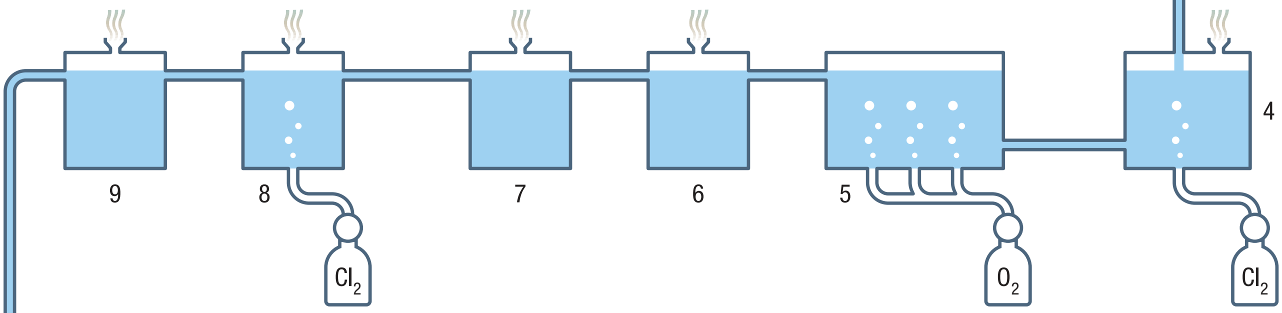
3. Filtration: The water flows by gravity through efficient filters made up of layers of sand and anthracite (carbon). This removes the remaining floc, other chemical and physical impurities, and most of the biological impurities (bacteria, etc.).



6. Softening: Sodium carbonate and calcium hydroxide or a phosphate may be added to reduce water hardness by precipitating the calcium and magnesium ions in the water.

5. Aeration: Air, ozone, oxygen, "activated" charcoal, ammonia, chlorine dioxide, or potassium permanganate may be mixed with the water to further reduce taste and colour problems.

4. Disinfection: A disinfectant such as chlorine kills micro-organisms. Alternative disinfectants include ozone or chlorine dioxide, ammonia, potassium permanganate, and even ultraviolet light.



9. Ammoniation: Ammonia stabilizes the chlorine so that it remains dissolved in the treated water for longer periods of time.

8. Post-chlorination: A final chlorine disinfection treatment kills any remaining micro-organisms. The pH level in the water is adjusted to be slightly basic because even slightly acidic water will corrode metallic pipes.

7. Fluoridation: Some municipalities add a small quantity of fluoride to drinking water because it makes the enamel layer of teeth more resistant to decay.

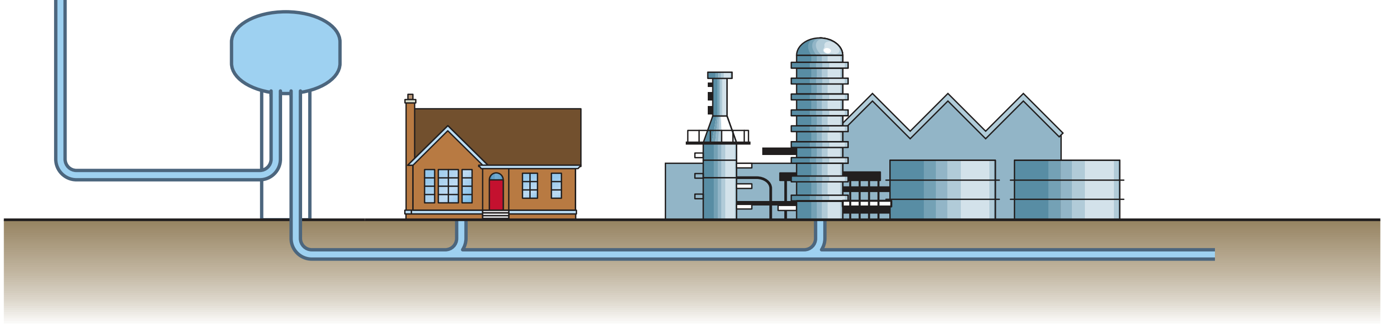


Figure 7 These are the steps that are typically involved in a municipal water treatment process. Some of these steps may be omitted depending on the characteristics of the source water.

Research This

What Is in Your Water?

Skills: Researching, Evaluating, Communicating — SKILLS HANDBOOK A5.1

Municipalities test drinking water regularly to ensure that contaminant concentrations do not exceed the maximum acceptable concentrations set by the province or territory. These water quality data are available to the public on the municipality website.

1. Find the water quality reports for your municipality or a municipality near you. Locate the most recent test results.
2. Find the maximum acceptable concentrations (MAC) of contaminants from the Health Canada website if they are not given in the municipality's report.

A. Identify one biological contaminant mentioned in the report.

B. Identify two inorganic substances and two organic substances in the water. Compare the concentrations of these substances with their MAC, as suggested by Health Canada. How did the concentrations of these substances compare with the MAC?

C. Give an example of data that you found to be particularly interesting or surprising.

D. Based on this report, evaluate the quality of the water from this source.

E. Search for comparable information on a locally available brand of bottled water. Using your findings, decide whether tap water or bottled water is safer to drink.

GO TO NELSON SCIENCE

hard water water that has a relatively high concentration of Ca^{2+} and Mg^{2+}

soft water water that has a relatively low concentration of Ca^{2+} and Mg^{2+}

WATER SOFTENING

If you have ever washed your hair with rainwater, you may have noticed that it is shinier and more manageable than if you used tap water. The difference is that tap water is usually "harder" than rainwater. **Hard water** contains relatively high concentrations of the ions Ca^{2+}, Mg^{2+}, and occasionally Fe^{3+}. These ions dissolve in the water as it moves through the ground. Water with low concentrations of these ions is called **soft water**. Hard and soft water look and feel the same but they react differently with soap. Hard water does not lather and leaves behind a soap scum (**Figure 8**). The scum is actually a precipitate formed when cations in the water combine with soap anions. A typical net ionic equation for this reaction is

$$\underset{}{Ca^{2+}(aq)} + \underset{\text{(soap ion)}}{2\,CH_3(CH_2)_{16}CO_2^{-}(aq)} \rightarrow \underset{\text{(bathtub ring precipitate)}}{Ca(CH_3(CH_2)_{16}CO_2)_2(s)}$$

Hard water is also a nuisance because it forms precipitates that clog hot-water pipes and the heating elements of water heaters and kettles (**Figure 9**).

UNIT TASK BOOKMARK

In the Unit Task on page 498 you will soften a sample of hard water and then test the effectiveness of your softening method.

Figure 8 Hard water and soap react to leave a precipitate in sinks and bathtubs.

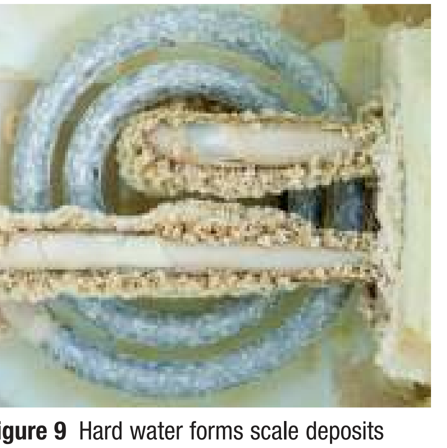


Figure 9 Hard water forms scale deposits inside hot-water pipes and on kettle heating elements.

Detergents were specifically designed so that they do not form precipitates as often in hard water. For this reason, most shampoos contain detergents rather than soaps.

One way to deal with hard water is to remove the cations that cause the water to be hard. A home water-softening device does this by exchanging the "hardness" ions with sodium ions. A typical home water softener consists of a large tank full of beads of an ion-exchange resin (**Figure 10(a)**). The resin is a plastic-like material made up of very long ionic compounds containing many groups of atoms that are similar to polyatomic ions. The charge of each group is balanced by a sodium ion. As hard water passes through the resin, calcium ions, $Ca^{2+}(aq)$, displace sodium ions, $Na^{+}(aq)$, and bind to the polyatomic groups (**Figure 10(b)**). Eventually, the resin becomes saturated with hard water ions (**Figure 10(c)**). The resin can be recharged with sodium by passing a concentrated sodium chloride solution through it.

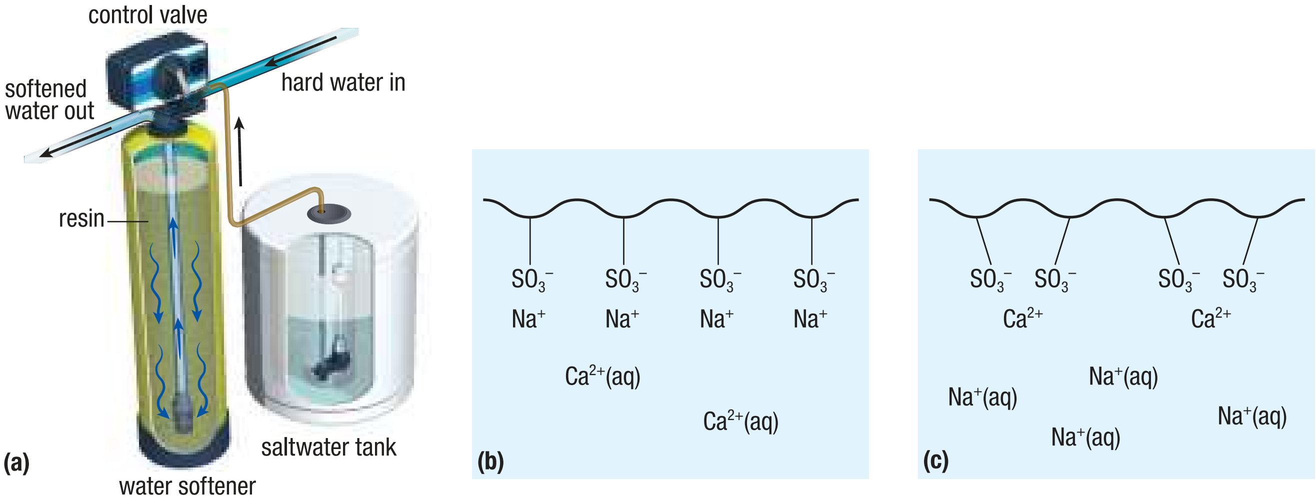


Figure 10 (a) Hard water fills the tank of an ionic exchange water softener and passes through the resin. (b) Sodium ions bonded to the resin are displaced by calcium ions in the water. (c) By the time the water gets to the bottom of the tank it is soft because it has fewer calcium ions.

Mini **Investigation**

Testing a Cationic Exchange Resin

Skills: Questioning, Controlling Variables, Performing, Observing, Analyzing, Evaluating, Communicating

SKILLS HANDBOOK A1.2, A2.4

As water passes through a cationic exchange resin, the ions responsible for water hardness are replaced with sodium or potassium ions. In this activity, you will conduct a chemical test to compare the hardness of water before and after it passes through a cationic exchange resin. Ethylenediamine tetracetic acid (EDTA) is an anion that reacts with the "hardness" ions: calcium and magnesium. Once this reaction is complete, the indicator Eriochrome Black T changes colour from pink to blue.

Equipment and Materials: chemical safety goggles; lab apron; resin pipette from Investigation 9.2.1; test tube; test-tube rack; 10 mL graduated cylinder; spot plate; toothpick; Eriochrome Black T indicator (solid); dropper bottles of tap water, pH 10 buffer solution, and EDTA (0.01 mol/L)

pH 10 buffer solution, Eriochrome Black T, and EDTA are all irritants to the eyes, skin, and respiratory system. Avoid skin and eye contact. If you spill any of these substances on your skin, wash the affected area with plenty of cool water and inform your teacher.

EDTA is toxic if ingested. If you ingest this substance, inform your teacher immediately. Wash your hands thoroughly after this investigation.

1. Put on your chemical safety goggles and lab apron.
2. Place the resin pipette in a test tube in the test-tube rack.
3. Pass about 5 mL of tap water through the resin.
4. Place 5 drops of the original water sample in one well of the spot plate.
5. Place 5 drops of the filtered water into another well.
6. Add 1 drop of pH 10 buffer solution to each of the two wells.
7. Under your teacher's supervision, add a few grains of solid Eriochrome Black T indicator to each well. (Use less than will fit on the end of a toothpick.)
8. Add EDTA to each well and count the number of drops required to turn the Eriochrome Black T from pink to blue.

A. Compare the hardness of the two water samples. T/I
B. Why was the indicator necessary? T/I
C. How did you control the variables in this activity? T/I

UNIT TASK **BOOKMARK**

The Unit Task (page 498) is very similar to the Mini Investigation on page 435. Both investigations analyze hard and soft water, both use EDTA in a reaction with hard water ions, and both use Eriochrome Black T as an indicator. Consider the Mini Investigation as practice for the Unit Task.

9.2 Summary

- Water contaminants can be physical, biological, or chemical.
- Leachate is a fluid carrying waste material extracted from mines or dump sites.
- The purpose of water treatment is to make water potable or safe to drink, not to purify water.
- Federal, provincial, and municipal governments work together to ensure that drinking water is safe to drink.
- Water with contaminants that exceed the maximum acceptable concentrations (MAC) is unsafe to drink.
- Hard water contains unusually high concentrations of calcium and magnesium ions. Soft water contains low concentrations of these metal ions.
- Hard water is a nuisance because it forms precipitates that clog pipes and appliances. It also produces a scum-like precipitate with soap.
- Water can be softened by passing it through an ion exchange resin.

9.2 Questions

1. List the three categories of water contaminants. Give one example of each. K/U
2. Why are sanitation and the quality of drinking water often related? T/I
3. Tetrachloroethene (also called perchloroethylene or perc) is an organic solvent widely used in the dry-cleaning industry because of its ability to dissolve grease. Unlike many organic substances, the density of perc is greater than the density of water. T/I A
   (a) What type of water contaminant is perc?
   (b) Based on the fact that perc can dissolve grease, is perc a polar or non-polar compound? Explain.
   (c) Equal volumes of perc and water are added to a beaker. What would you expect to observe in the beaker? Why?
   (d) Could a perc spill in a lake be treated in the same way as cleaning up an oil spill? Explain.
4. Why do you think it is important that the maximum acceptable concentrations (MAC) of water contaminants be reconsidered from time to time? T/I
5. Assuming that tap water contains the maximum acceptable concentration of lead given in Table 2, what is the mass of lead in a 270 L bathtub full of water? T/I
6. Use the water treatment diagram (Figure 7) to answer the following questions:
   (a) What is floc? Why is floc formation used in water treatment?
   (b) Identify two chemical alternatives and one non-chemical alternative to using chlorine to disinfect water.
   (c) Why is ammonia added to water late in the treatment process? T/I
7. Describe a simple chemical test that you could do to determine if your tap water is hard. T/I C
8. Water is softened on a large scale for industrial use by precipitating calcium ions as calcium carbonate. T/I C
   (a) Write the net ionic equation for this reaction.
   (b) Suggest an appropriate carbonate compound that could be the source of the carbonate ion. (Refer to Table 1 in Section 9.1.) Justify your choice.
9. (a) Why is a water softener often necessary in communities that get their water from underground sources?
   (b) Why is it necessary to regenerate the resin in a home water-softening system? K/U T/I
10. EDTA is an important additive in some detergents. EDTA reacts with calcium ions, which keeps calcium ions in solution. Why is this an advantage?
11. During a boil-water advisory, the local public health unit warns the public to boil tap water before drinking it. What type(s) of water contaminants does boiling eliminate? What type(s) of contaminants does boiling not eliminate? Why? T/I
12. Research a "green" way to remove hard-water scale from the inside of a kettle. If possible, write a chemical equation for the reaction that occurs. T/I C A
13. Chlorine is commonly used to disinfect water during the treatment process. A potential problem with the use of chlorine is that it can react with organic compounds in the water to form a class of chemicals called trihalomethanes. Some of the compounds are known carcinogens. T/I A
   (a) Research the potential health risks of trihalomethanes.
   (b) Evaluate the benefits and risks of using chlorine. In your opinion, do the benefits outweigh the risks? What criteria did you use to make this decision? Defend your decision.
14. Hair, toenails, and fingernails are often collected during research studies of toxins in the human body. What makes these materials good markers of toxins in the body? T/I A

GO TO NELSON SCIENCE

Chemical Analysis

9.3

"Your child has three times the normal concentration of lead in his blood." The parents of eight-month-old Yuri felt shock and disbelief as their doctor read the results of Yuri's blood test. The presence of lead in the blood can impair brain and nervous-system development. What could be the source of the lead, they wondered? It couldn't be their home. A building inspector had verified that all the old lead paint in the house had been correctly removed when they renovated their 80-year-old home. They remembered seeing reports in the news about the recall of toys contaminated with lead. Their suspicions were soon confirmed. Using a lead test kit purchased from the hardware store, they tested the paint on Yuri's favourite toy truck. Sure enough, the red colour indicated that lead was present (**Figure 1(a)**). Just to be sure, a health department official analyzed the toy using a device called an X-ray diffraction analyzer (**Figure 1(b)**). The analysis concluded the paint contained about 3000 ppm of lead—five times Health Canada's acceptable limit for lead in consumer products.

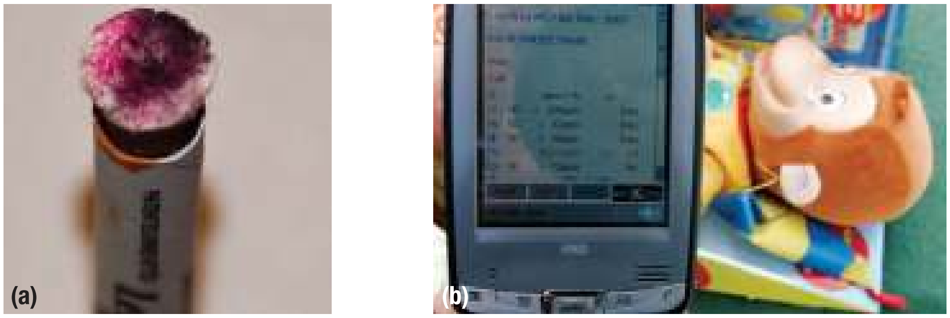


Figure 1 (a) The colour change observed using a home lead test kit indicates that lead may be present. (b) A technician tests a toy using an X-ray diffraction analyzer. This is a quantitative test to determine the concentration of lead in an object.

As you learned in Chapter 6, chemical analysis may be qualitative or quantitative. Qualitative analysis, you will recall, involves the identification of a substance. The colour change observed while using the home lead test kit is an example of qualitative analysis. Quantitative analysis involves determining how much of a substance is present. Determining the concentration of lead using X-ray diffraction is quantitative analysis.

Qualitative Analysis

SKILLS HANDBOOK B4

Many substances can be identified by a unique physical or chemical property. Nickel, for example, can be detected by the characteristic pink colour produced when it reacts with a compound called dimethylglyoxime, $C_4H_8N_2O_2$. As you learned in Chapter 4, skin allergies from exposure to nickel are becoming more common. The presence of nickel on metal objects like jewellery or the outer rim of a cellphone can be confirmed using a nickel test kit (**Figure 2**).

Figure 2 A qualitative test shows the presence of nickel ions in the casing of a watch.

WEB LINK

The acceptable limit for lead in children's toys is 600 ppm, according to Health Canada. To find out more about acceptable lead levels,

GO TO NELSON SCIENCE

UNIT TASK BOOKMARK

You will use a form of chemical analysis when you perform the Unit Task, which is described in detail on page 498. In this case, the test will involve a colour change.

Investigation 9.3.1

Cation Qualitative Analysis (p. 452)
In this investigation you will be given a solution containing a number of ions. You will plan and carry out a qualitative analysis, involving a series of precipitation tests, to identify the ions.

We can determine the presence of some metal ions in solution by the colour of the precipitates they form (**Figure 3**). Lead(II) ions, $Pb^{2+}(aq)$, for example, produce a characteristic bright yellow precipitate with iodide ions, $I^-(aq)$:

$$Pb^{2+}(aq) + 2\,I^-(aq) \rightarrow \underset{\text{yellow}}{PbI_2(s)}$$

Figure 3 We can add a solution containing iodide ions to precipitate any lead(II) ions from an unknown solution.

Qualitative Analysis Involving Differences in Solubility

The identification of lead cations in a solution is based on the low solubility of lead(II) iodide. Chemists have developed a series of qualitative tests to identify other cations as well, based on the precipitation of compounds with low solubility. These tests often must be conducted in a specific order to successfully identify the ions present. We can use the following example to illustrate how these tests apply to identifying cations in a mixture.

Suppose that you are given an unknown solution and suspect that it may contain the cations $Pb^{2+}(aq)$, $Ba^{2+}(aq)$, or both. What chemical tests could you perform to identify each ion in solution? Based on the solubility table, you know that lead is the only one of the two cations that forms a precipitate with iodide. Therefore, adding a solution of a highly soluble compound like potassium iodide, KI(aq), to the mixture will produce a yellow precipitate of lead(II) iodide if lead cations are present.

Confirming that barium ions are present, in a solution containing both barium and iodide ions, is more difficult. According to the solubility table, the addition of a solution containing sulfate ions to a solution containing barium ions should produce a precipitate of barium sulfate. But if lead ions are also present, a precipitate of lead(II) sulfate will also form. Therefore, all the lead ions must first be removed before the presence of barium can be confirmed using sulfate ions. This can be done by adding an excess of potassium iodide to the unknown solution (Figure 3). An excess is necessary to ensure that all of the lead ions are precipitated. The lead(II) iodide precipitate is then filtered out. Any barium ions present, along with spectator ions from the precipitation reaction, remain in the filtrate. The **filtrate** is the clear liquid remaining after a mixture has been filtered. The presence of barium ions in the filtrate is confirmed if the addition of a sodium sulfate solution produces a precipitate.

filtrate the clear liquid (solvent and any dissolved substances) collected after a mixture is filtered to remove any solid components

In summary, the qualitative analysis of this unknown solution involves two tests. To correctly identify the ions present, the tests have to be performed in a specific sequence. They could not be done in reverse order since both lead and barium ions precipitate with sulfate. The flow chart in **Figure 4** summarizes the qualitative analysis.

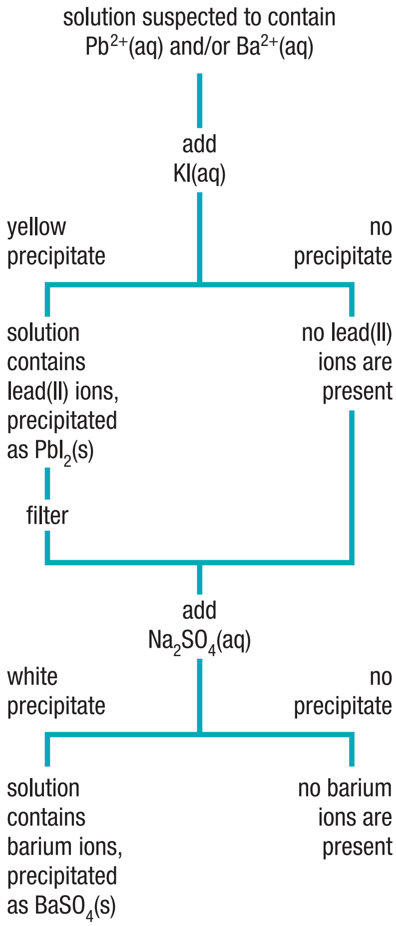


Figure 4 The flow chart for the qualitative analysis of a solution that may contain $Pb^{2+}(aq)$ and/or $Ba^{2+}(aq)$

Tutorial 1 Planning a Qualitative Analysis

Follow these tips for conducting a qualitative analysis of cations in an unknown solution. These tips also apply for conducting a qualitative analysis for anions. Simply switch "cation" for "anion" (and vice versa) in each tip.

1. Use a solubility table to identify an anion that may be used to precipitate each of the cations in the solution.
2. Determine the order in which the precipitation reactions must occur, and select an appropriate solution for each test.
3. Construct a flow chart like the one in Figure 4 to illustrate the test sequence.

Sample Problem 1: Developing a Flow Chart for a Qualitative Analysis

Construct a flow chart for the qualitative analysis of a solution that may contain silver ions, $Ag^+(aq)$, and/or zinc ions, $Zn^{2+}(aq)$.

Step 1. Use a solubility table to identify an anion that may be used to precipitate each of the cations in the solution.

According to the solubility table, sulfide ions precipitate both silver and zinc. However, chloride will precipitate silver but not zinc.

Step 2. Determine the order in which the precipitation reactions must occur, and select an appropriate solution for each test.

Silver ions must be precipitated first by adding a solution of a soluble chloride compound such as sodium chloride, NaCl(aq).

If the test is positive, filter the mixture to remove the silver chloride precipitate.

Adding a solution of sodium sulfide should produce a precipitate of zinc sulfide, if zinc ions are present.

Step 3. Construct a flow chart to illustrate the test sequence.

Your flow chart should be similar to **Figure 5**.

Practice

1. Construct a flow chart for the qualitative analysis of a solution that may contain each of the following pairs of ions:
   (a) Sr^{2+}(aq) and/or Mn^{2+}(aq)
   (b) Cu^{+}(aq) and/or Fe^{3+}(aq)
   (c) OH^{-}(aq) and/or $C_2H_3O_2^{-}$(aq)

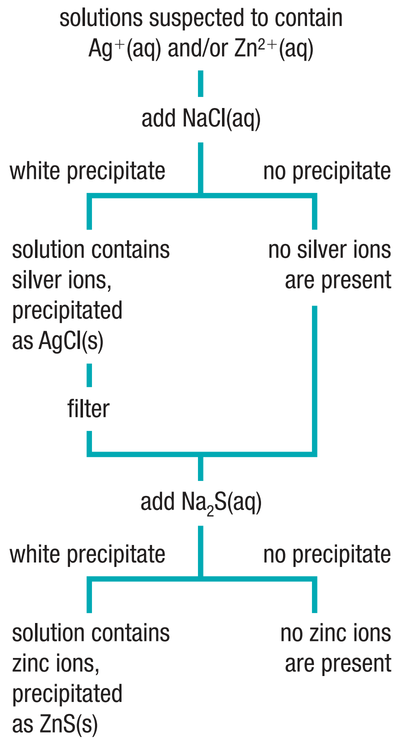


Figure 5 Flow chart for Sample Problem 1

Qualitative Analysis Involving Flame Colour

A limitation of qualitative analysis using precipitation reactions is that some cations, like those of the alkali metals, are highly soluble. They tend not to form precipitates in water. There is an alternative. In Unit 1, you learned that many cations produce a characteristic colour when they are placed in a flame. This type of analysis is called a **flame test**. Flame tests can be done in many ways. In an activity from Unit 1, you dipped a loop of nichrome wire into an ionic compound and then placed the loop into a Bunsen burner flame. The flame colour observed was caused by a cation in the compound. For example, copper(II) sulfate, copper(II) nitrate, and even copper metal all produce the same flame colour: green. Therefore, you can conclude that the element copper is responsible for the flame colour. **Figure 6** shows the characteristic flame colours of selected elements. Flame colours of more elements are listed in **Table 1** (on the next page). Identifying cations using flame tests is most successful when only one cation is present. In mixtures, the flame colour of one cation may mask another, making identification more difficult.

flame test a diagnostic test used to identify a specific element

WEB LINK

There are many videos of flame tests online. To link to some of them,

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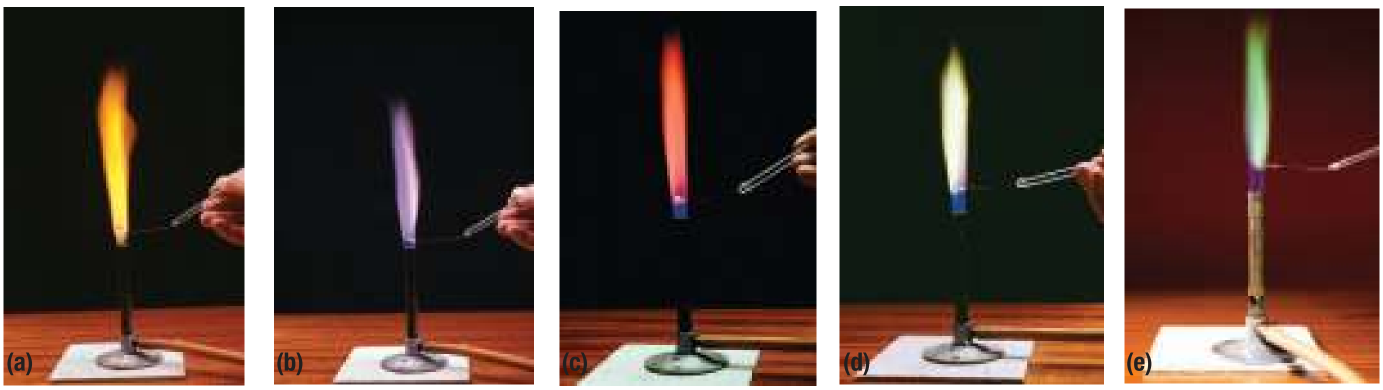

Figure 6 The flame tests for (a) sodium, (b) potassium, (c) strontium, (d) barium, and (e) copper

Spectroscopy

Successfully identifying a cation using a flame test requires the hot, nearly colourless flame of a Bunsen burner. This device was developed in the mid-1850s by the German chemist Robert Bunsen. Prior to the 1850s, gas burners used in laboratories produced sooty, relatively cool flames. Frustrated with these inefficient burners,

Table 1 Colours of Flames

Ion	Flame colour
H^+	colourless
Li^+	bright red
Na^+	yellow (Figure 5(a))
K^+	violet (Figure 5(b))
Ca^{2+}	yellow-red
Sr^{2+}	bright red (Figure 5(c))
Ba^{2+}	yellow-green (Figure 5(d))
Cu^{2+}	green (Figure 5(e))
Pb^{2+}	light blue-grey
Zn^{2+}	whitish green

CAREER LINK

Astronomers are interested in the chemistry of distant objects. One way to find out about their chemical composition is to analyze the spectrum of light and other radiation that they emit. To find out more about the work of an astronomer,

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Bunsen began to tinker with their design. Bunsen found that he could produce a hotter flame if air and fuel were allowed to mix first, before burning. This is why you twist open the barrel on a modern burner to produce a blue flame.

The hot and nearly colourless flame of the Bunsen burner was an important technological advance. It enabled chemists to investigate the light given off during a flame test. In 1858, Robert Bunsen and Gustav Kirchoff developed a device called a spectroscope, which could analyze the light emitted during a flame test. Later, chemists learned to extend this technology to analyze light beyond the visible portion of the electromagnetic spectrum. This ultimately led to a new branch of chemical analysis called spectroscopy. Thus technological developments allowed scientific knowledge to move forward.

Modern spectroscopy has numerous applications in both quantitative and qualitative analysis. Environmental chemists use spectroscopy to measure concentrations of environmental pollutants as low as parts per billion or even, in some cases, parts per trillion. Astronomers use spectroscopy to determine the composition of distant objects, including whether or not they contain water (**Figure 7**).

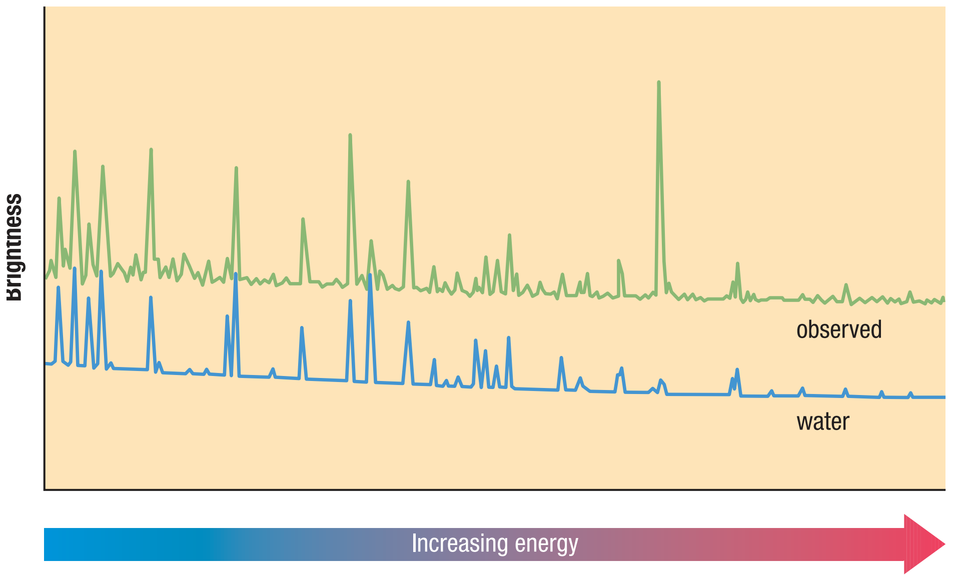

Figure 7 The green line indicates light emitted by a distant cluster of stars, about 1000 light years from Earth. These data were collected by NASA's Spitzer space telescope. The blue line indicates light emitted by very hot water. The alignment of the peaks confirms the presence of water near the stars.

9.3 Summary

- The presence of certain substances in our environment can be a health risk. Identifying these substances, and their concentrations, is the first step to minimizing our exposure to them.
- Qualitative analysis is the identification of a substance.
- Quantitative analysis is the process of determining the quantity of the substance.
- The presence of some cations or anions in a solution can be determined using differences in the solubility of the compounds they form.
- Some cations can be identified by their characteristic flame colour.
- Spectroscopy involves the analysis of the colours of emitted light.

9.3 Questions

1. Classify each of the following observations as qualitative or quantitative analysis: T/I
   (a) Testing an unknown gas with a burning splint
   (b) Measuring the density of a pure metal
   (c) Conducting a flame test to identify cations
   (d) Determining the concentration of carbon dioxide in the atmosphere
   (e) Testing the pH of rainwater and finding that it is 5.3
2. During a qualitative analysis of cations, more than enough sodium hydroxide solution is added to precipitate all of the copper ions, $Cu^{2+}(aq)$. Identify the limiting reagent and excess reagent. Justify your choice. T/I
3. Qualitative analysis of solutions often involves precipitation reactions. Why is it necessary to perform the steps in the correct order? K/U
4. Give a specific example of how advances in technology are necessary before advances in science can occur. K/U A
5. Refer to Table 1 to predict the flame colour when samples of each of the following compounds are placed in a Bunsen burner flame:
   (a) $BaCl_2$ (b) $ZnCO_3$ (c) $CuCl_2$ (d) NaCl
6. Some of the colours produced in a fireworks display are produced by adding compounds of selected cations. What cations may be present in **Figure 8**? K/U

Figure 8 Can you identify the elements responsible for the colours in these fireworks?

7. Why do you think it is necessary to use distilled water rather than tap water when preparing solutions for a qualitative analysis involving precipitation reactions? K/U T/I
8. Design a procedure to test and identify the presence of each of the following pairs of ions in solution: K/U T/I
   (a) silver and lead(II)
   (b) calcium and sodium
   (c) copper(I), iron(III), and barium
   (d) sulfate and chloride
9. Three unlabelled bottles containing white solid compounds were found in the "chlorides" section of the chemical storeroom. The compounds in these bottles were most likely sodium chloride, NaCl(s), magnesium chloride, $MgCl_2(s)$, and potassium chloride, KCl(s). What qualitative tests could be conducted to confirm the identity of these three compounds? T/I
10. Examine the flowchart in Figure 5. Would the analysis be effective if the two steps were reversed? Explain your answer. K/I T/I
11. A sample of industrial waste water is thought to contain $Fe^{3+}(aq)$, $Ba^{2+}(aq)$, and $Ag^{+}(aq)$ ions. K/I T/I C
    (a) Select a compound that could be used to precipitate all three ions from solution.
    (b) Plan a procedure that will precipitate these ions, one at a time, from the solution. Write out your procedure as a flow chart, similar to Figure 4.
12. A spectrophotometer is a device used to determine the concentration of coloured solutions (**Figure 9**). Research how this device works and list some of its applications. A

Figure 9 A spectrophotometer

13. (a) Research how a Breathalyzer determines blood alcohol content (BAC).
    (b) Inexpensive pocket-sized Breathalyzers have recently become available in Canada. In your opinion, what are the advantages and disadvantages of the availability of these devices? K/U T/I A

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Drugs in Drinking Water

ABSTRACT

Extremely low concentrations of prescription medications have recently been discovered in the drinking water of communities worldwide. It is possible that these drugs have been in the water for quite some time. They have only been detected recently, however, because we have better technologies to analyze water. Further improvements in this technology will likely uncover other contaminants which, at present, remain undetected. For policy makers who oversee how drinking water is prepared, these disturbing discoveries raise some important questions:

- What, if anything, can be done about these contaminants?
- At what concentration do medications in drinking water pose a health threat?
- What are the long-term effects of exposure to these contaminants?

Introduction

Pain relievers, antidepressants, tranquilizers, and even birth control hormones are just a few of the classes of medications that scientists have recently detected in the drinking water of many communities worldwide. The source of these medications is not irresponsible dumping by drug manufacturers. Rather, most of these medications are in the wastewater being flushed from individual homes. Some of these drugs are removed by standard water treatment processes like chlorination and the use of activated carbon filtration. But a surprising number are not (**Figure 1**).

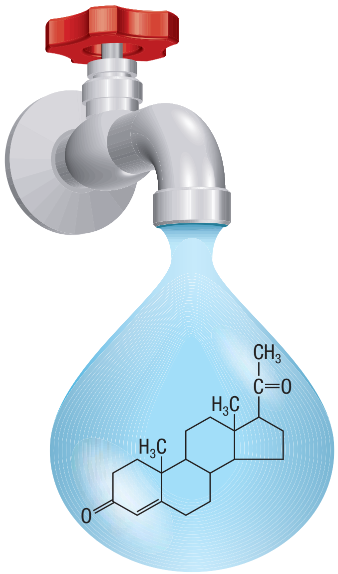

Figure 1 Should we be worried about the pharmaceuticals coming out of our taps?

The thought of consuming someone else's medication can be disturbing. But the good news is that the concentrations of these medications are extremely low, measured mostly in parts per trillion. (One part per trillion is equivalent to one drop of water in 20 Olympic-sized swimming pools, or 3 seconds in 100 000 years.) Medication concentrations this low are at least tens of thousands of times more dilute than prescription doses.

So far there has been no published evidence to suggest that these concentrations of medications in drinking water have any effect on human health. As a result, many water quality experts remain cautiously optimistic that drinking water remains safe. However, they also point out that this is a recent discovery and that further research is necessary, particularly regarding any long-term effects that may result.

How Do Drugs Get into the Water?

Some of the medications in drinking water result from people flushing unused medicines directly down the drain. Others leach from discarded medications in landfill sites. The majority, however, are excreted medications. Ingested medications are not always completely used up by the body. What remains is filtered out by the kidneys, excreted in urine, and flushed down the toilet. As a result, the wastewater from a home is a snapshot of the household's medication use. Wastewater can be an indicator of illicit drug use as well. Studies have shown that the analysis of wastewater can be useful to track patterns in the use of illicit drugs like cocaine and ecstasy in a neighbourhood.

Why Now?

Prescription medications have been used for decades. The first reported detection of a medication in drinking water was in the 1970s, but the majority of medications in drinking water have only been discovered in the past few years. This does not necessarily mean that more medications are getting into the water system. Instead, it is a reflection of how sensitive the detection tools have become.

Around the middle of the twentieth century, the lowest detectable concentrations of contaminants in drinking water were measured in parts per million. Advances in water-testing technology over the next few decades made the detection of parts per billion possible. Today, experts can routinely detect contaminant concentrations in parts per trillion (**Figure 2**). It may only be a matter of time before experts will be able to detect concentrations in parts per quadrillion.

Detecting contaminants is one thing. Knowing what to do with this information is quite another.

Figure 2 High-tech equipment is required when testing for very low concentrations of contaminants.

What Should We Do?

Conventional water treatment processes remove most of the contaminants from drinking water. Activated carbon filters, for example, remove some medications and organic contaminants. Some medications are destroyed when water is chlorinated. But because there is such a broad range of medications in drinking water, no single treatment option could eliminate all of them.

Because the concentrations of medications in drinking water are unimaginably low, many water industry experts warn against taking drastic measures to "ultra-clean" our water. Such measures would certainly be "ultra-expensive" and energy intensive. The energy used to operate current water treatment plants already makes up a significant portion of a municipality's total energy bill. The additional energy and capital investment required to "ultra-clean" water would certainly be a financial burden to both a city and its taxpayers. The cost of additional purification methods means that less money would be available to upgrade the aging water distribution infrastructure of pipes and pumps in our cities. Consuming more energy also results in increased greenhouse gas emissions. Hence, a potential fix for one environmental problem could aggravate another.

Removing medications from drinking water is clearly difficult. However, there are relatively simple strategies that can be implemented to help reduce drug contamination of drinking water in the first place. City councils can create public service campaigns to remind the public not to dispose of unused prescription drugs down the drain. Pharmacies can set up drop boxes to collect unused medications so that they can dispose of them appropriately.

Clearly, we need further research into the potential long-term effects of medications in our water. In the meantime, we can reduce the impact by disposing of our own medications properly.

Further Reading

Boyd, G.R., Reemtsma, H., Grimm, D. A., & Mitra, S. (2003). Pharmaceuticals and personal care products (PPCPs) in surface and treated waters of Louisiana, USA and Ontario, Canada. *The Science of the Total Environment*, 311, 135–149.

Snyder, S., Lue-Hing, C., Cotruvo, J., Drewes, J.E., Eaton, A., Pleus, R.C., & Schlenk. D. (2010). Pharmaceuticals in the Water Environment.

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9.4 Questions

1. Identify two ways in which we unintentionally contaminate our water with medications. K/U
2. Why were medications in drinking water only discovered recently? K/U
3. Why is it not feasible to remove all traces of medication from drinking water? K/U
4. In your opinion, does the long-term exposure to medications in drinking water pose a greater threat to humans or to aquatic life? Why? T/I
5. If you were a researcher in the field of contaminants in water, what questions would you like to try to answer? Why? T/I A
6. Find out whether your local pharmacy collects leftover medications. If it does, are there any restrictions on what can be dropped off there? T/I
7. There are hormones in our water supply that originate from a variety of medications. You may have heard that some of these hormones could be affecting people's fertility and changing the relative numbers of male and female babies being born. Is there any truth to these statements? Research to find out. Summarize your findings for a page on a Chem-myth-try website. Be sure to support your report with credible sources of information. T/I C A

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9.5 Stoichiometry of Solutions

Figure 1 Chemists frequently work with aqueous solutions because they can control many variables by adjusting the temperature and/or concentrations.

Picture a stereotypical chemist at work. There are probably several bottles and jars of mysterious liquids in your mental image. It might be a stereotype, but a great deal of chemistry really does take place in aqueous solutions (**Figure 1**). This is partly because water is cheap and easily accessible. It is also good at dissolving many substances that it comes in contact with. Chemicals mix more completely when they are dissolved, resulting in faster reactions. Allowing reactions to occur in solution also makes it easier to control how quickly the reactions occur. The investigator can adjust the solution temperature or the reactant concentrations.

Being able to predict the quantity of product or the quantity of reactants required is critical to the success of a chemical process. For example, in Section 9.1, you learned that patients consume a "barium meal" prior to having X-rays taken of their gastrointestinal tract. The barium sulfate used can be prepared by combining solutions of barium chloride and sodium sulfate:

$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2\ NaCl(aq)$$

Barium ions, however, are quite toxic. No free barium ions should remain in the final solution. Therefore, when barium sulfate is prepared, the quantities of both reactants have to be carefully measured. Chemists must determine the correct amount of sodium sulfate required using stoichiometry.

Stoichiometry and Reactions

In Chapter 7 you learned how we can use stoichiometry to predict the quantity of one chemical required to react with another. Recall that we can use a balanced chemical equation to determine the ratios of chemicals involved in the reaction. It is likely that we will know the amount concentrations and volumes of one or more of the reactant solutions. We will make a stoichiometric prediction regarding the reaction. First, though, we must convert the known quantities of chemicals used to amounts (**Figure 2**). We can apply this concept to making a barium meal that is safe to consume.

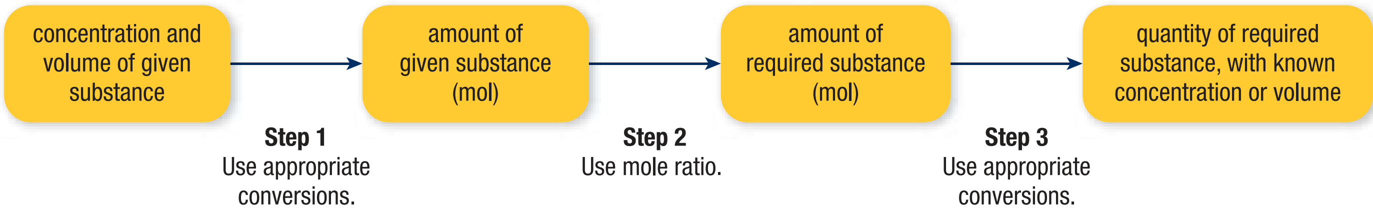


Figure 2 The general strategy for solving stoichiometry problems

Tutorial 1 Solving Stoichiometry Problems Involving Solutions

The strategies you learned to solve stoichiometric problems in Chapter 7 apply to solutions as well. The only difference is that amount concentrations and volumes of solutions are involved.

Sample Problem 1: Determining Volume to Precipitate a Compound

Determine the minimum volume of 0.42 mol/L sodium sulfate, $Na_2SO_4(aq)$, that is required to react completely with all the barium ions in 500.0 mL of a 0.100 mol/L barium chloride, $BaCl_2$, solution.

Given: $c_{Na_2SO_4} = 0.42$ mol/L; $c_{BaCl_2} = 0.100$ mol/L; $V_{BaCl_2} = 500.0$ mL

Required: volume of 0.42 mol/L sodium sulfate, $V_{Na_2SO_4}$

Analysis:

$$c = \frac{n}{V}$$

Solution:

Step 1. Convert all volumes of the solutions to litres, if necessary.

$$V_{BaCl_2} = 500.0\ \cancel{mL} \times \frac{1\ L}{1000\ \cancel{mL}}$$
$$V_{BaCl_2} = 0.5000\ L$$

Step 2. Write a balanced equation for the reaction, listing the given value(s), required value(s), and amount concentrations below the substances being considered in the problem.

$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2\ NaCl(aq)$$
$$V_{BaCl_2} = 0.5000\ L \qquad V_{Na_2SO_4}$$
$$c_{BaCl_2} = 0.100\ mol/L \qquad c_{Na_2SO_4} = 0.42\ mol/L$$

Step 3. Rearrange the concentration equation to solve for the amount of barium chloride, n_{BaCl_2}.

$$n_{BaCl_2} = c_{BaCl_2}V_{BaCl_2}$$
$$= 0.100\ mol/L \times 0.5000\ L$$
$$n_{BaCl_2} = 0.0500\ mol$$

Step 4. Determine the amount of the substance whose volume is required from the amount of the substance whose volume and concentration are given. To do this, multiply the amount calculated in Step 3 by a suitable conversion factor derived from the mole ratio in the balanced equation. In this case, the conversion factor is either

$$\frac{1\ mol_{BaCl_2}}{1\ mol_{Na_2SO_4}} \text{ or } \frac{1\ mol_{NaSO_4}}{1\ mol_{BaCl_2}}$$

Since we are determining the amount of Na_2SO_4 from the amount of $BaCl_2$, we use the conversion factor $\frac{1\ mol_{Na_2SO_4}}{1\ mol_{BaCl_2}}$, as follows:

$$n_{Na_2SO_4} = 0.0500\ \cancel{mol_{BaCl_2}} \times \frac{1\ mol_{Na_2SO_4}}{1\ \cancel{mol_{BaCl_2}}}$$
$$n_{Na_2SO_4} = 0.0500\ mol_{Na_2SO_4}$$

Step 5. Determine the volume of the substance whose volume is required by rearranging the amount concentration equation and substituting the values determined in Steps 2 and 4.

$$c_{Na_2SO_4} = \frac{n_{Na_2SO_4}}{V_{Na_2SO_4}}$$
$$V_{Na_2SO_4} = \frac{n_{Na_2SO_4}}{c_{Na_2SO_4}}$$
$$= \frac{0.0500\ mol}{\frac{0.42\ mol}{1\ L}}$$
$$= 0.0500\ \cancel{mol} \times \frac{1\ L}{0.42\ \cancel{mol}}$$
$$V_{Na_2SO_4} = 0.12\ L \text{ or } 120\ mL$$

Statement: The minimum volume of 0.42 mol/L sodium sulfate required to react with 500.0 mL of 0.100 mol/L of a barium chloride solution is 120 mL.

UNIT TASK **BOOKMARK**

In your analysis of your Unit Task data (p. 498), you will calculate the concentrations of cations in two different solutions. You will then use your findings to compare the solutions.

LEARNING **TIP**

Determining Amount

There is an alternative to using the conversion factor method for finding the amount of barium chloride in Step 3. You could rearrange the concentration equation and solve for *n*.

LIMITING REAGENT PROBLEMS

If the volume and concentration of both of the reactants are given, you will have to determine which of them is the limiting reagent. (Recall that the limiting reagent is the reactant that is completely used up during the reaction.) Sample Problem 2 includes this additional step.

Sample Problem 2: Predicting the Mass of Precitate Expected

Predict the mass of precipitate expected when 1.50 L of 0.800 mol/L sodium carbonate, Na_2CO_3, is mixed with 850 mL of a 1.00 mol/L aluminum nitrate, $Al(NO_3)_3$, solution.

Given: $V_{Na_2CO_3} = 1.50$ L; $c_{Na_2CO_3}$0.800 mol/L; $V_{Al(NO_3)_3} = 850$ mL; $c_{Al(NO_3)_3} = 1.00$ mol/L

Required: mass of aluminum carbonate precipitate, $m_{Al_2(CO_3)_3}$

Solution:

Step 1. Convert all volumes to litres, if necessary.

$$V_{Al(NO_3)_3} = 850\ \cancel{mL} \times \frac{1\ L}{1000\ \cancel{mL}}$$

$$V_{Al(NO_3)_3} = 0.850\ L$$

Step 2. Write a balanced equation for the reaction, listing the given value(s), required value(s), and amount concentrations below the substances being considered in the problem.

$3\ Na_2CO_3(aq)$	$+\ 2\ Al(NO_3)_3(aq) \rightarrow$	$Al_2(CO_3)_3(s) + 6\ NaNO_3(aq)$
$V_{Na_2CO_3} = 1.50\ L$	$V_{Al(NO_3)_3} = 0.850\ L$	$m_{Al_2(CO_3)_3}$
$c_{Na_2CO_3} = 0.800\ mol/L$	$c_{Al(NO_3)_3} = 1.00\ mol/L$	

Step 3. Determine the amounts of both substances given using the given volumes and concentrations as follows:

$$n_{Na_2CO_3} = 1.50\ \cancel{L} \times \frac{0.800\ mol}{1\ \cancel{L}}$$

$$n_{Na_2CO_3} = 1.20\ mol$$

$$n_{Al(NO_3)_3} = 0.850\ \cancel{L} \times \frac{1.00\ mol}{1\ \cancel{L}}$$

$$n_{Al(NO_3)_3} = 0.850\ mol$$

Step 4. Determine which reactant is the limiting reagent. Since the amounts of two reactants are now known, one of the reactants is a limiting reagent. To determine which one is the limiting reagent, determine the amount of one reactant required to react completely with the other reactant as follows:

$$n_{Na_2CO_3} = 0.850\ \cancel{mol_{Al(NO_3)_3}} \times \frac{3\ mol_{Na_2CO_3}}{2\ \cancel{mol_{Al(NO_3)_3}}}$$

$$n_{Na_2CO_3} = 1.28\ mol$$

Therefore, 0.850 mol $Al(NO_3)_3$ requires 1.28 mol of Na_2CO_3. Since only 1.2 mol of Na_2CO_3 is present, sodium carbonate is the limiting reagent.

Step 5. Use the amount of the limiting reagent determined in Step 4 to determine the amount of the substance whose mass is required, $n_{Al_2(CO_3)_3}$.

$$n_{Al_2(CO_3)_3} = 1.2\ \cancel{mol_{Na_2CO_3}} \times \frac{1\ mol_{Al_2(CO_3)_3}}{3\ \cancel{mol_{Na_2CO_3}}}$$

$$n_{Al_2(CO_3)_3} = 0.40\ mol_{Al_2(CO_3)_3}$$

Step 6. Use the amount calculated in Step 5 to determine the mass of the substance whose mass is required, $m_{Al_2(CO_3)_3}$.

$$m_{Al_2(CO_3)_3} = 0.400\ \cancel{mol} \times \frac{233.99\ g}{1\ \cancel{mol}}$$

$$m_{Al_2(CO_3)_3} = 93.6\ g$$

Statement: Combining 1.50 L of 0.80 mol/L sodium carbonate with 0.850 L of a 1.00 mol/L aluminum nitrate produces 93.6 g of an aluminum carbonate precipitate.

Practice

SKILLS HANDBOOK A6

1. Sodium carbonate, Na_2CO_3, can be used to precipitate strontium ions from a solution of strontium chloride, $SrCl_2$:

   $SrCl_2(aq) + Na_2CO_3(aq) \rightarrow SrCO_3(s) + 2\ NaCl(aq)$

   You have 150 mL of a 0.25 mol/L strontium chloride solution. T/I
   (a) What volume of 0.500 mol/L sodium carbonate is required to precipitate all the strontium ions from this solution? [ans: 0.075 L or 75 mL]
   (b) What mass of precipitate is expected? [ans: 5.5 g]
2. A jelly-like precipitate of iron(III) hydroxide, $Fe(OH)_3$, forms when solutions of iron(III) nitrate, $Fe(NO_3)_3$, and potassium hydroxide, KOH, are combined:

   $Fe(NO_3)_3(aq) + 3\ KOH(aq) \rightarrow Fe(OH)_3(s) + 3\ KCl(aq)$

   What mass of precipitate is expected to form when 70.0 mL of 0.80 mol/L potassium hydroxide is added to 40.0 mL of a 0.50 mol/L iron(III) nitrate solution? T/I [ans: 2.0 g]

Stoichiometry of Ions in Solution

The bright glow of a conductivity tester indicates that a solution of calcium chloride is a good conductor of electricity (**Figure 3**). Our explanation is that calcium chloride completely dissociates into its ions as it dissolves (**Figure 4**):

$$\begin{array}{lll} CaCl_2(s) \rightarrow & Ca^{2+}(aq) + & 2\ Cl^-(aq) \\ 1\text{ mol} & 1\text{ mol} & 2\text{ mol} \end{array}$$

This equation is called a dissociation equation. It indicates that for each mole of calcium chloride in the solution, 1 mol of calcium ions and 2 mol of chloride ions dissociate from each other. For example, in a 0.25 mol/L calcium chloride solution, the concentration of calcium ions is 0.25 mol/L and the concentration of chloride ions is 0.50 mol/L.

LEARNING **TIP**

Ion Concentrations
The concentration of the ions is not always the same as the concentration of the compound they came from. In this case, when $CaCl_2$ dissociates, the chloride concentration will always be double the calcium chloride concentration since 9 chloride ions are released per formula unit of $CaCl_2$.

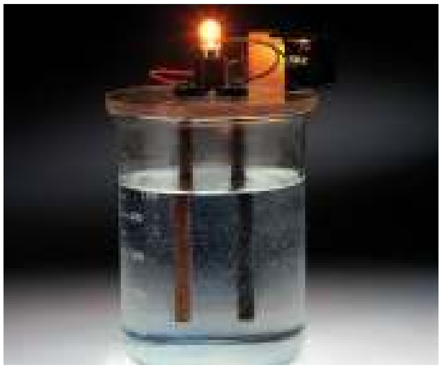


Figure 3 Calcium chloride is a good electrolyte because it dissolves in water to produce a solution that conducts electricity well.

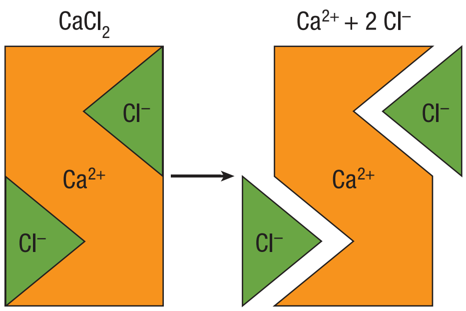


Figure 4 How do the final concentrations of chloride and calcium ions compare to the initial concentration of the compound?

Tutorial 2 Determining the Amount Concentration of Ions in Solution

When calculating the amount concentration of each type of ion released when an ionic compound dissolves, you will need to write a dissociation equation.

Investigation 9.5.1

Percentage Yield of a Precipitation Reaction (p. 453)

In this investigation you will first prepare two standard solutions, one of which contains calcium ions. You will then use the other solution to remove the calcium ions in a precipitation reaction. What is the percentage yield of the reaction?

Sample Problem 1: Calculating the Amount Concentration of Ions

A technician dissolves 17.1 g of aluminum sulfate, $Al_2(SO_4)_3$, to prepare 250.0 mL of solution. What are the amount concentrations of aluminum ions and sulfate ions in the solution?

Given: $m_{Al_2(SO_4)_3} = 17.1$ g; $V_{Al_2(SO_4)_3} = 250.0$ mL

Required: amount concentration of aluminum and sulfate ions, $c_{Al^{3+}}$; $c_{SO_4^{\,2-}}$

Solution:

Step 1. Convert given volumes to litres, if necessary.

$$V_{Al_2(SO_4)_3} = 250.0\ \cancel{mL} \times \frac{1\ L}{1000\ \cancel{mL}}$$

$$V_{Al_2(SO_4)_3} = 0.2500\ L$$

Step 2. Determine the amount of solute from the given mass of solute, if necessary.

$$n_{Al_2(SO_4)_3} = 17.1\ \cancel{g} \times \frac{1\ mol}{342.14\ \cancel{g}}$$

$$n_{Al_2(SO_4)_3} = 0.0500\ mol$$

Step 3. Determine the amount concentration of the solution.

$$c_{Al_2(SO_4)_3} = \frac{n_{Al_2(SO_4)_3}}{V_{Al_2(SO_4)_3}}$$

$$= \frac{0.0500\ mol}{0.250\ L}$$

$$c_{Al_2(SO_4)_3} = 0.200\ mol/L$$

Step 4. Write a dissociation equation listing the calculated amounts and the required value(s) below the substances being considered in the problem.

$$Al_2(SO_4)_3(s) \rightarrow 2\ Al^{3+}(aq) + 3\ SO_4^{\,2-}(aq)$$

$$c_{Al_2(SO_4)_3} = 0.200\ mol/L \qquad c_{Al^{3+}} \qquad c_{SO_4^{\,2-}}$$

Step 5. Convert the concentration of the compound into concentration of the ions.

$$c_{Al^{3+}} = 0.200\ \frac{\cancel{mol_{Al_2(SO_4)_3}}}{L} \times \frac{2\ mol_{Al^{3+}}}{1\ \cancel{mol_{Al_2(SO_4)_3}}}$$

$$c_{Al^{3+}} = 0.400\ mol/L$$

$$c_{SO_4^{\,2-}} = 0.200\ \frac{\cancel{mol_{Al_2(SO_4)_3}}}{L} \times \frac{3\ mol_{SO_4^{\,2-}}}{1\ \cancel{mol_{Al_2(SO_4)_3}}}$$

$$c_{SO_4^{\,2-}} = 0.600\ mol/L$$

Statement: The aluminum and sulfate ion concentrations in 250.0 mL of solution containing 17.1 g of aluminum sulfate are 0.400 mol/L and 0.600 mol/L respectively.

Practice

SKILLS HANDBOOK A6

1. Find the amount concentration of the anion in each of the following solutions: T/I
   (a) 0.50 mol/L barium chloride, $BaCl_2$ [ans: 1.00 mol/L]
   (b) 6.0 mol/L potassium hydroxide, KOH [ans: 6.0 mol/L]
   (c) 0.10 mol/L aluminum chlorate, $Al(ClO_3)_3$ [ans: 0.30 mol/L]
2. Calculate the amount concentration of the ammonium ion in 100.0 mL of a solution containing 14.4 g of ammonium carbonate. T/I [ans: 3.00 mol/L]
3. Calculate the mass of sodium phosphate required to prepare 1.75 L of solution in which the sodium ion concentration is 0.25 mol/L. T/I [ans: 24 g]

9.5 Summary

- Many applications involve determining the required quantities of reactants.
- In solution stoichiometry problems, concentrations and volumes are used to determine the amount of a given chemical using the equation $n = cv$.
- Stoichiometry problems may involve determining the limiting reagent.
- In limiting reagent problems, the limiting reagent determines the amount of product expected.
- For highly soluble ionic compounds, the amount concentration of each ion released can be calculated from a dissociation equation. Ion concentration equals the amount concentration of the compound multiplied by the coefficient of that ion in the dissociation equation.

9.5 Questions

1. Combining solutions of silver nitrate and sodium chloride produces a silver chloride precipitate:

   $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

   A researcher discovered that 32.0 mL of 0.100 mol/L silver nitrate is required to precipitate all the chloride ions in 25 mL of a solution of sodium chloride. T/I
   (a) What is the amount concentration of sodium ions in the initial sodium chloride solution?
   (b) What is the concentration, in g/L, of sodium chloride in the initial sodium chloride solution?
2. Nickel ions can be removed from a nickel(II) sulfate solution, $NiSO_4(aq)$, by precipitating them with a solution of sodium hydroxide, $NaOH(aq)$. T/I C
   (a) Write the balanced chemical equation for this reaction.
   (b) Predict the mass of precipitate expected when 50.0 mL of 0.45 mol/L nickel(II) sulfate solution is combined with 25.0 mL of 1.00 mol/L sodium hydroxide solution.
3. Combining solutions of sodium carbonate and calcium chloride produces a calcium carbonate precipitate:

   $Na_2CO_3(aq) + CaCl_2(aq) \rightarrow CaCO_3(s) + 2\ NaCl(aq)$

   (a) During an investigation, 15.2 g of calcium carbonate was collected using 200.0 mL of sodium carbonate solution and an excess of calcium chloride solution. What was the amount concentration of the original sodium carbonate solution?
   (b) What volume of a 0.500 mol/L calcium chloride solution would produce the same mass of precipitate as in (a)? T/I
4. The steel industry uses large volumes of concentrated hydrochloric acid to remove rust, which is essentially Fe_2O_3, from the surface of steel. This process is called "pickling:"

   $Fe_2O_3(s) + 6\ HCl(aq) \rightarrow 2\ FeCl_3(aq) + 3\ H_2O(l)$

   What volume of 12.0 mol/L HCl(aq) is required to remove 224 g of iron(III) oxide? T/I
5. When aluminum metal is placed in copper(II) sulfate solution, the aluminum ions displace the copper(II) ions in a single displacement reaction. T/I C
   (a) Write the chemical equation for the reaction.
   (b) What mass of aluminum is required to remove all the copper ions from 150 mL of a 0.100 mol/L solution of copper(II) sulfate?
6. Sodium hydroxide, $NaOH(aq)$, is used in the production of paper, textiles, cleaners, and detergents. Sodium hydroxide is produced industrially by passing electricity through a concentrated sodium chloride solution. The chemical equation for this reaction, which involves water, is

   $2\ NaCl(aq) + 2\ H_2O(l) \rightarrow 2\ NaOH(aq) + Cl_2(g) + H_2(g)$

   As much as 45×10^6 t of sodium hydroxide is produced around the world each year. What volume of a 6.0 mol/L sodium chloride solution is required to produce this mass of sodium hydroxide? T/I
7. An unknown solution was prepared by dissolving 0.42 g of sodium chloride, $NaCl(s)$, or calcium chloride, $CaCl_2(s)$, in enough water to make 100.0 mL of solution. 15.0 mL of a 0.50 mol/L silver nitrate solution was required to precipitate all the chloride ions from this solution. The two possible precipitation reactions are

   $NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

   $CaCl_2(aq) + 2\ AgNO_3(aq) \rightarrow 2\ AgCl(s) + Ca(NO_3)_2(aq)$

   Was the unknown solute calcium chloride or sodium chloride? T/I
8. Calculate the amount concentration of the cations in the following solutions: T/I
   (a) 0.5 mol/L sodium carbonate, Na_2CO_3
   (b) 0.2 mol/L ammonium sulfite, $(NH_4)_2SO_3$
   (c) 1.5 mol/L iron(III) sulfate, $Fe_2(SO_4)_3$
9. What mass of sodium carbonate, $Na_2CO_3(s)$, is required to prepare 200.0 mL of a solution in which the sodium concentration is 0.85 mol/L? T/I

Investigation 9.2.1 OBSERVATIONAL STUDY

SKILLS MENU

- Questioning
- Researching
- Hypothesizing
- Predicting
- Planning
- Controlling Variables
- Performing
- Observing
- Analyzing
- Evaluating
- Communicating

How Effective Are Water Filters?

Drinking water is an aqueous solution of many different substances. The characteristics of water can be influenced by the substances dissolved in it. For example, excess calcium ions make water "hard."

Today, most Canadians have several options for their drinking water. Most choose tap water. Many choose to further treat their water by passing it through a cartridge-type filter (**Figure 1**).

In this investigation, you will first observe the positive chemical test for a specific set of ions. You will then test for these ions in filtered and unfiltered water.

Figure 1 The beads inside home water filter cartridges are designed to remove some dissolved substances as water passes through them.

Purpose

To test the effectiveness of the contents of a home water filter cartridge at removing dissolved ions

Equipment and Materials

SKILLS HANDBOOK A1

- chemical safety goggles
- lab apron
- 24-well plate
- 100 mm test tube
- test-tube rack
- labelled waste beaker
- scoopula
- graduated cylinder
- disposable plastic pipette with the top cut off
- dropper bottles containing solutions of:
  - sodium chloride, $NaCl(aq)$ (0.1 mol/L)
  - sodium sulfate, $Na_2SO_4(aq)$ (0.1 mol/L)
  - calcium chloride, $CaCl_2(aq)$ (0.1 mol/L)
  - iron(III) chloride, $FeCl_3(aq)$ (0.1 mol/L)
  - silver nitrate, $AgNO_3(aq)$ (0.1 mol/L)
  - barium chloride, $BaCl_2(aq)$ (0.1 mol/L)
  - sodium phosphate, $Na_3PO_4(aq)$ (0.1 mol/L)
  - potassium thiocyanate, $KSCN(aq)$ (0.1 mol/L)
- resin from a home water filtration product

Silver nitrate stains the skin. Avoid skin and eye contact. If you spill solutions containing silver nitrate on your skin, wash the affected area with plenty of cool water and inform your teacher.

Barium compounds are toxic. Avoid skin and eye contact. If you spill solutions containing barium on your skin, wash the affected area with plenty of cool water.

Most of the solutions used in this investigation are irritants. Avoid skin and eye contact. If you spill any of these substances on your skin, wash the affected area with plenty of cool water and inform your teacher.

Procedure

SKILLS HANDBOOK A6

Part A: Chemical Tests for Ions

1. Put on your chemical safety goggles and lab apron.
2. Combine two drops of sodium chloride and two drops of silver nitrate in one well of the well plate (**Figure 2**). Record your observations in a table. This is the positive test for chloride ions.

Figure 2 Adding solutions to a well plate

3. Repeat Step 2 for the other ion combinations in **Table 1**.

Table 1 Chemical Tests for Selected Ions

Chemical test	Solution A	Solution B
chloride, Cl^-	sodium chloride, $NaCl(aq)$	silver nitrate, $AgNO_3(aq)$
sulfate, SO_4^{2-}	sodium sulfate, $Na_2SO_4(aq)$	barium chloride, $BaCl_2$
calcium, Ca^{2+}	calcium chloride, $CaCl_2(aq)$	sodium phosphate, $Na_3PO_4(aq)$
iron, Fe^{3+}	iron(III) chloride	potassium thiocyanate, KSCN(aq)

4. Discard the contents of the well plate as directed by your teacher.

Part B: Water Treatment

5. Fill the pipette to within 2 cm of the top with resin that your teacher has extracted from the water filter (**Figure 3**).

Figure 3 Resin in a plastic pipette

6. Place the pipette in a test tube.
7. Slowly pour about 5 mL of tap water or "simulated hard water" through the resin in the pipette (**Figure 4**). Collect the treated water in the test tube.
8. Test unfiltered water and the filtered water for the presence of chloride, sulfate, calcium, and iron(III) ions using the same procedure as in Part A.

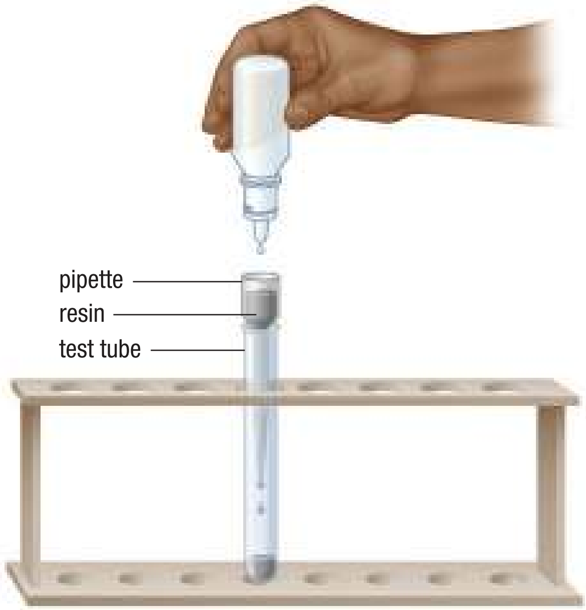


Figure 4 Adding water to the resin

9. Dispose of the waste materials as directed by your teacher.

Analyze and Evaluate

(a) Write the net ionic equations for the first 3 reactions that occurred in Part A. K/U C

(b) How effective was the resin at removing ions from water? T/I

(c) The chemical tests conducted in this investigation cannot absolutely confirm the presence of the test ion in water. Why? T/I

(d) How could you alter the design of the water treatment process in Part B to make it more effective? T/I

Apply and Extend

(e) The resin used in Part B is an example of a cationic exchange resin. Research how these resins remove ions and how they can be recharged once they no longer function. T/I

(f) In your opinion, should every household in Canada use a countertop water filtration system, such as the one you just tested? Explain your position. T/I A

(g) Test different water samples using a commercial water-testing kit. K/U T/I

(h) What happens to the resin in water filters after we use it to filter ions out of our drinking water? How is it disposed of? Are there any environmental implications? Research this topic. T/I

GO TO NELSON SCIENCE

Investigation 9.3.1 CONTROLLED EXPERIMENT

SKILLS MENU

• Questioning	• Planning	• Observing
• Researching	• Controlling Variables	• Analyzing
• Hypothesizing	• Performing	• Evaluating
• Predicting		• Communicating

Cation Qualitative Analysis

While exploring the chemical storeroom, a new chemistry teacher discovered a bottle labelled "unknown cation mixture for qualitative analysis." Rather than disposing of the mixture, the teacher decided to determine the cations it contained so that she could use the solution for a class experiment. Based on the chemical bottles stored nearby, she suspected that the unknown solution may be a mixture of barium nitrate, silver nitrate, and/or zinc nitrate.

Testable Question

Which of the following ions are present in the unknown cation solution: $Ba^{2+}(aq)$, $Ag^{+}(aq)$, and/or $Zn^{2+}(aq)$?

Hypothesis

SKILLS HANDBOOK A2.3

Write a hypothesis in the form of a flow chart describing the qualitative analysis of the solution. Each step should include an "If … then … because …" statement. The available test solutions are listed in the Equipment and Materials list.

Variables

The manipulated variable will be the solution(s) that you add to the unknown cation solution. The responding variable will be the presence or absence of a precipitate. Consider which variables you will control throughout the experiment.

Experimental Design

Plan a sequence of steps, based on your flow chart, that will be used to determine whether the unknown solution contains $Ba^{2+}(aq)$, $Ag^{+}(aq)$, and/or $Zn^{2+}(aq)$. Plan to precipitate and collect one cation at a time from the mixture.

Equipment and Materials

SKILLS HANDBOOK A1

- chemical safety goggles
- lab apron
- funnel
- retort stand with ring clamp **(Figure 1)**
- small beaker
- test-tube rack
- three 100 mm test tubes with stoppers
- labelled waste beaker
- filter paper
- dropper bottles containing:
  - unknown solution
  - sodium carbonate solution, $Na_2CO_3(aq)$ (0.5 mol/L)
  - sodium chloride solution, $NaCl(aq)$ (0.5 mol/L)
  - sodium sulfate solution, $Na_2SO_4(aq)$ (0.5 mol/L)

The unknown solution may be corrosive and/or toxic, and it may stain the skin. Avoid skin and eye contact. If you spill the solution on your skin, wash the affected area with plenty of cool water and inform your teacher.

Sodium carbonate is an irritant. Avoid skin and eye contact. If you spill sodium carbonate on your skin, wash the affected area with plenty of cool water.

Figure 1 Filtration apparatus

Procedure

1. Once you have your teacher's approval for your flow chart and Experimental Design, put on your safety goggles and lab apron, and proceed with your tests.
2. Dispose of all waste as directed by your teacher. Wash your hands thoroughly at the end of the investigation.

Analyze and Evaluate

(a) Explain the sequence of steps you chose for your analysis. K/U

(b) Identify the cations in the mixture. What evidence supports this conclusion? K/U T/I

(c) Write the net ionic equation for each precipitation reaction that occurred. C

(d) Why are sodium compounds used to precipitate the cations rather than compounds of other metals? K/U T/I

(e) Why do you think nitrate compounds were used to prepare the unknown solution? K/U T/I

Apply and Extend

(f) Plan an investigation to use flame tests on each of the precipitates to confirm the identity of the cations. Include an Equipment and Materials list and all safety precautions in your plan. Proceed once you have your teacher's approval. K/U T/I

Investigation 9.5.1 OBSERVATIONAL STUDY

SKILLS MENU

• Questioning	• Planning	• Observing
• Researching	• Controlling Variables	• Analyzing
• Hypothesizing	• Performing	• Evaluating
• Predicting		• Communicating

Percentage Yield of a Precipitation Reaction

Some municipalities soften water during the treatment process by adding a highly soluble carbonate compound and then separating the precipitate by filtration.

Ideally, the precipitation reaction should be very efficient at removing the hardness ions. In other words, the reaction should have a high percentage yield.

In this investigation, you will design and conduct an procedure to discover whether potassium carbonate can efficiently remove hardness ions from water. Rather than using untreated water, you will conduct your investigation with standard solutions of calcium chloride (representing hard water) and potassium carbonate.

The chemical reaction involved is

$$CaCl_2(aq) + K_2CO_3(aq) \rightarrow CaCO_3(s) + 2\ KCl(aq)$$

You will add an excess of potassium carbonate solution to a known volume of calcium chloride solution. You will then collect the calcium carbonate precipitate by filtration, dry it, and determine its mass.

Purpose

To calculate the percentage yield of calcium carbonate

Equipment and Materials

SKILLS HANDBOOK A1

- chemical safety goggles
- lab apron
- 250 mL beaker
- two 100 mL volumetric flasks and stoppers
- balance
- scoopula
- funnel
- dropper
- wash bottle of distilled water
- retort stand and ring clamp
- filter paper
- calcium chloride, $CaCl_2(s)$
- potassium carbonate, $K_2CO_3(s)$

Calcium chloride and potassium carbonate are both irritants. Avoid skin and eye contact. If you spill either of these compounds on your skin, wash the affected area with plenty of cool water.

Procedure

SKILLS HANDBOOK A2.4, A3

Part A: Solution Preparation (Optional)

1. Plan a procedure to prepare 100 mL of a 0.40 mol/L calcium chloride solution and 100 mL of a 0.60 mol/L potassium carbonate solution.
2. Put on your safety goggles and lab apron.
3. With your teacher's approval, proceed with your plan.

Part B: Testing the Efficiency of the Precipitation Reaction

4. Plan a procedure to determine the percentage yield of the precipitation reaction.
5. Include a test for the completeness of the reaction in your plan. This can be done by adding 2 to 3 drops of the potassium carbonate solution to the clear solution above the precipitate (**Figure 1**).

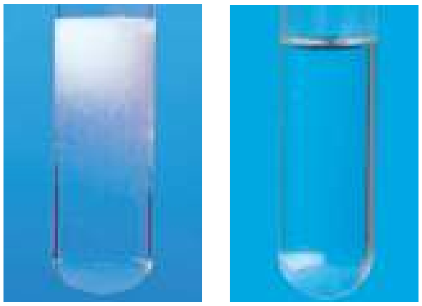


Figure 1 Two results are possible when drops of the potassium carbonate solution are added to test for the completeness of the reaction. Which observation indicates that the reaction is complete? How will you adjust your procedure if you observe the other result?

6. With your teacher's approval, conduct the test and collect your observations.

Analyze and Evaluate

SKILLS HANDBOOK A6

(a) Why was an excess of potassium carbonate used? T/I

(b) Calculate the percentage yield of calcium carbonate. T/I

(c) Evaluate your procedure. What improvements do you recommend if the investigation is repeated? Why? T/I

Apply and Extend

(d) If the distilled water used to prepare the solutions is slightly acidic, carbonate ions may decompose to produce bubbles of carbon dioxide. How might this affect the percentage yield of this investigation. Why? T/I

(e) Solid potassium carbonate decomposes over time. What effect might using old potassium carbonate have on the percentage yield of this investigation. Why? T/I

(f) Suggest another suitable carbonate compound to test. Check the MSDS of this compound for any safety precautions. T/I A

(g) With your teacher's approval, repeat the investigation using the compound that you recommended in (f). Compare the percentage yield of the two reactions. T/I

Summary Questions

1. Create a study guide based on the points listed in the margin on page 422. For each point, create three or four sub-points that provide further information, relevant examples, explanatory diagrams, or general equations.
2. Look back at the Starting Points questions on page 422. Answer these questions using what you have learned in this chapter. Compare your latest answers with those that you wrote at the beginning of the chapter. Note how your answers have changed.

Vocabulary

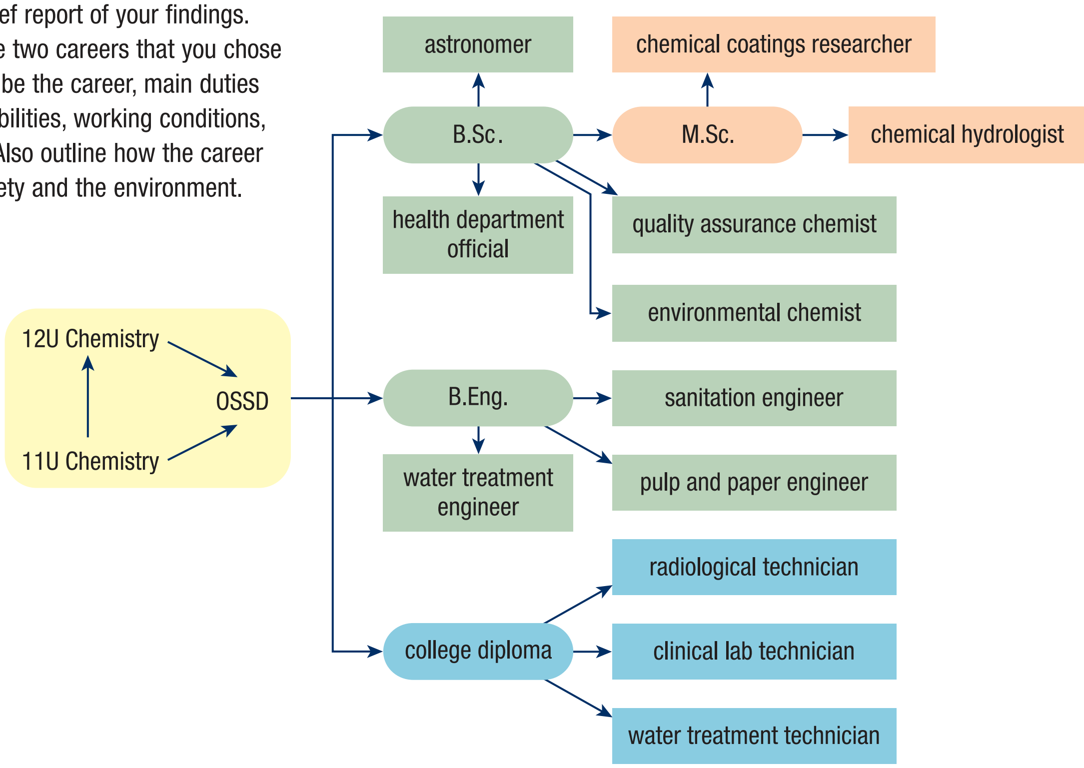
formula equation (p. 425)
total ionic equation (p. 425)
spectator ions (p. 426)
net ionic equation (p. 426)
leachate (p. 431)
hard water (p. 434)
soft water (p. 434)
filtrate (p. 438)
flame test (p. 439)

CAREER PATHWAYS

SKILLS HANDBOOK A7

Grade 11 Chemistry can lead to a wide range of careers. Some require a college diploma or a B.Sc. degree. Others require specialized or post-graduate degrees. This graphic organizer shows a few pathways to careers mentioned in this chapter.

1. Select two careers, related to solutions and their reactions, that you find interesting. Research the educational pathways that you would need to follow to pursue these careers. What is involved in the required educational programs? Prepare a brief report of your findings.
2. For one of the two careers that you chose above, describe the career, main duties and responsibilities, working conditions, and setting. Also outline how the career benefits society and the environment.



GO TO NELSON SCIENCE

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

For each question, select the best answer from the four alternatives.

1. Which of the following chemical equations shows the total ionic equation for a reaction between zinc metal and sulfuric acid? (9.1) K/U
   (a) $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$
   (b) $Zn(s) + 2\ H^+(aq) + SO_4^{2-}(aq) \rightarrow Zn^{2+}(aq) + SO_4^{2-}(aq) + H_2(g)$
   (c) $Zn(s) + 2\ H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
   (d) $Zn(s) + 2\ H^+(aq) + SO_4^{2-}\ (aq) \rightarrow ZnSO_4(aq) + H_2(g)$
2. The following chemical equation is an example of what type of equation?
   $NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
   (9.1) K/U
   (a) total ionic equation
   (b) skeleton equation
   (c) net ionic equation
   (d) formula equation
3. Use Table 1 in Section 9.1 to identify which of the following ions is always a spectator ion in solution: (9.1) K/U
   (a) Cu^+
   (b) SO_4^{2-}
   (c) K^+
   (d) Cl^-
4. Which of the following contaminants is an example of a chemical contaminant in water? (9.2) K/U
   (a) virus
   (b) twig
   (c) pesticide
   (d) protozoa
5. When treating water, what type of material causes small particles in the water to clump together? (9.2) K/U
   (a) floc
   (b) coagulant
   (c) disinfectant
   (d) fluoride
6. Suppose that you have a mixture of four different cations that you want to separate. What is the first step in making a plan to separate the ions? (9.3) K/U
   (a) Determine the order in which the precipitation reactions must occur.
   (b) Draw a flow chart that shows the order in which you will add solutions.
   (c) Start adding solutions to the original solution.
   (d) Determine the anion, if any, that can be used to precipitate each cation.
7. A flame test produces a violet flame. What cation is present in the solution? (9.3) K/U
   (a) K^+
   (b) Ca^{2+}
   (c) Sr^{2+}
   (d) Ba^{2+}
8. Most of the drugs that are currently found in drinking water come from what source? (9.4) K/U
   (a) flushing unused medicines down the drain
   (b) drugs unused by the body and contained in urine
   (c) irresponsible dumping by drug manufacturers
   (d) leaching of medicines placed in landfills
9. What type of chemical equation is the following?
   $Na_2SO_4(aq) \rightarrow 2\ Na^+(aq) + SO_4^{2-}(aq)$
   1 mol 2 mol 1 mol (9.5) K/U
   (a) formula equation
   (b) net ionic equation
   (c) dissociation equation
   (d) total ionic equation

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

10. According to the solubility table, ammonium hydroxide, NH_4OH, is slightly soluble. (9.1) K/U
11. Fluorine is added to kill micro-organisms during municipal water treatment. (9.2) K/U
12. In water treatment, ammonia stabilizes fluoride so that it remains stable in water for a longer period of time. (9.2) K/U
13. Spectroscopy is limited to quantitative analysis. (9.3) K/U
14. Dissolved ions can be present in a filtrate. (9.3) K/U
15. It is likely that the concentration of drugs found in drinking water is hazardous to human health. (9.4) K/U
16. The most precise measurements of the concentration of medications found in water sources are in units of ppm. (9.4) K/U
17. Before making a stoichiometric prediction, you must convert known quantities of chemicals to amounts. (9.5) K/U
18. The amount of product from a reaction depends on the amount of limiting reagent present. (9.5) K/U

To do an online self-quiz,
GO TO NELSON SCIENCE

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

Knowledge

For each question, select the best answer from the four alternatives.

1. Use Table 1 in Section 9.1 to identify which of the following ions is always a spectator ion in solution: (9.1) K/U
   (a) $C_2H_3O_2^-$
   (b) NO_3^-
   (c) Ag^+
   (d) S_2^-
2. Which of the following is a complete list of the spectator ions in the reaction represented by the chemical equation below?
   $Mg(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$ (9.1) K/U
   (a) H^+, Cl^-
   (b) Cl^-
   (c) Mg^{2+}, H^+
   (d) Mg^{2+}, H^+, Cl^-
3. Two solutions, each containing a dissolved ionic compound, are mixed together. Use the solubility table (Table 1 in Section 9.1) to help you decide which of the following pairs of solutions would result in the formation of a precipitate: (9.1) K/U
   (a) sodium chloride, $NaCl(aq)$, and potassium nitrate, $KNO_3(aq)$
   (b) ammonium iodide, $NH_4I(aq)$, and iron(II) sulfate, $FeSO_4(aq)$
   (c) potassium chloride, $KCl(aq)$, and silver nitrate, $AgNO_3(aq)$
   (d) lithium sulfide, $Li_2S(aq)$, and sodium hydroxide, $NaOH(aq)$
4. Which of the following items is an example of *chemical* contamination if it is present in drinking water? (9.2) K/U
   (a) viruses
   (b) pesticides
   (c) suspended clay
   (d) protozoa
5. The presence of bacteria in a water supply is an example of what type of contamination? (9.2) K/U
   (a) biological
   (b) physical
   (c) particulate
   (d) chemical
6. During water treatment, water is softened to remove what materials? (9.2) K/U
   (a) calcium and magnesium ions
   (b) fluorides
   (c) chlorides
   (d) debris and other large particles
7. Which of the following is an example of quantitative analysis? (9.3) K/U
   (a) Barium sulfate appears as a white precipitate.
   (b) A sample of potassium iodide that contains 39.1 g of potassium also contains 127 g of iodide.
   (c) A precipitate of silver chloride forms when solutions of silver nitrate and sodium chloride mix.
   (d) Copper deposits form on an iron nail placed in a copper(II) nitrate solution.
8. What is the most likely reason for the increased number of drugs detected in drinking water, compared to years ago? (9.4) K/U
   (a) Detection methods are better now.
   (b) More drugs are present in the water.
   (c) There is more illegal drug use now than there was years ago.
   (d) Water treatment removes fewer drugs.
9. Suppose 29 g of sodium hydroxide, $NaOH(s)$, is combined with a known quantity of another reactant. The two substances react. What is the first thing that needs to be done to make a stoichiometric prediction about the reaction? (9.5) K/U
   (a) Determine whether the sodium hydroxide will dissociate in solution.
   (b) Measure the volume of sodium hydroxide solution that was used.
   (c) Convert the mass of sodium hydroxide to the amount of NaOH.
   (d) Determine whether or not sodium hydroxide is the limiting reagent.

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

10. Net ionic equations are limited to precipitation reactions. (9.1) K/U
11. Good sanitation has a direct, positive impact on education. (9.2) K/U
12. Quantitative analysis involves the identification of a substance. (9.3) K/U
13. Certain qualitative tests involve identifying cations based on the solubility of certain compounds of the cations. (9.3) K/U
14. None of the drugs found in water supplies are removed by standard water treatment processes. (9.4) K/U
15. Some of the drugs in water sources come from the leaching of discarded medicines in landfills. (9.4) K/U
16. A compound is a good electrolyte if it partially dissociates into ions as it dissolves in water. (9.5) K/U

17. In stoichiometric problems, volume should be in units of millilitres. (9.5) K/U
18. Stoichiometry enables you to predict exactly how much of one chemical will react with a certain amount of another chemical. (9.5) K/U

Match each cation on the left with the colour of its flame on the right.

19. (a) K^+ (i) yellow-red
    (b) Ba^{2+} (ii) bright red
    (c) Li^+ (iii) yellow-green
    (d) Ca^{2+} (iv) green
    (e) Cu^{2+} (v) yellow
    (f) Na^+ (vi) violet (9.3) K/U

Write a short answer to each question.

20. What is a spectator ion? (9.1) K/U
21. Compare and contrast total ionic equations and net ionic equations. (9.1) K/U
22. What two types of chemical reactions can be represented by total ionic equations and net ionic equations? (9.1) K/U
23. Most precipitation reactions are which type of chemical reaction? Explain your answer. (9.1) K/U
24. What type of contaminant is removed first during water treatment processes—physical, biological, or chemical? (9.2) K/U
25. What is sanitation and what is its effect on drinking water? (9.2) K/U
26. Explain the difference between hard water and soft water. (9.2) K/U
27. (a) What is leachate?
    (b) Describe two different ways leachate can pollute nearby water supplies. (9.2) K/U
28. How might a scientist use spectroscopy to determine whether hydrogen gas is present in a distant star? (9.3) K/U
29. Compare and contrast qualitative analysis and quantitative analysis. (9.3) K/U
30. Classify each of the following as qualitative analysis or quantitative analysis: (9.3) K/U
    (a) The formation of a reddish, scaly compound on the surface of a metal object indicates that the object contains iron.
    (b) Two moles of hydrogen gas react with one mole of oxygen gas to produce two moles of water.
31. (a) What is a filtrate?
    (b) What materials are contained in a filtrate? (9.3) K/U
32. Briefly describe what is done during a flame test. (9.3) K/U
33. What can communities do to reduce the quantities of drugs present in their potable water supplies? (9.4) K/U
34. How might increasing the measures to remove drugs from water sources cause additional environmental problems? (9.4) K/U
35. What are three advantages of conducting reactions in aqueous solutions? (9.5) K/U
36. The compound ammonium phosphate, $(NH_4)_3PO_4$, dissociates in solution. What is the ratio of ammonium ions to phosphate ions in the solution? (9.5) K/U
37. Explain how the amount of a chemical in solution can be determined if you know the amount concentration of the solution and the volume used. (9.5) K/U

Understanding

38. A reaction takes place between sodium iodide, NaI, and lead(II) chlorate, $Pb(ClO_3)_2$, producing lead(II) iodide, PbI_2, and sodium chlorate, $NaClO_3$. (9.1) K/U T/I C
    (a) Write a total ionic equation for the reaction. (Refer to Table 1 in Section 9.1.)
    (b) In the equation from Part (a), draw a line through each spectator ion to obtain the net ionic equation.
39. Not all chemical reactions can be represented by net ionic equations. Why can you not write a net ionic equation to show what happens in each of the following situations? (9.1) K/U
    (a) Copper metal, Cu(s), is placed into a solution of iron(III) chloride, $FeCl_3(aq)$. (Refer to Figure 4 in Section 4.4.)
    (b) Carbonic acid, $H_2CO_3(aq)$, decomposes into carbon dioxide and water.
40. Use Table 1 in Section 9.1 to classify each of the following ionic compounds as very soluble or slightly soluble at room temperature: (9.1) K/U
    (a) $AgNO_3$ (e) HI
    (b) Na_3PO_4 (f) CuCl
    (c) $(NH_4)_2CO_3$ (g) BaS
    (d) $PbSO_4$ (h) $Ca(NO_3)_2$

41. In each of the following situations, a chemical reaction occurs. For each reaction, write the balanced formula equation, the total ionic equation, and the net ionic equation. (Refer to Table 1 in Section 9.1.) (9.1) K/U C
   (a) Potassium carbonate, K_2CO_3, is used in the production of glass and soap. A solution of potassium carbonate is added to a solution of silver nitrate, $AgNO_3(aq)$, which is used in photography and as a medical disinfectant.
   (b) Calcium chloride, $CaCl_2$, is the primary component of road salt. A solution of calcium chloride is mixed with a solution of lead(II) nitrate, $Pb(NO_3)_2(aq)$.
   (c) Phosphoric acid, $H_3PO_4(aq)$, is used to make fertilizers and some detergents. In one chemical reaction, phosphoric acid is added to a calcium chloride solution, $CaCl_2(aq)$, to make a new phosphate compound.
   (d) Toxic mercury(I) nitrate, $Hg_2(NO_3)_2$, solution is added to a sodium hydroxide solution to form a yellow-orange precipitate.
   (e) In the "pickling" of metals, a strong acid is used to remove oxides and other impurities from the surface of the metals. During the process, the acid might react with the metal also. During one pickling process, hydrochloric acid, $HCl(aq)$, reacted with iron, producing hydrogen gas.
42. Refer to Table 1 in Section 9.1 to list two cations and one anion that are always spectator ions when they are in a chemical reaction. (9.1) K/U T/I
43. Write a formula equation, a total ionic equation, and a net ionic equation for each of the following chemical reactions. (Assume that a reaction does occur, in each case.) (Refer to Table 1 in Section 9.1.) (9.1) K/U C
   (a) $Hg_2(NO_3)_2(aq) + 2\ KCl(aq) \rightarrow$
   (b) $Mg(s) + 2\ HCl(aq) \rightarrow$
   (c) $2\ K(s) + MgSO_4(aq) \rightarrow$
44. Classify each of the following types of contaminants mixed with water as a homogeneous mixture or a heterogeneous mixture: (9.2) K/U
   (a) physical contaminants
   (b) biological contaminants
   (c) chemical contaminants
45. For each of the following types of contaminant, explain what part or parts of a municipal water treatment process would remove the contaminant: (9.2) K/U
   (a) twigs and pieces of plastic
   (b) suspended particles
   (c) sand and clay
   (d) calcium and magnesium ions
46. (a) List three micro-organisms that may contaminate a water source.
   (b) What two parts of a water purification process kill micro-organisms in the water?
   (c) What part of a water purification process actually removes biological impurities from the water? (9.2) K/U
47. At different stages in the water treatment process, different materials are added to the water. Why do you think the quantity of each material added has to be carefully controlled? (9.2) T/A A
48. How does spectroscopy analyze a tested substance both quantitatively and qualitatively? (9.3) K/U
49. Name two steps of the water treatment process that could remove some medications from water sources. (9.2, 9.4) K/U A
50. A 250 mL sample of a sodium hydroxide solution, $NaOH(aq)$, with a concentration of 1.2 mol/L, reacts with excess sulfuric acid, $H_2SO_4(aq)$, to produce water and dissolved sodium sulfate, $Na_2SO_4(aq)$. Refer to **Figure 1**, which shows how stoichiometry is used in chemical reactions, as you answer the following questions: (9.1, 9.5) K/U T/I C

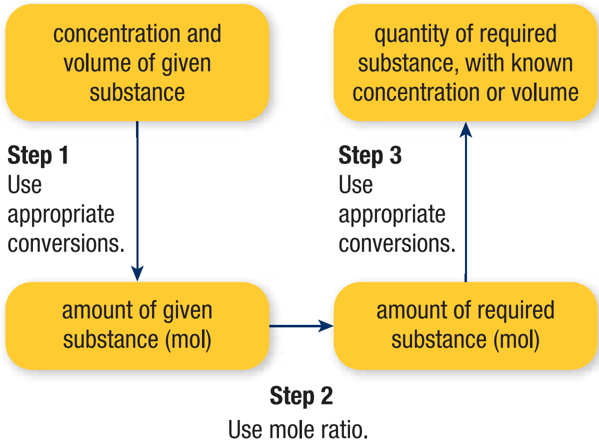


Figure 1

   (a) Write a balanced formula equation for the reaction.
   (b) How might you find the amount of sodium hydroxide present in the sample of solution?
   (c) How can you find the mole ratio needed for Step 2?
   (d) How would you convert the amount of sodium sulfate produced to the mass of sodium sulfate produced?
   (e) How would you find the concentration of the final sodium sulfate solution?

51. Suppose you are going to add 450 mL of a 0.10 mol/L hydrochloric acid, HCl(aq), solution to 250 mL of a 0.20 mol/L sodium hydroxide, NaOH(aq), solution. (9.5) K/U T/I
   (a) How do you determine which reactant is the limiting reagent?
   (b) What is the limiting reagent in this chemical reaction?
   (c) What is the theoretical yield of sodium chloride, NaCl?
52. (a) Write the dissociation equation for solid ammonium sulfate, $(NH_4)_2SO_4(s)$, in water. Include the ratio of moles underneath each formula.
   (b) What is the ratio of ammonium ions to sulfate ions in the solution?
   (c) How does the final concentration of ammonium ions compare to the initial concentration of ammonium sulfate?
   (d) How does the final concentration of sulfate ions compare to the initial concentration of ammonium sulfate? (9.5) K/U T/I C

Analysis and Application

53. (a) Write a net ionic equation for a neutralization reaction of your choice.
   (b) Explain one major advantage of using the total ionic equation for this reaction compared with using this net ionic equation. (9.1) T/I C
54. When an iron nail is placed in a solution of copper(II) nitrate, $Cu(NO_3)_2(aq)$, solid copper forms on the nail, and some of the iron in the nail goes into solution. Initially, the other product is iron(II) nitrate, $Fe(NO_3)_2(aq)$. (9.1) K/U C
   (a) Write a formula equation for the reaction that takes place.
   (b) Write a total ionic equation for the reaction.
   (c) Write a net ionic equation for the reaction.
55. Over several months, the water in a water trap at a golf course turned green. Then, all the fish in the water died. Answer the following questions about this situation. (Refer to Table 1 in Section 9.2.) (9.2) K/U C A
   (a) What is the most likely contaminant in the water?
   (b) Why did the water turn green?
   (c) What happened that caused the fish to die?
   (d) Suggest several ways of resolving this problem and preventing it from happening in the future.
56. One part of water treatment involves the processes of flocculation, coagulation, and sedimentation. Copy **Figure 2** into your notebook. In the cells, write the names of the processes, in the correct order, followed by a brief description of each process. (9.2) K/U

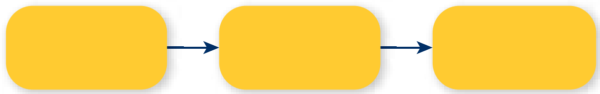


Figure 2

57. The goal of water treatment is not to purify water, but to make the water potable. How might the process of water treatment be different if the goal were to purify water instead of making it potable? (9.2) K/U T/I A
58. **Table 1** shows the relative solubilities of certain compounds in water at room temperature. In the table, "V" means "very soluble," and "S" means "slightly soluble." How might you separate the cations in a solution containing a mixture of Ag^+, Fe^{3+}, K^+, and Ca^{2+}? (9.1, 9.3) T/I A

Table 1 Solubility of Cation–Anion Combinations

	Anions			
Cations	Cl^-	OH^-	S^{2-}	CO_3^{2-}
Ag^+	S	S	S	S
Fe^{3+}	V	S	S	S
K^+	V	V	V	V
Ca^+	V	S	V	S

59. Mining operations can result in pollution involving heavy metals other than those being deliberately mined. For example, gold-mining operations, such as the Golden Giant Mine near Marathon, Ontario, can result in land and water being polluted by copper, cobalt, and nickel ions. One way such metal ions can be removed from soil and water is by electrolysis, during which the metal ions are deposited as metal on an electrode inserted into moist soil or water (Section 5.6). This technique is called "electroplating." How do you think electroplating can be used to remove metal ions from contaminated water? (9.2, 9.3) T/I A
60. When performing a flame test, you must be certain that the nichrome wire is clean. One common way to clean the wire is to dip it into an acid and then hold it in the flame until the flame is colourless. The colour of the flame of the cation in an acid does not mask the colour of the flame of any other cation. Using this information and Table 1 in Section 9.3, determine what cation is present in the acid. (9.3) T/I A

61. You can remove certain cations from solution by adding anions that form a compound with the cations that is only slightly soluble. Why do you need to add excess anions to remove all of the cations from the solution? (9.3) T/I

62. Some stores sell fireplace logs that burn with different colours of flames. These "logs" might be made of tightly rolled paper obtained from a paper recycling plant in Markham, Ontario; waste wood materials from the forestry industry near Kapuskasing; or even from real wood. (9.3) K/U A
    (a) Suggest what kind of substance might be added to the logs to make the flames different colours.
    (b) Name the cation that is used when a log produces each of the following colours of flame. If there is more than one possible cation, include all possibilities. (Refer to Table 1 in Section 9.3.)
        (i) bright red
        (ii) yellow-green
        (iii) violet
        (iv) green

63. Determine the minimum volume of 0.35 mol/L potassium hydroxide solution, $KOH(aq)$, needed to precipitate all the calcium ions in 400 mL of a 0.20 mol/L calcium nitrate solution, $Ca(NO_3)_2(aq)$. (9.5) T/I

Evaluation

64. A student is designing a demonstration in which he will make magnetic letters, numbers, and other symbols to use on a magnetic board to show other students the difference between a formula equation, a total ionic equation, and a net ionic equation. (9.1) K/U T/I C
    (a) Which type of equation do you think he should show first? Second? Third? Explain your reasoning.
    (b) Choose an example of a double-displacement reaction that is a precipitation reaction.
    (c) Write a lesson plan that shows how the student can use your reaction to show the differences between the three types of equations.

65. Joaquin and Kim have a solution containing sodium ions, Na^+, barium ions, Ba^{2+}, and lead(II) ions, Pb^{2+}. They are going to use solutions of ammonium chloride, $NH_4Cl(aq)$, and potassium sulfate, $K_2SO_4(aq)$, to separate the cations in the solution.
    - Joaquin adds excess ammonium chloride solution first, filters out the precipitate, then adds potassium sulfate solution to the filtrate.
    - Kim adds excess potassium sulfate solution first, filters out the precipitate, then adds ammonium chloride solution to the filtrate.

    Which procedure will separate the ions? Explain. (9.1, 9.3) K/U T/I A

66. A student has a solution containing both silver ions, Ag^+, and lead(II) ions, Cu^+. She wants to separate the ions. She says that if she adds a sodium chloride, $NaCl$, solution, the silver ions will precipitate as solid silver chloride, $AgCl(s)$. (9.1, 9.3) K/U T/I
    (a) Will this procedure successfully separate the silver ions from the lead(II) ions? Explain why or why not.
    (b) If your answer to Part (a) is negative, what would you do instead to separate the ions?

67. Flame tests are occasionally unsuccessful because of other ions contaminating the test. The most common contaminant is the sodium ion. The yellow sodium flame may mask the colour of many other cations also in solution. One common source of sodium ion contamination is the sodium ions present on human skin from perspiration. What do you think could be done to minimize sodium contamination in a flame test? (9.3) T/I A

Reflect on Your Learning

68. From what you know about solubility, do you agree with the statement "Even goldfish are slightly soluble"? Explain your reasoning. T/I

69. One reason why barium sulfate, $BaSO_4$, is used in tests in the human GI tract is that it is only slightly soluble. T/I A
    (a) Name two other compounds you think might have a similar solubility to that of barium sulfate. (Refer to Table 1 in Section 9.1.)
    (b) Think about how barium sulfate is used in GI tests. What factors other than solubility might need to be considered in choosing a different compound for this use?

70. Think about the water quality where you live. T/I A
    (a) How do you think your life would change if sanitation and water quality decreased? Try to include both direct and indirect effects in your answer.
    (b) Health Canada works with Indian and Northern Affairs Canada to assist First Nations in ensuring safe drinking water in any communities located south of latitude 60° north. The most common problem with drinking water in these communities is bacteriological contamination. What do you think could be done to ensure safe drinking water for these communities?

71. Think about the water that you use in your home. T/I A
    (a) Is your water hard or soft? What evidence do you have to support your answer?
    (b) Do you have a home water softener? If so, what are its benefits?

72. (a) Does the tap water at your home contain fluoride? If you live in a town or city, how could you find out?
(b) Why is fluoride added to some sources of potable water? K/U

73. When preparing barium sulfate, $BaSO_4$, for a "barium meal," is it more important that the amount of barium ion be measured exactly or that the amount of sulfate ion be measured exactly? Explain your answer. T/I A

74. Discuss the problem of medications in community water sources with members of your family. K/U U A
(a) How do members of your family dispose of prescription and over-the-counter medications that are outdated or otherwise unused?
(b) Develop a plan for proper disposal of these medications so that they do not enter the water supply.

75. The number of ions formed when an ionic compound dissociates helps determine the properties of the solution. For example, the freezing point of water is lowered for each entity dissolved in the water. Use this knowledge to predict which compound—calcium chloride, $CaCl_2$, or sodium chloride, $NaCl$—would be better at preventing the formation of ice on roads in winter. Assume the same amount of each compound is used. T/I A

Research

GO TO NELSON SCIENCE

76. Formation of a precipitate is only one indication of a double-displacement reaction. For each chemical reaction described below,
- research a typical reaction
- write a formula equation and a total ionic equation
- draw a line through each spectator ion
- name the molecular compound(s) that forms

T/I C A

(a) an acid–base reaction
(b) the reaction of an acid with a carbonate
(c) the reaction of an acid with a sulfide

77. Canadian cities need to have water treatment plants that can make drinking water safe for millions of people. Choose a Canadian city. Research its source of drinking water and what type of water treatment is used. A

78. The chemistry of fireworks is similar to what happens during a flame test. T/I A
(a) Are any materials other than ionic compounds used to produce colours in fireworks? If so, provide examples.
(b) Find out what cations are responsible for each colour of fireworks.
(c) What materials are used in fireworks to make the temperature high enough for cations to produce their colourful displays?

79. Different indicators are used in qualitative analysis to detect the presence of certain substances. Research each of the following reagents. Create a table, and use it to organize the following information on each indicator:
- the substance for which it tests
- the appearance of a positive test result
- why you might want to know these results
- whether or not the results provide any quantitative information T/I C A

(a) Benedict's solution
(b) Biuret solution
(c) Sudan III stain

CHAPTER

10 Acids and Bases

KEY CONCEPTS

After completing this chapter you will be able to

- understand some environmental effects of acids and bases and suggest ways to minimize negative effects
- identify the characteristic properties of acids and bases
- use the Arrhenius theory to explain the properties of acids and bases
- distinguish between ionization and dissociation
- explain how strong and weak acids differ
- use stoichiometry to determine the concentration of an acid or base
- conduct a titration to determine the concentration of an acid or base

What Are Acids and Bases?

Hydrangeas are popular flowering shrubs that can add a great deal of colour to a garden. What do hydrangeas have to do with chemistry? Hydrangeas have captured the curiosity of chemists because their colour depends on the acidity (pH) of their soil. Acidic soils produce blue hydrangeas, but basic soils produce pink hydrangeas. Gardeners have learned to manipulate soil pH by adding different substances to the soil. Adding compost or alum (hydrated aluminum potassium sulfate), $AlK(SO_4)_2 \cdot 12H_2O$, makes the soil acidic. Crushed limestone (mostly calcium carbonate, $CaCO_3$) makes the soil basic.

Chemists have discovered that aluminum ions also play a key role in determining the colour of hydrangeas. As soil becomes more acidic, the solubility of naturally occurring aluminum compounds in soil increases. This allows the plant to absorb more aluminum ions. High aluminum concentrations are toxic to many plants and animals, but not to hydrangeas. These ions bind to the pigment molecules in the petals of the plant to produce a new substance that is blue. When the soil is basic, the solubility of aluminum compound in soil is low. Without aluminum ions, the pigment molecules in hydrangeas change back to their pink form.

The pigment compounds in hydrangeas belong to a group of compounds called anthocyanins. Autumn leaves and brightly coloured fruits and vegetables such as berries, blood oranges, and red cabbage also contain anthocyanins. Like hydrangeas, the juice from these foods also changes colour as the pH changes. In fact, red cabbage juice changes colour at least four times as its pH increases from 1 to 14.

Effect on pigment colour and pH are two characteristic properties of acids and bases. In this chapter, you will review other properties of acids and bases and examine theories that help to explain these properties.

STARTING **POINTS**

Answer the following questions using your current knowledge. You will have a chance to revisit these questions later, applying concepts and skills from the chapter.

1. What acids and bases do you use, beyond the science laboratory?
2. What are acids and what properties do they have in common?
3. What are bases and what properties do they have in common?
4. Why do acids have similar properties? Why do bases have similar properties?
5. Why do acids and bases vary in strength?
6. How can you predict the quantity of acid required to neutralize a base?

Mini Investigation

The pH Rainbow Column

Skills: Performing, Observing, Communicating

SKILLS HANDBOOK A1, A2.4, A3.4

Universal indicator is an acid–base indicator that undergoes several colour changes as its pH changes (**Table 1**). In this activity you will use the colour changes of universal indicator to monitor the neutralization of an acid with a base.

Equipment and Materials: chemical safety goggles; lab apron; large test tube; Berol pipette; two 150 mL beakers; test-tube rack; dropper bottles containing saturated sodium carbonate solution, $Na_2CO_3(aq)$, dilute hydrochloric acid, HCl(aq) (0.1 mol/L), and universal indicator solution

The solutions of sodium carbonate and hydrochloric acid are both irritants. Avoid skin contact. If some does spill on your skin, wash the affected area with plenty of cool water. Report any spills to your teacher.

1. Put on your chemical safety goggles and lab apron.
2. Fill the test tube to within 4 cm of the top with hydrochloric acid. Pour the acid into a beaker.
3. Add about 5 drops of universal indicator solution to the acid in the beaker. Swirl to mix.
4. Carefully pour the contents of the beaker back into the test tube. Record your observations.
5. Pour about 15 mL of sodium carbonate solution into the second beaker. Fill the pipette with this solution.
6. Tilt the test tube slightly. Slowly add the sodium carbonate solution to the test tube, letting it flow alongside the acid solution. The carbonate solution sinks to the bottom because it is denser than the acid.
7. Place the test tube in the test-tube rack and observe.
8. Dispose of all the chemicals as directed by your teacher and wash your hands thoroughly.

A. Estimate the pH of each part of the mixture. T/I

B. Suggest an explanation for the variation in pH. T/I

C. Why was it necessary to add the sodium carbonate solution from the top? T/I

Table 1 Universal Indicator Colours

Colour		Solution is . . .
	red	very acidic
	orange/yellow	moderately acidic
	green	neutral
	blue	moderately basic
	purple	very basic

10.1 Properties of Acids and Bases

Canadians' teeth have never been healthier. Advances in dental care during the twentieth century have greatly reduced the incidence of dental decay. One remaining concern, however, is the number of Canadians whose teeth are wearing out faster than expected. The cause? Acid erosion (**Figure 1**). Our modern diet is rich in acidic foods. These include natural, nutritious foods such as citrus fruits, pure fruit juices, and tomatoes. But they also include processed foods such as sour candy and pop. Citric acid occurs naturally in oranges and limes. The tart taste of cola is caused by a combination of phosphoric and carbonic acids. In fact, pop would taste almost as sour as vinegar if not for its sweeteners. Acids prematurely wear away teeth, which can make teeth more sensitive. A simple way to solve this problem is to consume fewer acidic foods and drinks, especially those of little nutritional value like pop. Otherwise, you might need extensive dental work to the repair the damage.

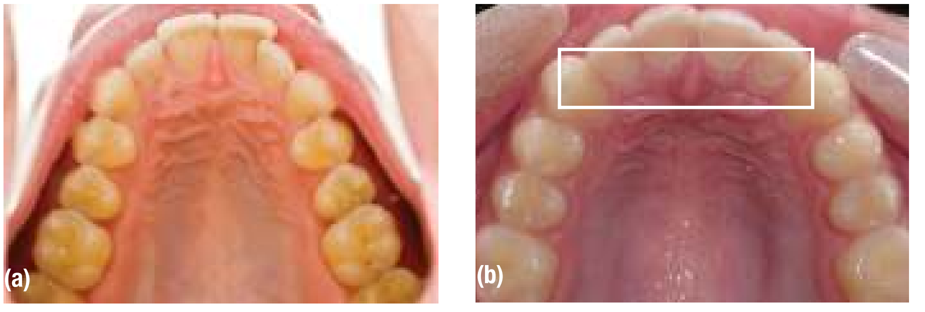


Figure 1 Compare (a) a teenager's healthy teeth and (b) a teenager's teeth damaged by acid erosion. The frame highlights where the enamel has worn away on the incisors.

CAREER **LINK**

Dentists and dental hygienists can advise you on how to care for your teeth. To find out more about these careers,

GO TO NELSON SCIENCE

LEARNING **TIP**

"Alkali" in the Periodic Table
The elements in Group 1 are sometimes called "alkali metals." Recall that the oxides of these metals form basic solutions. Similarly, Group 2 elements are the alkaline earth metals. They, too, react to form basic solutions. Basic = alkali.

To understand why acidic foods can be so damaging, we can examine the properties and structure of tooth enamel. Tooth enamel is the hard material on the outer surface of a tooth. Much of tooth enamel consists of the mineral calcium hydroxylapatite, $Ca_{10}(PO_4)_6(OH)_2(s)$. Hydrogen ions in acidic fluids can break down or demineralize this compound:

$$Ca_{10}(PO_4)_6(OH)_2(s) + 8\ H^+(aq) \rightarrow 10\ Ca^{2+}(aq) + 6\ HPO_4{}^{2-}(aq) + 2\ H_2O(l)$$

This reaction is similar to the reaction that occurs when vinegar is used to remove the mineral buildup in a clogged shower head. Demineralization softens teeth, making them more easily worn away. That is why brushing your teeth immediately after drinking cola (pH 3) may not be wise. This is when tooth enamel is softest.

Acids have a number of properties in common. Reactivity with certain minerals and low pH are two characteristic properties of acids.

Acids and Bases in History

We have known about acids and bases for centuries. Ancient winemakers, for example, knew that exposure to air makes wine sour. Today we know that wine sours because oxygen reacts with ethanol in the wine to produce ethanoic acid (acetic acid). Recall that vinegar is a dilute solution of ethanoic acid. In fact, the word "acid" comes from the Latin word *acidus,* meaning sour.

The use of bases (also known as alkalis) also has a long history. The word "alkali" comes from the Arabic term *al-qilwi*, meaning "basic." Many ancient civilizations discovered that the ashes from a wood fire produced a corrosive mixture when dissolved in water. They also discovered that cooking wood ashes with animal fats or vegetable oils produced an effective cleaner.

Wood ashes are rich in potassium carbonate, K_2CO_3. You will soon learn how potassium carbonate dissolves in water to produce a basic solution. Soap is still produced today using essentially the same ingredients (a base and fat) in a process called saponification (**Figure 2**).

Figure 2 Soap can be made by cooking vegetable oils in a concentrated solution of sodium hydroxide.

People in many cultures developed a practical or empirical understanding of acids and bases, founded on their properties, long before we knew their chemical structures. **Table 1** compares some of the properties of acids and bases.

Table 1 Properties of Acids and Bases

Property		Acid	Base
pH		less than 7	greater than 7
electrical conductivity		conductivity varies	conductivity varies (**Figure 3**)
taste (Never taste chemicals in the lab.)		sour	bitter
feel (Never touch chemicals that are labelled "corrosive.")		no special feel	slippery
colour with acid–base indicator	bromothymol blue		
	phenolphthalein		
	methyl orange		
	litmus		
neutralization reactions		acids neutralize bases	bases neutralize acids

Figure 3 The fluid paste in an alkaline cell must be a good conductor of electricity in order for the cell to generate electricity. These cells are called "alkaline" because they contain the base potassium hydroxide, KOH.

Reactions of Acids

We can also distinguish acids from other substances by their characteristic reactions. In Chapter 4, you learned that acids react with metals above hydrogen on the activity series to produce hydrogen gas. For example, **Figure 4** shows a typical acid, hydrochloric acid, in two different reactions. With magnesium metal, hydrochloric acid reacts to release visible bubbles of hydrogen:

$$Mg(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

Hydrochloric acid also reacts with carbonate compounds to produce carbon dioxide:

$$CaCO_3(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$$

Figure 4 Acids react with (a) reactive metals (such as magnesium) to produce hydrogen gas and (b) carbonate compounds (such as calcium carbonate in chalk) to produce carbon dioxide gas.

Reactions of Bases

Bases also undergo some characteristic reactions: they react with carbon dioxide to form carbonates. For example, blowing carbon dioxide into a calcium hydroxide solution (limewater) produces a calcium carbonate precipitate (**Figure 5**):

$$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$$

Bases also react with oils and fats to produce soap. If a base comes in contact with your skin it feels slippery because it is reacting with oils on your skin to form a soap-like mixture.

Figure 5 When carbon dioxide bubbles through a straw into a solution of calcium hydroxide (limewater), a precipitate turns the liquid cloudy. This reaction is used in the familiar limewater test for carbon dioxide.

Nomenclature of Acids and Bases

Many traditional names have been used to identify acids and bases. Today, a few of these names remain in everyday use. For example, early chemists of the sixteenth century learned to make a highly corrosive fluid called "muriatic acid." Today, pool chemical suppliers sell muriatic acid (hydrochloric acid) to lower the pH of pool water. Similarly, caustic soda and lye are two common names for the same base: sodium hydroxide. Sodium hydroxide is the active ingredient in some drain cleaners.

Today, there are far too many acids and bases for the use of common names to be either safe or practical. Fortunately, chemists have developed a more systematic method of naming these compounds.

Naming Acids

You can usually tell that a compound is an acid if its chemical formula starts with one or more hydrogen atoms. (Water is an exception to this rule.) Acids are sometimes represented by the general formula HA(aq). "H" represents hydrogen and "A" represents the rest of the acid's formula. We add the state symbol (aq) because most acids show their acidic properties only when they are dissolved in water.

Chemists use two different systems to name acids, depending on whether or not the compounds contain oxygen.

Table 2 Acids without Oxygen

Acid name	Chemical formula
hydrofluoric acid	HF(aq)
hydrochloric acid	HCl(aq)
hydrobromic acid	HBr(aq)
hydroiodic acid	HI(aq)
hydrosulfuric acid	H_2S(aq)
hydrocyanic acid	HCN(aq)

ACIDS WITHOUT OXYGEN

The names of acids without oxygen start with the prefix *hydro-* and end with *-ic acid*. The stem of the acid name comes from the element or group following hydrogen in the formula. For example, HCl(aq) is called hydrochloric acid because it contains chlorine. *Chlor*, from "chlorine," is the stem of the acid's name. Halogens have similar properties and similar electron arrangements. You would therefore expect halogens to form acids with similar chemical formulas (**Table 2**). These include hydrofluoric acid, HF(aq), hydrochloric acid, HCl(aq), and hydroiodic acid, HI(aq).

ACIDS WITH OXYGEN

oxyacid an acid that includes oxygen in its formula

The name of an acid that contains oxygen is based on the name of its oxyanion. These acids are called **oxyacids**. **Table 3** lists some oxyacids. Note that their names do not start with the prefix *hydro-*. Recall from Section 2.4 that a negatively charged polyatomic ion that contains oxygen is called an oxyanion.

If the oxyanion name ends in *-ate*, the acid name ends in *-ic acid*. For example,

ion	acid
NO_3^-	HNO_3(aq)
nitrate ion	nitric acid

If the oxyanion name ends in *-ite*, the acid name ends in *-ous acid.* For example,

ion	acid
NO_2^-	HNO_2(aq)
nitrite ion	nitrous acid

Similarly, the hypochlorite ion is related to hypochlorous acid. **Table 4** lists the oxyanions that contain the same elements as the chlorate ion, ClO_3^-. Note that all the halogens follow this pattern for the names and formulas of their oxyanions.

Table 3 Acids with Oxygen (Oxyacids)

Acid name	Chemical formula	Parent oxyanion
nitric acid	$HNO_3(aq)$	nitrate, NO_3^-
chloric acid	$HClO_3(aq)$	chlorate, ClO_3^-
carbonic acid	$H_2CO_3(aq)$	carbonate, CO_3^{2-}
sulfuric acid	$H_2SO_4(aq)$	sulfate, SO_4^{2-}
phosphoric acid	$H_3PO_4(aq)$	phosphate, PO_4^{3-}

Table 4 Oxyanions of Chlorine

Anion	Formula of anion
perchlorate	ClO_4^-
chlorate	ClO_3^-
chlorite	ClO_2^-
hypochlorite	ClO^-

LEARNING **TIP**

Oxyanion Review

Recall that *per, ite,* and *hypo___ite* are used to indicate the number of oxygen atoms in the oxyanion as compared to the −*ate* anion:
per means 1 more oxygen atom.
ite means 1 less oxygen atom.
hypo + *ite* means 2 less oxygen atoms. You can apply the pattern of oxyanion names in Table 4 to any of the parent oxyanions in Table 3. This allows you to write ions such as the nitrite ion, NO_2^-. With practice you will become familiar with the oxyanions, the number of oxygen atoms in each, their charges, and their related oxyacids.

Tutorial **1** Naming Acids

Figure 6 summarizes a procedure that you can follow when naming acids. This procedure assumes that you already know that the compound is an acid: HA(aq). You will also need to refer to Tables 3 and 4 when naming acids that include oxygen in their formulas.

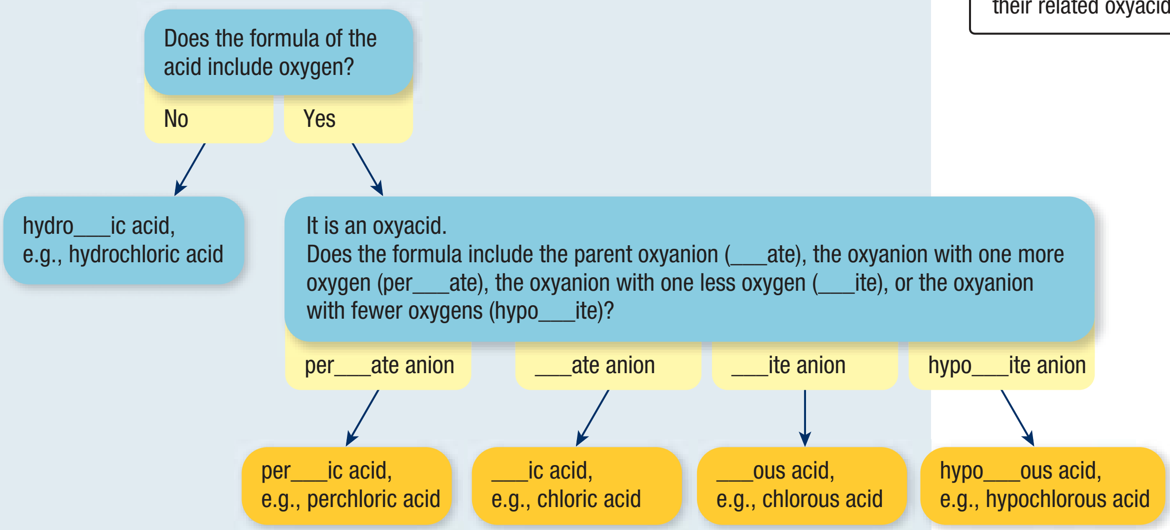


Figure 6 A summary of how to name acids

Sample Problem 1: Naming an Acid That Includes Oxygen

Write the name of the acid with formula $H_2SO_3(aq)$.

Step 1. Start at the top of the flow chart in Figure 6. Establish whether oxygen is present.

Oxygen is present, so it is an oxyacid. It is related to an oxyanion.

Step 2. Identify the oxyanion.

The oxyanion is SO_3^{2-}, sulfite ion.

Step 3. Establish the name of the acid from the name of the oxyanion.

The acid with formula $H_2SO_3(aq)$ is sulfurous acid.

Sample Problem 2: Writing the Formula of an Oxyacid

Write the chemical formula of bromic acid.

Step 1. Use the name to identify the type of acid.

Since the prefix *hydro-* is not present, bromic acid is an oxyacid.

Step 2. Use the stem of the acid's name and any prefixes or suffixes in the name to determine the oxyanion in the acid's formula.

Since the stem of the acid's name is *brom-* and the suffix is *-ic*, the oxyanion is the bromate ion, BrO_3^-.

LEARNING **TIP**

Non-Oxygen Name

The name of the acid comes from the non-oxygen element in the anion. For example, the acid related to the sulfite ion (which contains sulfur and oxygen) is called "sulfurous acid." Notice that the stem of the acid name changes from *sulf-* to *sulfur-*.

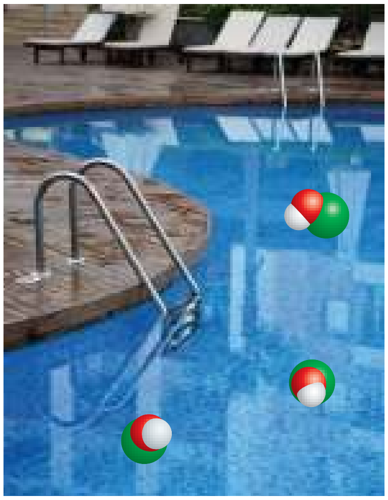


Figure 7 Hypochlorous acid forms in pool water from the chemicals used to disinfect the pool. This acid kills bacteria by passing through the cell membrane and bleaching the cell's interior.

Step 3. Use the zero-sum rule to determine the number of hydrogen ions that are required to produce a neutral acid molecule.

Since the charge of the bromate ion is –1, only one H^+ ion is required to produce a neutral acid molecule.

Step 4. Combine the correct number of H^+ ions with the oxyanion to write the formula for the acid.

The formula of bromic acid is $HBrO_3$(aq).

Practice

1. Name the following acids: K/U
   (a) HBr(aq)
   (b) H_3PO_3(aq)
   (c) H_2S(aq)
   (d) H_2SO_3(aq)
   (e) HIO_4(aq)
   (f) HFO(aq)
2. Write the chemical formula for each of the following acids: K/U
   (a) hydrofluoric acid
   (b) sulfurous acid
   (c) hypochlorous acid (**Figure 7**)
   (d) perbromic acid

Naming Bases

A number of substances produce basic solutions when dissolved in water. However, for this course you only have to learn how to name bases that are ionic hydroxides, such as sodium hydroxide, NaOH(aq), and calcium hydroxide, $Ca(OH)_2$(aq). (You will encounter other bases in future chemistry courses.) Note that the name of the base is the same as the name of the ionic hydroxide.

Hydroxide bases only show their basic properties when they are in solution. Because of this, we always use the symbol (aq) in the chemical formula of a base.

10.1 Summary

- Acids taste sour, have a pH less than 7, conduct electricity (to varying degrees), change the colours of acid-base indicators, and neutralize bases.
- Acids react with metals above hydrogen on the activity series.
- Acids react with carbonate compounds.
- Bases taste bitter, feel slippery, have a pH greater than 7, conduct electricity to some degree, change the colours of acid-base indicators, and neutralize acids.
- Acids that do not have oxygen in the formula are named hydro____ic acid.
- The names of acids with oxygen are based on the name of their oxyanion (Figure 6).
- The names of bases that are ionic hydroxides are simply the names of the ionic hydroxides.

10.1 Questions

1. Chewy sour candies sometimes remain stuck to teeth for several minutes. Why is this a problem? K/U A
2. The reaction of sodium hydrogen carbonate (baking soda), $NaHCO_3$, and vinegar is a safe and effective method of opening a clogged drain (**Figure 8**). C A
   (a) Write the chemical equation for this reaction.
   (b) Explain why using sodium hydrogen carbonate and vinegar may be a greener alternative to using household drain openers, many of which contain sodium hydroxide.

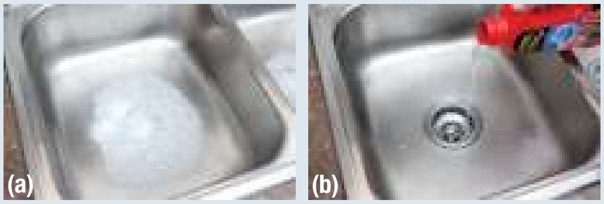


Figure 8 You can open a clogged drain using (a) baking soda and vinegar or (b) a household product that may contain sodium hydroxide.

3. Sulfuric acid is dripped onto samples of three metals: silver, zinc, and aluminum. K/U T/I C
   (a) Use the activity series to determine whether a reaction will occur with each metal.
   (b) Write a chemical equation for each reaction that occurs.
4. Name the following acids: T/I
   (a) $H_2CO_3(aq)$ (in carbonated beverages)
   (b) $HI(aq)$ (used to make organic iodine compounds)
   (c) $H_2S(aq)$ (rotten-egg smell)
   (d) $H_3PO_4(aq)$ (rust remover)
   (e) $HNO_3(aq)$ (used to make fertilizers)
   (f) $HF(aq)$ (used to etch glass)
   (g) $HNO_2(aq)$
   (h) $H_2SO_3(aq)$
   (i) $H_3PO_3(aq)$
   (j) $HIO_4(aq)$
   (k) $HClO(aq)$
5. What information in the chemical formula indicates that a compound is likely to be an acid? K/U
6. Write the chemical formula for each of the following acids: T/I C
   (a) hydrobromic acid
   (b) perchloric acid
   (c) chlorous acid
   (d) hydroiodic acid
7. Name the following bases: T/I
   (a) $Mg(OH)_2(aq)$ (b) $KOH(aq)$
8. How are the chemical formulas of chloric acid and hydrochloric acid similar? How do they differ? K/U T/I
9. Given that the elements chlorine and iodine are in the same chemical family, you would expect them to form compounds with similar chemical formulas. T/I C
   (a) Predict the chemical formula of iodic acid.
   (b) Use your answer from (a) to predict the chemical formulas of iodous acid and hypoiodous acid.
10. The stomach secretes a corrosive mixture of digestive juices that includes hydrochloric acid. Research why the stomach does not digest itself. A
11. Despite its name, the medical condition known as "heartburn" has nothing to do with the heart. Research the cause of heartburn and what can be done to soothe its symptoms. T/I A
12. Research the role of carbon dioxide in the formation of caves in limestone deposits (**Figure 9**). T/I A

Figure 9 Carlsbad Caverns in New Mexico

GO TO NELSON SCIENCE

10.2 Theoretical Acid–Base Definitions

People have had a practical understanding of acids and bases and their applications for centuries. As our understanding of chemistry has grown, several theories have been proposed to explain the properties of acids and bases. In 1777, Antoine Lavoisier discovered that air contained an invisible gas, which he called "vital air," that was necessary for combustion. He also observed that burning sulfur and phosphorus in air produced acidic solutions. This observation led Lavoisier to believe that "vital air" was an element that was common to all acids. He was so confident in this conclusion that he renamed this element "oxygen" from the Greek words *oxys* (sour) and *genes* (born).

The first challenge to Lavoisier's oxygen theory of acids came in 1789. This was when Claude Louis Berthollet showed that hydrocyanic acid, HCN(aq), did not contain oxygen. However, the acidic properties of HCN(aq) were so mild that chemists doubted whether it was a true acid. The oxygen theory of acids remained the dominant acid theory until the early 1800s. That was when Humphrey Davy discovered that muriatic (hydrochloric) acid did not contain oxygen. Rather, it consisted only of hydrogen and a new element that Davy named "chlorine."

CAREER LINK

Theoretical chemists may work in any branch of chemistry, proposing and testing explanations for observations. Many years of education and laboratory experience are necessary. To learn about this career,

GO TO NELSON SCIENCE

Arrhenius Theory of Acids and Bases

By the 1880s, Swedish chemist Svante Arrhenius had developed a theory about electrolytes such as sodium chloride. His theory explained why solutions of electrolytes conduct electricity. According to Arrhenius, when an electrolyte dissolves, its ions dissociate (come apart). This allows them to move freely in the solution and to conduct electricity. In the mid-1880s, Arrhenius modified his theory to include acids and bases. Since acidic and basic solutions also conduct electricity, he concluded that these solutions must also contain ions (**Figure 1**).

Figure 1 Solutions of hydrochloric acid, sodium hydroxide, and sodium chloride are good conductors of electricity because of the ions they contain. According to Arrhenius, pure water is a poor conductor because it does not contain ions.

Arrhenius proposed that a base is an ionic compound that dissociates into cations and hydroxide ions, OH^-, as it dissolves in water. Hydroxide ions give bases their characteristic properties. The dissociation of a base is similar to that of other ionic compounds discussed in Chapter 9. For example, the equation for dissociation of calcium hydroxide is

$$Ca(OH)_2(s) \rightarrow Ca^{2+}(aq) + 2\ OH^-(aq)$$

Arrhenius also proposed that an acid is a molecular compound that ionizes to produce hydrogen ions in water. Hydrogen ions give acids their characteristic properties. For example, the ionization equation for nitric acid is

$$HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$$

Ionization and dissociation are both events in which a compound breaks apart in water, causing the presence of ions in the water. However, there are subtle differences in how these processes occur. Dissociation occurs when water molecules pull the positive and negative ions of an ionic hydroxide (or any soluble ionic compound) apart. The ions dissociate from each other. **Ionization**, however, involves the formation of ions from uncharged molecules. Previously uncharged entities become ionized. Arrhenius likely assumed that the ionization process involved water but did not explain how this occurred.

ionization the formation of ions from uncharged molecules

In summary, ionic compounds dissociate in water; molecular compounds ionize. The chemical equations for these two processes look the same.

LEARNING **TIP**

Dissociation and Ionization

Both dissociation and ionization result in the presence of ions in solution. Dissociation separates ions that already exist in the neutral compound.

$$M^+X^-(s) \rightarrow M^+(aq) + X^-(aq)$$

During ionization, new ions form from a neutral compound.

$$HA(aq) \rightarrow H^+(aq) + A^-(aq)$$

Neutralization Reactions

For an acid–base theory to be useful, it should be able to explain acid–base properties. Let's see how Arrhenius applied his theory to explain the neutralization of hydrochloric acid with sodium hydroxide. Hydrochloric acid conducts electricity very well—much better than some other acids at the same concentration. Arrhenius therefore assumed that it completely ionizes in water, leaving no hydrogen chloride molecules in solution:

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

Similarly, sodium hydroxide completely dissociates into its ions:

$$NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$$

When these solutions are combined, the ions move about freely and react. When we combine the two equations above and consider how the ions react, we are writing a total ionic equation for the reaction:

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow H_2O(l) + Na^+(aq) + Cl^-(aq)$$

The total ionic equation includes some spectator ions. Remember that spectator ions are ions that appear on both sides of the equation and are not involved in the reaction.

$$H^+(aq) + \cancel{Cl^-(aq)} + \cancel{Na^+(aq)} + OH^-(aq) \rightarrow H_2O(l) + \cancel{Na^+(aq)} + \cancel{Cl^-(aq)}$$

After removing the spectator ions, this equation simplifies to

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Since acids ionize to produce hydrogen ions, and bases dissociate to release hydroxide ions, we can simplify any neutralization reaction to

$$H^+(aq) + OH^-\ (aq) \rightarrow H_2O(l)$$

We can test this idea with the neutralization of hydrochloric acid with aluminum hydroxide (the active ingredient in some antacid tablets).

Balanced chemical equation:

$$3\ HCl(aq) + Al(OH)_3(aq) \rightarrow 3\ H_2O(l) + AlCl_3(aq)$$

Total ionic equation:

$$3\ H^+(aq) + 3\ \cancel{Cl^-(aq)} + \cancel{Al^{3+}(aq)} + 3\ OH^-(aq) \rightarrow 3\ H_2O(l) + \cancel{Al^{3+}(aq)} + \cancel{3\ Cl^-(aq)}$$

Net ionic equation:

$$3\ H^+(aq) + 3\ OH^-(aq) \rightarrow 3\ H_2O(l)$$

which simplifies to

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

The Arrhenius theory of acids and bases provides a simple and consistent explanation of the properties of acids and bases. However, as with all scientific theories, it does have limitations. For example, some substances have basic properties and yet do not contain hydroxide ions in their chemical formula. Household ammonia cleaner, $NH_3(aq)$, for example, has a pH of 11 to 12 and can neutralize acids (**Figure 2**). A different theory of acids and bases explains this observation. You may study this theory in future chemistry courses.

Figure 2 An ammonia solution is basic, with a pH of almost 12.

Scientific knowledge changes over time. An essential feature of science is the existence of competing theories to explain the same phenomenon. Multiple competing

Figure 3 Powdered magnesium metal reacts with (a) hydrochloric acid and with (b) ethanoic acid. The foam in test tube (a) indicates that oxygen is produced much more rapidly than in test tube (b).

theories do not make science invalid in any way. Rather, the existence of competing theories pushes scientists to test and refine their theories so that they explain as many examples as possible.

Strong and Weak Acids

The concentration of ethanoic acid (acetic acid) in vinegar is similar to the concentration of the hydrochloric acid that you have used in investigations. Yet vinegar is safe to consume, while handling a comparable solution of hydrochloric acid requires special precautions. Hydrochloric acid also reacts more vigorously with metals (**Figure 3**). Why is there such a difference in the strength of these acids?

According to Arrhenius, the strength of an acid depends on the extent to which it ionizes. Hydrochloric acid is a **strong acid** because it completely ionizes in solution. All of the hydrogen chloride molecules react to form hydrogen ions and chloride ions:

$$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

This formation of many ions explains why hydrochloric acid is a good conductor of electricity (**Figure 4(a)**). We classify ethanoic acid as a **weak acid** because only a small fraction of its molecules ionize:

$$HC_2H_3O_2(aq) \xrightarrow{\text{(partial ionization)}} H^+(aq) + C_2H_3O_2^-(aq)$$

This explains why ethanoic acid is a weak electrolyte and does not conduct electricity as well as hydrochloric acid at the same concentration (**Figure 4(b)**).

strong acid a substance that ionizes completely in water; strong acids have strong acidic properties (e.g., low pH and vigorous reactivity)

weak acid a substance that only partially ionizes in water; weak acids have weak acidic properties (e.g., moderate pH and mild reactivity)

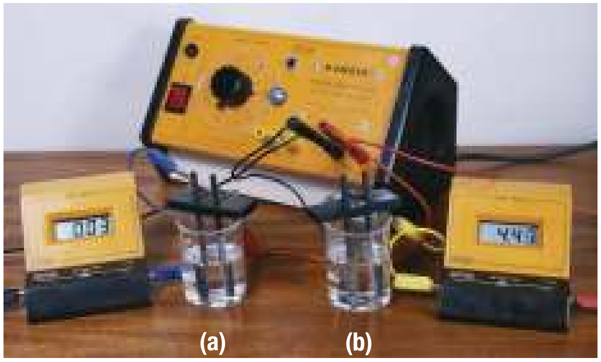


Figure 4 If (a) an ethanoic acid solution has the same concentration as (b) a hydrochloric acid solution, then the hydrochloric acid solution will be a better conductor of electricity.

In general, the strength of an acid, HA(aq), depends on how much it ionizes in aqueous solution:

$$HA(aq) \rightarrow H^+(aq) + A^-(aq)$$

Strong acids ionize almost completely. Weak acids ionize only partially: about 2 % (**Figure 5**).

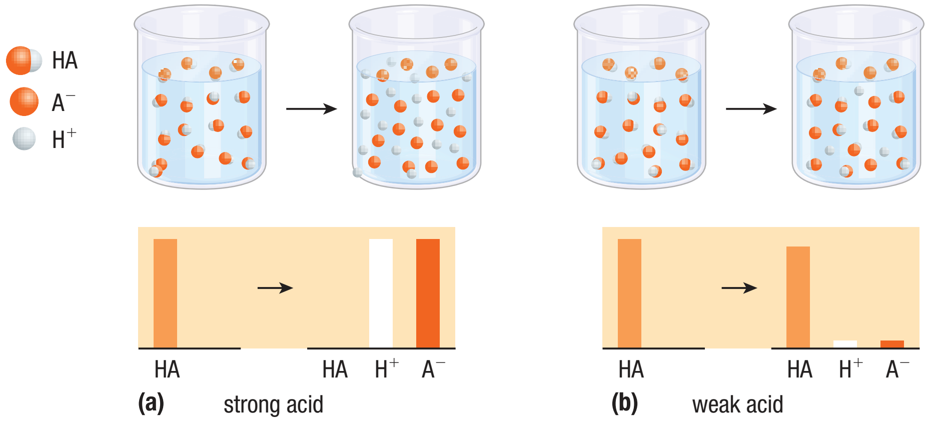


Figure 5 (a) A strong acid, HA(aq), almost completely ionizes into H^+(aq) and A^-(aq). Virtually no molecules of a strong acid remain in solution. (b) A weak acid, HA(aq), partially ionizes into H^+(aq) and A^-(aq). Most molecules of a weak acid remain intact.

LEARNING TIP

"Strong" Is Not Necessarily "Concentrated"

Be sure that you understand the difference between "strong" (completely ionized or dissociated) and "concentrated" (having a relatively large quantity of solute in a given volume of solution). Similarly, try not to mix up "weak" (only slightly ionized or dissociated) and "dilute" (having a relatively small quantity of solute in a given volume of solution).

Table 1 lists some common strong acids and weak acids.

Table 1 Common Strong Acids and Weak Acids

Strong acids	Application	Weak acids	Application
hydrochloric acid, $HCl(aq)$	• stomach acid • used for cleaning or "pickling" steel	phosphoric acid, $H_3PO_4(aq)$	• rust remover • an ingredient in pop
nitric acid, $HNO_3(aq)$	• used in the production of fertilizers and rocket fuel	ethanoic acid, $HC_2H_3O_2(aq)$	• the acid in vinegar • an ingredient in pickled vegetables
perchloric acid, $HClO_4(aq)$	• a powerful bleaching agent	methanoic acid (formic acid), $HCO_2H(aq)$	• produced by ants
sulfuric acid, $H_2SO_4(aq)$	• the acid in car batteries • used in the production of detergents and plastics	carbonic acid, $H_2CO_3(aq)$	• makes rain naturally acidic • an ingredient in pop

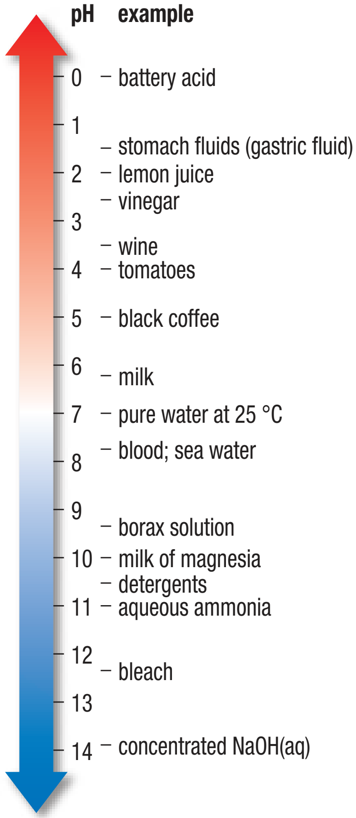


Figure 6 The pH scale. What patterns do you see in the data?

pH AND ACIDITY

You have already learned that the pH is a measure of how acidic or basic a solution is. The pH scale (**Figure 6**) is a numerical scale of all the possible values of pH ranging from 0 to 14. A solution with a pH less than 7 is considered acidic while a solution with a pH greater than 7 is basic. A solution with a pH of 7 is neutral: neither acidic nor basic.

Quantitatively, a change of one pH unit represents a tenfold change in how acidic or basic a solution is. For example, lemon juice, with a pH of 2, is ten times more acidic than pop with a pH of 3. Conversely, aqueous ammonia, with a pH of 11, is ten times more basic than milk of magnesia, with a pH of 10. Two steps along the pH scale represent a 10×10 or a hundredfold change in how acidic or basic a solution is. Lemon juice, with a pH of 2, is 100 times more acidic than tomatoes, with a pH of 4. Conversely, a concentrated solution of sodium hydroxide used in drain openers, with pH 14, is $10 \times 10 \times 10 \times 10 \times 10 \times 10 \times 10 = 10\ 000\ 000$ times more basic than water (pH 7).

Living things can tolerate only small changes in pH. Most freshwater fish, for example, thrive in a pH range of 5.5 to 7.5, depending on the species. "Clean" rain can have a pH of about 5.6. Rain is naturally acidic due to the formation of some carbonic acid. Atmospheric pollution, however, can lower the pH of rain to 4.6. This means that acid precipitation is 10 times more acidic than normal rain. This change can devastate aquatic life in fragile ecosystems (**Figure 7**). In Section 5.3 you learned about some other effects of acidic conditions in the environment.

Figure 7 (a) A healthy salamander embryo and (b) a salamander embryo raised in acidic water

CAREER LINK

Environmental technicians collect data on many properties of soil and water, including their pH. If you are interested in this career, which generally involves fieldwork,

GO TO NELSON SCIENCE

Investigation 10.2.1

Not All Acids Are Created Equal (p. 486)

In this controlled experiment you will predict and explore the reactions of three acids to compare their reactivity.

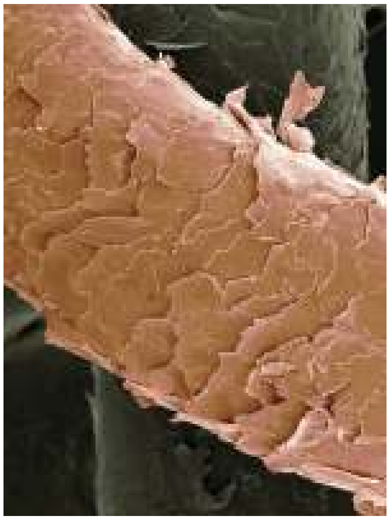


Figure 8 The cuticle of a human hair

In some cases, however, acidic environments can be beneficial. For example, the outer covering of human hair is called the cuticle and consists of tiny overlapping plates (**Figure 8**). Washing hair with a slightly acidic shampoo or a lemon rinse causes these plates of the cuticle to close tightly around the shaft of hair. This makes each hair strand smoother and less likely to tangle with its neighbours. Most shampoos are slightly acidic for this reason.

Mini **Investigation**

pH and Dilution

Skills: Planning, Performing, Observing, Communicating — SKILLS HANDBOOK **A1, A2.4, A3**

In this investigation you will observe the effect of dilution on the pH of an acidic solution.

Equipment and Materials: chemical safety goggles; lab apron; pH meter; 50 mL graduated cylinder; volumetric pipette; pipette pump or bulb; three 150 mL beakers; hydrochloric acid, HCl(aq) (0.1 mol/L); wash bottle of distilled water

Hydrochloric acid is an irritant. Avoid skin and eye contact. If acid spills in your eyes or on your skin, wash the affected area with plenty of cool water.

1. Plan a procedure to prepare 20 mL of a 0.01 mol/L solution of hydrochloric acid, using the equipment and materials listed. Ask your teacher to approve your procedure.
2. Put on your chemical safety goggles and lab apron.
3. Add 20 mL of 0.1 mol/L hydrochloric acid to a clean, dry beaker.
4. Measure the pH of the solution using a pH meter. Record your observations.
5. Rinse the electrode of the pH meter with distilled water.
6. Proceed with procedure, planned in Step 1, once it is approved by your teacher.
7. Measure and record the pH of your solution.
8. Use the same procedure to prepare 20 mL of 0.001 mol/L and 20 mL of 0.0001 mol/L hydrochloric acid. Measure and record the pH of each solution.
9. Dispose of the solutions as directed by your teacher.

A. By what factor was acid solution diluted each time? T/I

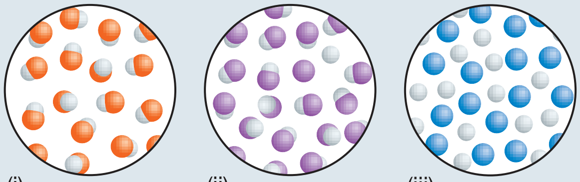
B. What effect did each dilution have on the measured pH value? T/I

C. Based on your observations, predict the pH of a 1.0×10^5 mol/L HCl(aq) solution. T/I

10.2 Summary

- According to the Arrhenius theory, acids ionize in water to produce hydrogen ions and bases dissociate in water, resulting in hydroxide ions. Neutralization reactions are reactions of hydrogen ions with hydroxide ions.
- The strength of an acid depends on how much it ionizes. Strong acids ionize completely. Weak acids ionize partially.
- pH is used to communicate the acidity of a solution. As the hydrogen ion concentration increases, the pH of the solution decreases.
- Each step along the pH scale corresponds to a tenfold change in acidity (for example, pH 2 is 10 000 times more acidic than pH 6).

1. Classify these substances as either Arrhenius acids or bases. Justify your choice. K/U
   (a) $HF(aq)$
   (b) $Ba(OH)_2(aq)$
   (c) $H_2SO_4(aq)$
   (d) $NH_4OH(aq)$
2. Why are theories often described as being "valid" rather than "correct"? Use the Lavoisier oxygen theory as an example. T/I
3. **Figure 9** represents acids and related entities at the molecular level. The white spheres represent hydrogen atoms or ions. The coloured spheres represent the other part of the acid. Which diagram best represents each of the following substances? Explain why in each case. K/U T/I
   (a) hydrochloric acid
   (b) ethanoic acid
   (c) hydrogen chloride gas



(i) (ii) (iii)

Figure 9

4. Explain how an acid can be both weak and concentrated. Use an example to illustrate. T/I
5. Describe a situation in which solutions of a strong acid and a weak acid can have the same pH. K/U
6. Write the total ionic and net ionic equations for each of the following reactions: K/U T/I C
   (a) hydrobromic acid with potassium hydroxide
   (b) nitric acid with barium hydroxide
7. Imagine that you have two solutions of exactly the same concentration. One is a strong acid and the other is a weak acid. Compare the pH of the two solutions. Why is there a difference? K/U
8. (a) Stomach acid has a pH around 2, and orange juice has a pH around 4. How much more acidic is stomach acid than orange juice?
   (b) Assuming that your stomach is healthy, why is the acidity of orange juice not a health risk? T/I
9. What safety precautions should you use when working with strong acids and bases in the laboratory? T/I
10. According to the data in Figure 6, approximately how much more acidic is
    (a) lemon juice than black coffee?
    (b) black coffee than milk of magnesia?
    (c) battery acid than pure water at 25 °C?
    (d) tomatoes than aqueous ammonia T/I
11. According to the data in Figure 6, approximately how much more basic is
    (a) aqueous ammonia than pure water?
    (b) 1 mol/L $NaOH(aq)$ than milk of magnesia?
    (c) tomatoes than lemon juice?
    (d) milk of magnesia than black coffee? T/I
12. **Figure 10(a)** shows the conductivity test for sulfuric acid. A similar result is observed for barium hydroxide. K/U T/I C
    (a) How would Arrhenius explain these observations?
    (b) Write the chemical equation for the neutralization of sulfuric acid by barium hydroxide. Consult the solubility table (Section 4.6) to determine the state of the ionic product.
    (c) **Figure 10(b)** shows the final mixture produced when sulfuric acid is completely neutralized by barium hydroxide. Why is there a difference in conductivity? What would be observed if sodium hydroxide were used instead? Why?

(a) **(b)**

Figure 10 (a) Sulfuric acid alone is an excellent conductor of electricity, as the lit bulb shows. (b) When sulfuric acid is completely neutralized with barium hydroxide, the resulting mixture does not conduct electricity and the bulb no longer lights up.

13. It took several years before Arrhenius's theory of acids and bases became widely accepted. Arrhenius first presented his theory in 1884 while completing his Ph.D. studies in chemistry at the University of Uppsala, Sweden. Initially, his professors were not impressed. In fact, they found his theory so controversial that they gave him the lowest possible passing mark. A
    (a) Research the impact of this rejection on Arrhenius.
    (b) How did the scientific community finally recognize Arrhenius for his contributions to chemistry?

GO TO NELSON SCIENCE

10.3

Acid–Base Stoichiometry

CAREER LINK

Food scientists are concerned about the freshness of the ingredients used in food products. To learn more about the work of a food scientist,

GO TO NELSON SCIENCE

An important application of neutralization reactions is finding the concentration of an unknown solution. The dairy industry, for example, uses a neutralization reaction to monitor the concentration of lactic acid in stored milk. Fresh milk has no lactic acid. Bacteria in milk consume lactose, a milk sugar, and produce lactic acid as a waste product. The older the milk, the more lactic acid is present. Finding the concentration of an unknown solute like lactic acid in milk involves an analysis called titration.

Titration

titration a procedure used to determine the concentration of a solution using a standardized solution

titrant the solution in the burette during a titration

burette a calibrated tube used to deliver variable known volumes of a liquid during a titration

A **titration** is an analytical procedure used to determine the concentration of a solution. During a titration, a measured volume of a standardized solution called the **titrant** is gradually added to a flask containing a measured volume of a solution of unknown concentration. Recall that a standardized solution is a solution whose concentration is precisely known. A burette is used to accurately measure the volume of titrant used. A **burette** is a tube that is typically calibrated with 0.1 mL divisions. The valve at one end of the burette is used to control the volume of titrant that leaves the burette. **Figure 1** shows the titration of sulfuric acid (in the flask) with sodium hydroxide (in the burette), using bromothymol blue indicator. The chemical equation for this neutralization reaction is

$$2\ NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l)$$

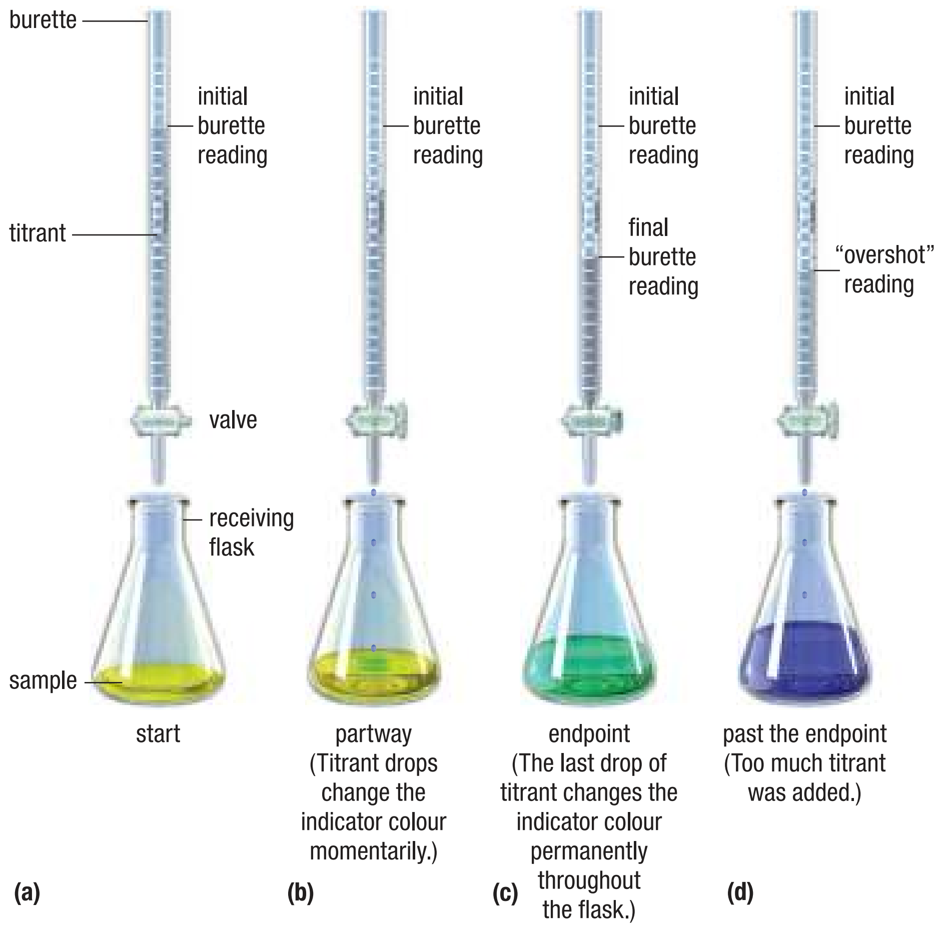


Figure 1 The titration of sulfuric acid with sodium hydroxide (a) A few drops of bromothymol blue indicator make the sulfuric acid solution initially yellow. (b) Adding sodium hydroxide produces a temporary colour change that fades as the flask is swirled. (c) The last drop of titrant changes bromothymol blue permanently to green at the endpoint. (d) The blue colour indicates that the titration has gone past the desired endpoint. This means too much base has been added.

WEB LINK

There are many online videos showing how to perform a titration and how to identify when the endpoint has been reached. To view some of these videos,

GO TO NELSON SCIENCE

The point during an acid–base titration when neutralization is complete is called the **equivalence point**. At this point, the amount of acid and base added to the flask exactly match their mole ratio in the chemical equation. They are stoichiometrically equal. To identify when the equivalence point is reached, chemists look for a sudden change in an observable property of the solution. This could be a sudden change in pH (detected using a pH meter) or the colour of an acid–base indicator. The point in a titration at which this sudden observable change occurs is called the **endpoint**. The endpoint and equivalence point are not the same (**Table 1**). In most cases, we chose an acid–base indicator that changes colour at the equivalence point of the reaction. The colour change is a visual clue that the equivalence point has been reached. The indicator changes colour permanently when a very slight excess of titrant is added. For example, the endpoint of the titration in Figure 1 occurs when bromothymol blue changes permanently to green.

You might wonder whether titrations are done in "real life." Indeed they are! For example, companies that process pickles have to be sure that the liquid in the pickle jar is acidic enough to preserve the pickles. Usually they use an electronic pH meter to monitor the acidity, but they also periodically check the accuracy of the pH meter by performing a titration.

When titrating, it is important to make accurate and precise observations. You need to ensure that you accurately determine the volume of titrant required to reach the equivalence point. To improve the quality of your data, you should repeat the procedure several times, determining the volume used each time. Three trials with results that are within 0.2 mL are normally required for a reliable analysis.

Although the equivalence point and the endpoint are not the same, they coincide if the correct indicator is selected. Thus, the endpoint tells us that the equivalence point has been reached.

equivalence point the point in a titration when neutralization is complete

endpoint the point during a titration when a sudden change in an observable property of the solution occurs; usually a change in colour of an acid–base indicator or a significant change in pH

Table 1 Comparing Equivalence Points and Endpoints

Endpoint	Equivalence point
sudden observable change (e.g., in colour, pH, or conductivity)	theoretical: cannot be observed
occurs when an acid–base indicator reacts with the titrant to produce a lasting colour change	occurs when an acid and a base have completely neutralized each other
depends on the type of indicator used	depends on the stoichiometry of the neutralization reaction

Choosing an Acid–Base Indicator

A titration gives accurate results only when its endpoint coincides with its equivalence point. This means that the acid–base indicator chosen for the titration must change colour when the equivalence point is reached. How do you know which is an appropriate indicator for a titration? To answer this question, we can first examine the pH changes that occur during the neutralization of a strong acid with a strong base (**Figure 2**). Imagine that a pH meter is recording the pH of the solution in the flask throughout the titration.

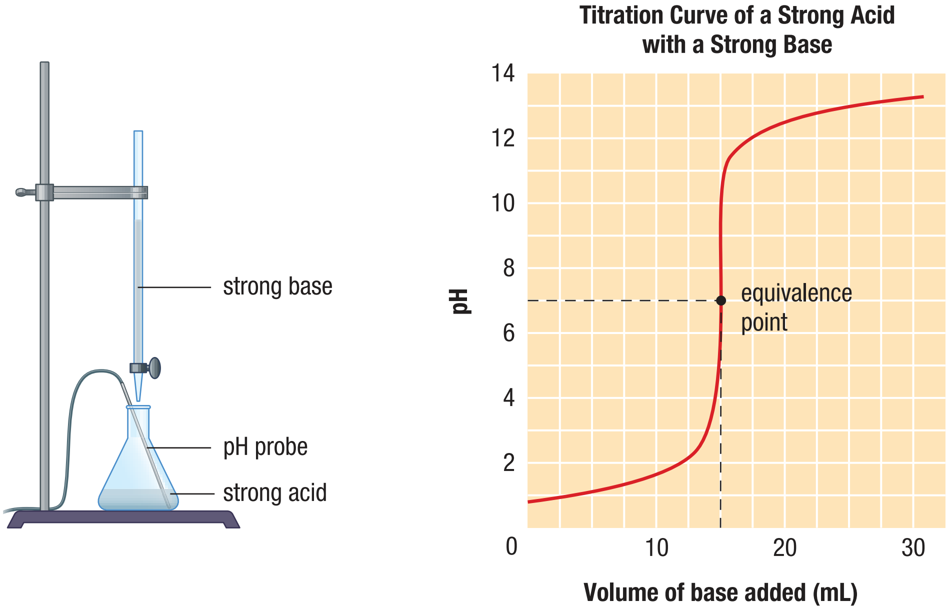


Figure 2 When a strong acid is titrated with a strong base, the pH increases quickly near the equivalence point. In this titration, the equivalence point is at pH 7.

The pH of the solution starts off very low, when there is only acid in the flask. When the base is dripped into the flask the pH increases, slowly at first, as the base neutralizes the acid. Then a rapid increase in pH occurs as the base neutralizes the last traces of the acid. This occurs during the steep portion of the graph that includes the equivalence point.

Acid–base indicators do not change colour at a specific pH value. Rather, we can observe the colour changes over a small pH range. **Table 2** lists some common acid–base indicators and pH ranges over which they change colour. The ideal acid–base indicator should change colour during the steep portion of the titration graph.

Table 2 Common Acid–Base Indicators and the pH Range of Their Colour Changes

Indicator	Acid colour	Base colour	pH range of colour change
thymol blue	red	yellow	1.2 to 2.8
	yellow	blue	8.0 to 9.6
methyl red	red	yellow	4.4 to 6.2
bromothymol blue	yellow	blue	6.0 to 7.6
phenolphthalein	colourless	pink	8.0 to 9.6
alizarin yellow	yellow	red	10.1 to 12.0

LEARNING **TIP**

Indicator Solvents
Some indicators are compounds dissolved in ethanol, which makes them flammable. They should be handled with care and kept away from flames and sparks.

Which of these indicators would be suitable to mark the equivalence point of the titration represented in Figure 2? Methyl red, bromothymol blue, and phenolphthalein are all suitable because their colour changes occur during the steep portion of the curve that includes the equivalence point (**Figure 3**). Thymol blue and alizarin yellow are not suitable because their colour changes do not occur in the steep portion of the graph.

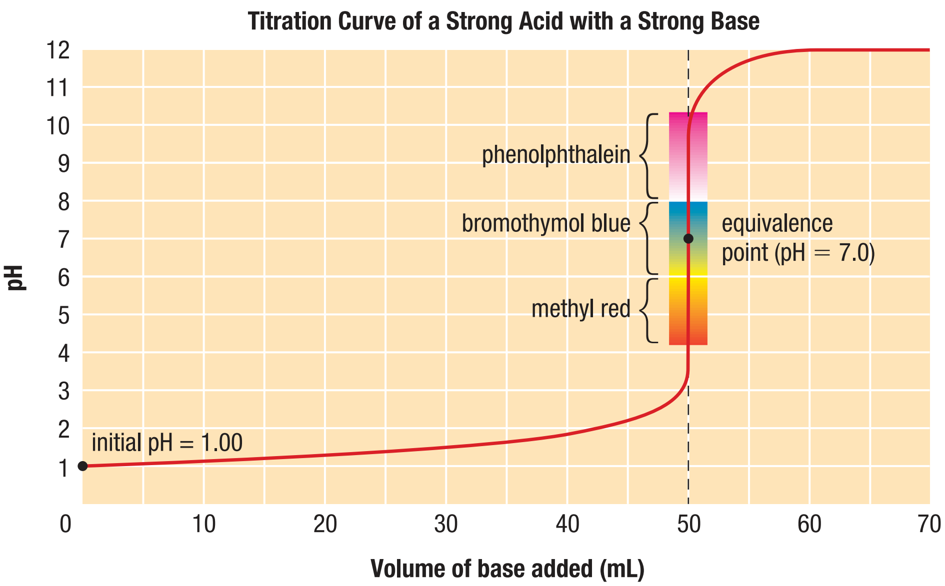


Figure 3 Methyl red, bromothymol blue, and phenolphthalein are appropriate indicators for the titration of a strong acid with a strong base because their colour changes occur along the steepest portion of the pH curve.

UNIT TASK **BOOKMARK**

In the Unit Task (page 498) you will have a chance to apply the skills that you learn in this section and in the investigations. The task involves a titration and a series of stoichiometric calculations. Your aim is to test a water-softening process that you have designed.

Tutorial 1 Performing a Titration

Titration is an analytical procedure used to determine the concentration of a solution. Titration involves using a burette to add a precise volume of one solution (the titrant) to a sample of another solution in an Erlenmeyer flask. Titrant is added until an observable endpoint is reached. The volume of titrant used is then determined by subtracting initial volume from final volume.

The following steps outline how to perform a titration.

Preparing the Burette

SKILLS HANDBOOK A3. 5

1. Assemble the required equipment: burette; burette brush; retort stand and burette clamp; small funnel; beaker; Erlenmeyer flask; meniscus reader; wash bottle of distilled water (**Figure 4**). You will also need the titrant solution and the sample solution. Check that the tip of the burette is not chipped or broken. Replace the tip or the entire burette if necessary. Also check that the valve turns easily and smoothly and that liquid flows easily out of the tip.
2. Close the valve. Rinse the inside of the burette with small volumes of distilled water from the wash bottle. Open the valve to allow the water to drain through the tip (**Figure 5**). If water clings to the inner wall of the burette, carefully scrub the inside of the burette with the brush and rinse again.
3. Close the valve. Clamp the burette to the stand. Pour about 10 mL of the titrant solution into the burette using the funnel (**Figure 6**). You may have to raise the funnel slightly to allow the solution to flow into the burette.
4. Carefully remove the burette from its clamp. Tilt and roll the burette in your fingers so that its inner surface is coated with titrant (**Figure 7**). Allow some titrant to flow out through the tip into a waste beaker. Pour the remaining titrant out of the top of the burette into the beaker. Replace the burette in its clamp and repeat Steps 3 and 4.

Figure 4 Titration equipment

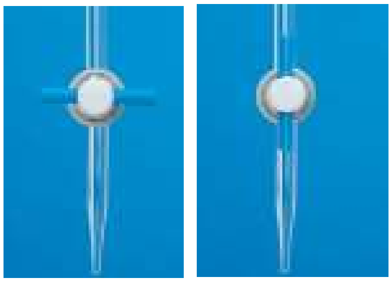


Figure 5 The valve is closed when it is perpendicular to the burette and open when it is parallel to the burette.

Figure 7 Roll the burette in your fingers to wet the inner surface with titrant.

Figure 6 A tight-fitting funnel prevents the titrant from draining into the flask. Raise the funnel slightly if necessary.

5. Using the funnel, fill the burette with titrant until the volume level is about 2 mL higher than the 0.00 mL mark.
6. Place the waste beaker under the burette. Briefly open the valve fully to clear any air bubbles that may have formed in the valve and tip. Then almost close the valve (**Figure 8**, on next page) and allow liquid to run slowly out of the burette until the level is between 0.00 mL and 1.00 mL. Record this reading to the nearest 0.02 mL. Always read a liquid level at the bottom of the meniscus. Read at eye level rather than on an angle (**Figure 9**, on next page). You might find it helpful to use a meniscus reader: a card with a black stripe that makes it easier to see the meniscus.
7. Touch the tip of the burette to the side of the waste beaker, if necessary, to remove any clinging drops. Use the wash bottle of distilled water to rinse the outside of the tip, to ensure that there are no drops of titrant on the outside that might affect your readings.

LEARNING TIP

Titration Precautions

When you are preparing a burette for a titration, follow all the usual safety precautions for handling toxic or corrosive solutions. In addition, take particular care with laboratory glassware. Always wear chemical safety goggles, an apron, and protective gloves.

Figure 8 Air bubbles in the tip and valve must be cleared before starting a titration.

Figure 9 Always take your reading from the bottom of the meniscus.

Figure 10 The burette tip should remain in the flask throughout the titration.

Preparing the Sample

SKILLS HANDBOOK A3. 4

8. You will need a beaker containing the sample solution and a pipette fitted with a pipette pump. Pipette the required volume of sample solution into the Erlenmeyer flask. (Instructions for pipetting are given in the Skills Handbook, A3.4.)
9. Add an appropriate indicator to the flask, if necessary. Use the smallest possible volume of indicator. Since many indicators are weak acids, using too much of the indicator solution will give an inaccurate endpoint.
10. Place a sheet of white paper under the flask to make colour changes easier to see.

Titrating an Unknown Solution

11. Record the initial burette volume.
12. Raise the titration flask or lower the burette so that the burette tip is in the mouth of the flask (**Figure 10**). This will help prevent spillage of titrant.
13. Add titrant to the flask, quickly at first (**Figure 11**). Then, as you near the endpoint (when you see the first signs of a change in colour), add titrant drop by drop, swirling the flask gently between drops.
14. Use a wash bottle to rinse the sides of the flask and the tip of the burette to ensure that all the titrant has reacted (**Figure 12**).
15. Stop titrating when the endpoint is reached: when the addition of a drop of titrant produces a permanent colour change. If you are unsure whether you have reached the endpoint, record the burette volume. Then add another drop of titrant. If a dramatic colour change occurs, you were likely at the endpoint and have now passed it.

Figure 11 Allow the titrant to flow into the flask. Prepare to reduce the flow as you approach the endpoint.

Figure 12 Rinse the tip and sides of the titration flask with distilled water prior to the endpoint.

LEARNING **TIP**

Cleaning Equipment
All laboratory equipment should be cleaned before being stored. To clean burettes and pipettes, first drain them, then rinse them with acetic acid (vinegar), and then with distilled water.

Questions

1. Why is it necessary to keep the volume of indicator used in an acid–base titration to a minimum? K/U
2. Why do you think it is necessary to rinse a burette with titrant after it is rinsed with water? T/I
3. A student forgot to clear an air bubble prior to starting a titration. What effect would this have on the measured volume of titration required to reach the endpoint? T/I
4. Why do you think rinsing the sides of a titration flask near the end of the titration is recommended? T/I
5. A student titrates a hydrochloric acid solution (in the flask) with a sodium hydroxide solution (from the burette) to a phenolphthalein endpoint. The solution in the flask turns pink momentarily when a small volume of sodium hydroxide is added (**Figure 13**). The pink then disappears as the flask is swirled. Explain why these colour changes are observed. T/I

Figure 13 The endpoint has been reached when a single drop of titrant results in a colour change that does not fade when the mixture is swirled.

Standardizing Solutions of Acids and Bases

Titrations require standard solutions. Recall, from Section 8.6, that we know the precise concentration of a standard solution. Standard solutions are sometimes prepared using a precisely measured mass of a solid solute or volume of a concentrated solution. In practice, however, this is not always possible. Some solids are difficult to obtain in a highly pure form. Others change once they are exposed to air. Sodium hydroxide, for example, absorbs water vapour from the air. As a result, the mass of solid sodium hydroxide increases slightly as it is being weighed. Sodium hydroxide solutions also react with carbon dioxide in the air to form carbonate compounds (**Figure 14**). This decreases the sodium hydroxide concentration in the solution.

Because of these problems, chemists must first standardize the solution of sodium hydroxide needed for an analysis. Standardizing a solution involves accurately determining its concentration. This is done by titrating the solution against a primary standard. A **primary standard** is a chemical that is highly pure and chemically stable. It should react completely and almost instantaneously and should ideally have a large molecular mass. Potassium hydrogen phthalate, $KHC_8H_4O_4$ (abbreviated to KHP), is a common primary standard used to standardize basic solutions. During this process, a measured mass of KHP is titrated by the basic solution. Chemists use the amount of KHP present to determine the amount concentration of base in the titrant. The equation for the neutralization of sodium hydroxide with KHP is

$$KHP(aq) + NaOH(aq) \rightarrow H_2O(l) + NaKP(aq)$$

Sodium carbonate is a common primary standard used to standardize acidic solutions. The equation for the neutralization of hydrochloric acid with sodium carbonate is

$$2\ HCl(aq) + Na_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g) + 2\ NaCl(aq)$$

Investigation 10.3.1

Standardization of a Sodium Hydroxide Solution (p. 487)
In this investigation you will learn and practise the techniques of titration as you determine the concentration of a basic solution.

Figure 14 Sodium hydroxide reacts with carbon dioxide to form solid sodium carbonate. If the mouth and glass stopper of this bottle are not cleaned regularly, sodium carbonate can glue the stopper to the bottle, making it almost impossible to open.

primary standard a highly pure and stable chemical used to determine the precise concentration of acids or bases

Tutorial 2 Titration Calculations

Titrations involve several trials to ensure that the data collected are reliable. Data from the three or four most consistent trials are then averaged and used to determine the concentration of the solution being analyzed. Calculations involving titration data follow the same general process as any stoichiometry calculation.

Investigation **10.3.2**

What Is the Amount Concentration of Ethanoic Acid in Vinegar? (p. 488)
Put your skills and understanding to the test to experimentally determine the concentration of a common solution.

Sample Problem 1: Determining Concentration Using Titration Data

Several 10.00 mL samples of sulfuric acid solution of unknown concentration are titrated with a 0.100 mol/L solution of sodium hydroxide. (Note that the burette contains the sodium hydroxide solution.) The endpoint was determined using phenolphthalein indicator. The acceptable observations from the titration are summarized in **Table 3**. Use these data to determine the amount concentration of the acid solution.

Table 3 Titration Data for the Titration of Sulfuric Acid with Sodium Hydroxide Solution

Trial	1	2	3
final burette volume reading (mL)	12.52	24.98	37.62
initial burette volume reading (mL)	00.10	12.52	25.10
volume of base (titrant) added (mL)	12.42	12.46	12.52

Given: volume of acid, $V_{H_2SO_4} = 10.00$ mL; concentration of base, $c_{NaOH} = 0.100$ mol/L; volume of base, V_{NaOH} (see Table 3)

Required: concentration of acid, $c_{H_2SO_4}$

Analysis: $c = \dfrac{n}{V}$

Solution:

Step 1. Calculate the average volume of titrant used. In this example, the titrant is the base.

$$V_{NaOH\,(average)} = \frac{12.42 \text{ mL} + 12.46 \text{ mL} + 12.52 \text{ mL}}{3}$$

$$V_{NaOH\,(average)} = 12.4667 \text{ mL } [2 \text{ extra digits carried}]$$

Step 2. Convert the given volumes to litres.

$$V_{NaOH\,(average)} = 12.47 \cancel{\text{mL}} \times \frac{1 \text{ L}}{1000 \cancel{\text{mL}}}$$

$$V_{NaOH\,(average)} = 0.01247 \text{ L}$$

$$V_{H_2SO_4} = 10.00 \cancel{\text{mL}} \times \frac{1 \text{ L}}{1000 \cancel{\text{mL}}}$$

$$V_{H_2SO_4} = 0.01000 \text{ L}$$

Step 3. Write a balanced equation for the reaction, listing the given value(s), required value(s), molar masses of solids (if any), and amount concentrations of solutions below the entities being considered in the problem.

$2\ NaOH(aq) + H_2SO_4(aq) \rightarrow 2\ H_2O(l) + Na_2SO_4(aq)$
0.01247 L 0.01000 L
0.100 mol/L $c_{H_2SO_4}$

Step 4. Use the concentration equation to determine the amount of the substance whose volume and concentration are given, n_{NaOH}.

$$c_{NaOH} = \frac{n_{NaOH}}{V_{NaOH}}$$

$$n_{NaOH} = c_{NaOH}V_{NaOH}$$

$$= 0.100 \frac{\text{mol}}{\cancel{\text{L}}} \times 0.01247 \cancel{\text{L}}$$

$$n_{NaOH} = 1.247 \times 10^{-3} \text{ mol}$$

Step 5. Use the amount of the substance determined in Step 4 and the mole ratio in the balanced equation to determine the amount of the required entity, $n_{H_2SO_4}$.

$$n_{H_2SO_4} = 1.247 \times 10^{-3}\ \cancel{mol_{NaOH}} \times \frac{1\ mol_{H_2SO_4}}{2\ \cancel{mol_{NaOH}}}$$

$$= 6.233 \times 10^{-4} mol\ [1 \text{ extra digit carried}]$$

Step 6. Use the amount of the required substance (determined in Step 5), the volume of the required substance, and the concentration equation to determine the amount concentration of the required substance, $c_{H_2SO_4}$.

$$c_{H_2SO_4} = \frac{n_{H_2SO_4}}{V_{H_2SO_4}}$$

$$= \frac{6.233 \times 10^{-4}\ mol}{1.000 \times 10^{-2}\ L}$$

$$c_{H_2SO_4} = 6.233 \times 10^{-2}\ \frac{mol}{L}$$

Statement: The amount concentration of the sulfuric acid solution is 6.233×10^{-2} mol/L.

The strategy used in Sample Problem 1 also applies to problems involving a primary standard. The only difference is that the mass of the primary standard is used rather than a volume.

Sample Problem 2: Determining Concentration Using a Primary Standard

A 0.306 g mass of potassium hydrogen phthalate, $KHC_8H_4O_4(s)$ or KHP, is dissolved in 50 mL of water. This solution is titrated with a sodium hydroxide solution of unknown concentration. It is determined that 14.80 mL of the sodium hydroxide solution is needed to titrate the potassium hydrogen phthalate sample to the endpoint at which phenolphthalein indicator changes colour. Calculate the amount concentration of the sodium hydroxide solution.

Given: mass of acid, $m_{KHP} = 0.306$ g; volume of base, $V_{NaOH} = 14.80$ mL

Required: amount concentration of sodium hydroxide solution, c_{NaOH}

Solution:

Step 1. Convert the volume of titrant used to litres.

$$V_{NaOH} = 14.80\ \cancel{mL} \times \frac{1\ L}{1000\ \cancel{mL}}$$

$$V_{NaOH} = 0.01480\ L$$

Step 2. Write a balanced equation for the reaction, listing the given value(s), required value(s), molar masses of solid(s), and amount concentrations below the entities being considered in the problem.

$KHP(aq)$	+	$NaOH(aq) \rightarrow H_2O(l) + NaKP(aq)$
0.306 g		0.01480 L
204.23 g/mol		c_{NaOH}

Step 3. Convert mass of given substance (KHP) into amount of given substance.

$$n_{KHP} = 0.306\ \cancel{g} \times \frac{1\ mol}{204.23\ \cancel{g}}$$

$$n_{KHP} = 1.4983 \times 10^{-3}\ mol\ [2 \text{ extra digits carried}]$$

Step 4. Use the amount of the substance (determined in Step 3) and the mole ratio in the balanced equation to determine the amount of the required entity, n_{NaOH}.

$$n_{NaOH} = 1.4983 \times 10^{-3}\ \cancel{mol_{KHP}} \times \frac{1\ mol_{NaOH}}{1\ \cancel{mol_{KHP}}}$$

$$n_{NaOH} = 1.4983 \times 10^{-3}\ mol\ [2 \text{ extra digits carried}]$$

LEARNING **TIP**

Masses and Moles

The amount of a pure substance in a sample is determined only by its mass and not by the volume of water it is dissolved in. That is why the 50 mL volume of water was not involved in calculating the amount of KHP in this problem.

Step 5. Use the amount of the required substance (determined in Step 4), the volume of the required substance, and the concentration equation to determine the amount concentration of the required substance, c_{NaOH}.

$$c_{NaOH} = \frac{n_{NaOH}}{V_{NaOH}}$$

$$= \frac{1.4983 \times 10^{-3} \text{ mol}}{1.480 \times 10^{-2} \text{ L}}$$

$$c_{NaOH} = 0.101 \text{ mol/L}$$

Statement: The amount concentration of the sodium hydroxide solution is 0.101 mol/L.

Practice

SKILLS HANDBOOK A6.2, A6.5

1. A 0.500 mol/L solution of nitric acid, $HNO_3(aq)$, is used to titrate several 25.00 mL samples of a sodium hydroxide solution, NaOH(aq). The concentration of the sodium hydroxide solution is unknown. The endpoint is detected using bromothymol blue indicator. The acceptable data collected from this titration are summarized in **Table 4**. T/I

Table 4 Titration Data for the Titration of Sodium Hydroxide Solution with Nitric Acid

Trial	1	2	3
final burette volume reading (mL)	8.00	16.12	24.62
initial burette volume reading (mL)	0.00	8.06	16.50

   (a) What is the average volume of nitric acid used? [ans: 8.06 mL]
   (b) Calculate the amount concentration of the sodium hydroxide solution. [ans: 0.161 mol/L]

2. A 1.00 g sample of potassium carbonate is dissolved in 100 mL water. A titration of this sample requires 24.20 mL of nitric acid to reach the endpoint. Calculate the amount concentration of the nitric acid. T/I [ans: 0.598 mol/L]

3. 15.52 mL of 0.100 mol/L hydrochloric acid is required to titrate 25.00 mL of a barium hydroxide solution to a bromothymol blue endpoint. Calculate the amount concentration of the barium hydroxide solution. T/I [ans: 3.10×10^{-2} mol/L]

WEB **LINK**

If you would like more practice solving problems involving titration,

GO TO NELSON SCIENCE

10.3 Summary

- Titrations are used to determine the concentration of a solution using another solution whose precise concentration is known. A measured volume of the titrant is added, from a burette, to a measured volume of the other solution until an endpoint is reached.
- The equivalence point of a titration is reached when the sample in the titration flask is completely neutralized.
- The endpoint of a titration is a visual cue used to indicate when the equivalence point is reached. The endpoint can be a sudden change in colour of an acid–base indicator or a sudden change in pH detected using a pH meter.
- A primary standard is a highly pure and stable chemical used in a titration to determine accurately the concentration of a solution.
- The pH of the solution in the flask changes sharply near the equivalence point.
- Appropriate acid–base indicators for a titration change colour during the steepest portion of the pH curve (near the equivalence point).

10.3 Questions

1. What is the purpose of an acid–base indicator in a titration? K/U
2. Distinguish between equivalence point and endpoint. K/U
3. A technician carefully measures 4.00 g of solid sodium hydroxide and dissolves it in water to prepare 100.0 mL of solution. T/I
   (a) Calculate the predicted concentration of this solution.
   (b) Titrating this solution against a primary standard showed that the sodium hydroxide concentration was actually 0.940 mol/L. Suggest why this concentration differs from the predicted concentration.
4. Why do you think it is necessary to re-standardize a solution after it has been stored for quite some time? K/U T/I
5. Why do you think storage bottles of solutions are often filled right to the top? K/U
6. Potassium hydrogen phthalate samples are generally dried in an oven before being used to standardize a base. Why do you think this step is necessary? What effect might omitting this step have on the final concentration of the base? Why? K/U T/I
7. Describe how a pH meter can be used to determine the equivalence point of a titration. What factor determines whether or not an acid–base indicator is suitable for a titration? K/U
8. Two different acids of the same concentration were titrated with sodium hydroxide solution. The pH changes occurring during these titrations were measured and plotted on the same graph (**Figure 15**). T/I

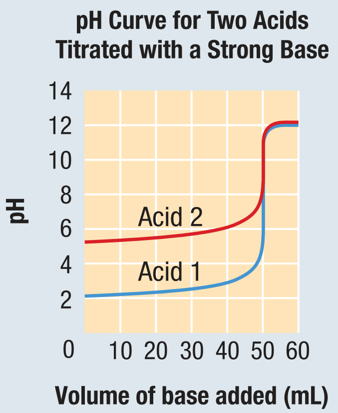


Figure 15

   (a) Suggest an explanation to account for the different initial pH of the acids.
   (b) What volume of base is required to neutralize these acids? Why is this volume the same for both acids?
   (c) Select an acid–base indicator from Table 2 that could be used for the titration of one of the acids but not for the other.
9. Calcium carbonate is the active ingredient in many antacid medications. An antacid tablet neutralizes 25.20 mL of 0.50 mol/L hydrochloric acid. T/I C
   (a) Write the chemical equation for the neutralization reaction.
   (b) What is the mass of calcium carbonate in the tablet? (Assume that calcium carbonate is the only ingredient that reacts with acid.)
10. An ammonium hydroxide solution is standardized by titrating 10.00 mL samples with a 0.25 mol/L solution of sulfuric acid. The titration data, indicating the volumes of acid added from the burette, are recorded in **Table 5**. T/I C

Table 5 Titration Data for the Titration of Ammonium Hydroxide Solution with 0.25 mol/L Sulfuric Acid

Trial	1	2	3
final burette volume reading (mL)	8.46	17.00	25.66
initial burette volume reading (mL)	0.00	8.50	17.00

   (a) Write the balanced chemical equation for the neutralization reaction.
   (b) Calculate the average volume of acid used.
   (c) Determine the amount concentration of the ammonium hydroxide solution.
   (d) Why do you think it is necessary to conduct more than one trial when standardizing a solution?
11. The chemical equation for the neutralization of lactic acid, $HC_3H_5O_2$, with sodium hydroxide is

$$HC_3H_5O_2(aq) + NaOH(aq) \rightarrow H_2O(l) + NaC_3H_5O_2(aq)$$

During a titration analysis for lactic acid in milk, several 10.00 mL milk samples were titrated using a 0.100 mol/L sodium hydroxide solution. Phenolphthalein indicator was used to indicate the endpoint. On average, 1.60 mL of the sodium hydroxide solution was required for each titration. Calculate the lactic acid concentration in the milk. T/I
12. Potassium hydrogen iodate, $KH(IO_3)_2$, is a primary standard sometimes used to standardize base solutions. For each of several trials, 1.50 g of $KH(IO_3)_2$ was first dissolved in water in the titration flask. An average of 25.42 mL of a potassium hydroxide solution was required to reach the endpoint. T/I C
   (a) Write a balanced equation for the neutralization of a solution of potassium hydroxide by a solution of potassium hydrogen iodate.
   (b) Determine the amount concentration of the potassium hydroxide solution.
13. Potassium carbonate, K_2CO_3, is a primary standard used to standardize a hydrochloric acid solution. 1.00 g samples of solid potassium carbonate are first dissolved in water to make 100 mL of solution. Over several trials, an average of 24.20 mL of hydrochloric acid is required to titrate the potassium carbonate solution to a phenolphthalein endpoint. Determine the amount concentration of the hydrochloric acid solution. T/I

CHAPTER 10 Investigations

Investigation 10.2.1 CONTROLLED EXPERIMENT

SKILLS MENU

- Questioning
- Researching
- Hypothesizing
- Predicting
- Planning
- Controlling Variables
- Performing
- Observing
- Analyzing
- Evaluating
- Communicating

Not All Acids Are Created Equal

In this experiment, you will compare three acids. You will investigate each acid's reactivity with magnesium metal, each acid's electrical conductivity, and the volume of base required to neutralize each acid.

Testable Question

What is the order of reactivity of the three acids: ethanoic acid, $HC_2H_3O_2(aq)$, hydrochloric acid, $HCl(aq)$, and sulfuric acid, $H_2SO_4(aq)$?

Hypothesis

SKILLS HANDBOOK A2.2

Use your knowledge of strong and weak acids to write a hypothesis ranking the three acids in order of reactivity.

Variables

Read the Procedure and identify the manipulated and the responding variables in the experiment. Identify which variable(s) must be controlled during the experiment.

Experimental Design

Write a brief outline of the experiment, stating how the variables will be manipulated, controlled, or measured.

Equipment and Materials

SKILLS HANDBOOK A1

- chemical safety goggles
- lab apron
- three test tubes
- test-tube rack
- timing device
- 6-well well plate
- conductivity tester
- three 1 cm strips of magnesium, $Mg(s)$
- dropper bottles of
  - 1.0 mol/L ethanoic acid, $HC_2H_3O_2(aq)$
  - 1.0 mol/L hydrochloric acid, $HCl(aq)$
  - 1.0 mol/L sulfuric acid, $H_2SO_4(aq)$
  - 1.0 mol/L sodium hydroxide, $NaOH(aq)$
  - phenolphthalein indicator solution

The acids and bases used in this investigation are corrosive. Any spills on the skin, in the eyes, or on clothing should be washed immediately with plenty of cool water. Report any spills to your teacher.

Phenolphthalein indicator solution is flammable. Keep it away from open flames.

Procedure

SKILLS HANDBOOK A3.9

1. Put on your chemical safety goggles and lab apron.
2. Place three test tubes in the test-tube rack. Label one "Ethanoic acid," one "Hydrochloric acid," and one "Sulfuric acid."
3. Add ethanoic acid to the first test tube, hydrochloric acid to the second, and sulfuric acid to the third, each to a depth of about 2 cm.
4. Add a 1 cm strip of magnesium to each test tube. Compare the length of time required for each acid to completely consume the magnesium. Record your observations.
5. Add 10 drops of ethanoic acid, 10 drops of hydrochloric acid, and 10 drops of sulfuric acid to three wells of the well plate.
6. Test the conductivity of each solution. Rinse and dry the conductivity tester after each test.
7. Add 1 drop of phenolphthalein indicator to each well.
8. Add sodium hydroxide solution, drop by drop, to the ethanoic acid well. Count the number of drops to produce the first faint pink colour that lasts for about 20 s.
9. Repeat Step 8, adding sodium hydroxide to the other acids.
10. Dispose of the waste materials as directed by your teacher.

Analyze and Evaluate

(a) Write ionization equations for the three acids. K/U C

(b) Write chemical equations for the reactions of the acids with magnesium. K/U C

(c) What variables were measured/recorded and/or manipulated in this experiment? T/I

(d) Rank the three acids according to their reactivity with magnesium. T/I

(e) Rank the three acids according to their electrical conductivity. Suggest an explanation for any difference in conductivity observed. K/U T/I

(f) Write a balanced chemical equation for the neutralization of each acid with sodium hydroxide.

(g) Rank the three acids according to the volume of base required to neutralize the acids. Account for any major differences observed. K/U T/I

(h) Does your answer to the testable question support your hypothesis? Explain. T/I

(i) How could you improve the design of this experiment? T/I

Apply and Extend

(j) One of the products of a neutralization reaction is an ionic compound. Experimental evidence suggests that the anion in this compound may react with water to produce a basic solution. (Recall Section 5.3, Elements and Their Oxides, and Figure 6 in Section 10.2.)

(i) Two of the ionic compounds produced in this investigation are sodium chloride, NaCl, and sodium ethanoate, $NaC_2H_3O_2$. Design an investigation to compare the pH of 0.1 mol/L solutions of sodium chloride and sodium ethanoate.

(ii) Conduct the investigation once you have your teacher's approval. Which anion produces the most basic solution?

(iii) Write a chemical equation to explain how the anion in (ii) raises the pH of water.

Investigation 10.3.1 OBSERVATIONAL STUDY

SKILLS MENU

• Questioning	• Planning	• Observing
• Researching	• Controlling Variables	• Analyzing
• Hypothesizing	• Performing	• Evaluating
• Predicting		• Communicating

Standardization of a Sodium Hydroxide Solution

Sodium hydroxide solutions are commonly used in analytical chemistry procedures. Sodium hydroxide is a white ionic solid that readily absorbs water and reacts with carbon dioxide from the atmosphere. These properties make it difficult to prepare sodium hydroxide solutions of a known concentration. In practice, sodium hydroxide solutions of approximate concentrations are first prepared and then titrated against a primary standard to determine their concentrations accurately.

In this investigation, you will use potassium hydrogen phthalate (abbreviated KHP) as a primary standard to determine the concentration of a sodium hydroxide solution. The chemical equation for the neutralization of KHP by sodium hydroxide is

$$\underset{\text{acid}}{KHP(aq)} + \underset{\text{base}}{NaOH(aq)} \rightarrow \underset{\text{water}}{H_2O(l)} + \underset{\text{ionic compound}}{KNaP(aq)}$$

The acid–base indicator used in this titration is phenolphthalein. The endpoint of the titration is the first faint pink colour that lasts for about 20 s. The sodium hydroxide solution that you standardize in this investigation will be used in the next investigation.

Purpose

To use a primary standard to determine the amount concentration of a sodium hydroxide solution

Equipment and Materials

SKILLS HANDBOOK A1

- chemical safety goggles
- lab apron
- scoopula
- two 250 mL Erlenmeyer flasks
- balance
- stirring rod
- 150 mL beaker
- 100 mL graduated cylinder
- 50 mL burette
- retort stand with burette clamp
- small funnel
- meniscus finder
- small beaker containing potassium hydrogen phthalate, $KHC_8H_4O_4(s)$
- wash bottle of distilled water
- dropper bottle of phenolphthalein indicator
- sodium hydroxide solution to be standardized (approximately 0.1 mol/L NaOH(aq))

Phenolphthalein is flammable. Keep it away from open flames.

Sodium hydroxide is corrosive. Avoid skin or eye contact. Should sodium hydroxide come in contact with your skin or eyes, wash the affected area for 15 minutes with cool water and notify your teacher.

Procedure

SKILLS HANDBOOK A2.4, A3.2, A3.5

1. Put on your chemical safety goggles and lab apron.
2. Using the balance, weigh out a sample of potassium hydrogen phthalate that has a mass of approximately 0.5 g. Record this mass to the nearest 0.01 g.
3. Transfer the potassium hydrogen phthalate to a flask.
4. Dissolve the potassium hydrogen phthalate in about 50 mL distilled water.
5. Add 5 drops of phenolphthalein indicator.
6. Obtain 60 mL of the sodium hydroxide solution in a clean, dry 150 mL beaker.
7. Fill the burette with the sodium hydroxide solution. Be sure to prepare the burette correctly and clear any air bubbles from the valve.
8. Record the initial volume level of sodium hydroxide in the burette to the nearest 0.02 mL.
9. Titrate the potassium hydrogen phthalate solution to the phenolphthalein endpoint (**Figure 1**).

Figure 1 Swirl the flask frequently as the titration nears the endpoint.

10. Record the final volume of sodium hydroxide to the nearest 0.02 mL.
11. Discard the waste solutions as directed by your teacher.
12. Repeat the titration until the results of three trials are within 0.2 mL. Your teacher will keep the remaining sodium hydroxide for the next investigation.
13. Wash your hands thoroughly.

Analyze and Evaluate

SKILLS HANDBOOK A6

(a) Determine the amount of potassium hydrogen phthalate used in each of your three best trials. T/I

(b) Calculate the amount concentration of the sodium hydroxide solution in each trial. T/I

(c) Use your three best trials to calculate the average concentration of the sodium hydroxide solution. T/I

(d) Evaluate the titration data collected. How confident are you in your determination of the amount concentration of sodium hydroxide? Do you think that your value is accurate? Explain. T/I

(e) Did you have to use a precisely measured volume of water to dissolve the KHP? Explain. T/I

(f) Usually, in the analysis of titration data, you average the volume of titrant used, then use that average volume to determine the amounts of reactants. In this investigation, however, you first calculated the concentration of sodium hydroxide for each trial, determined an average concentration, then used that average concentration to determine the amounts of reactants. Why do you think this approach was suggested? T/I

Apply and Extend

(g) A student added 34.50 mL of sodium hydroxide solution to the flask before realizing that he had forgotten to add the indicator. The solution in the flask turned bright pink as soon as the indicator was added. Suggest a lab procedure that he could use to salvage this titration. T/I

Investigation 10.3.2 OBSERVATIONAL STUDY

SKILLS MENU

• Questioning	• Planning	• Observing
• Researching	• Controlling Variables	• Analyzing
• Hypothesizing	• Performing	• Evaluating
• Predicting		• Communicating

What Is the Amount Concentration of Ethanoic Acid in Vinegar?

Vinegar is a dilute solution of ethanoic acid (acetic acid) in water. Vinegar manufacturers must ensure their product meets the acetic acid concentration claimed on the product label. One way of checking the concentration is by titrating samples of vinegar from each batch with a standardized solution of sodium hydroxide. A typical bottle of vinegar reports a concentration of 5.0 % V/V, which corresponds to an amount concentration of 0.87 mol/L.

In this investigation, you will determine the amount concentration of vinegar experimentally by titrating it with standardized sodium hydroxide solution. You will use the solution that you standardized in Investigation 10.3.1. Phenolphthalein is the indicator used in this titration. The phenolphthalein endpoint occurs when a drop of base produces a faint pink colour that lasts for about 20 s.

Purpose

To determine the amount concentration of ethanoic acid in vinegar

Equipment and Materials

SKILLS HANDBOOK A1

- chemical safety goggles
- lab apron
- 150 mL beaker
- 5 mL volumetric pipette
- pipette pump
- 100 mL graduated cylinder
- 50 mL burette
- retort stand with burette clamp
- small funnel
- two 250 mL Erlenmeyer flasks
- meniscus finder
- vinegar
- wash bottle of distilled water
- dropper bottle of phenolphthalein indicator
- standardized sodium hydroxide solution, NaOH(aq)

Phenolphthalein indicator is flammable. Keep it away from open flames.

Sodium hydroxide is corrosive. Avoid skin or eye contact. Should sodium hydroxide come in contact with your skin or eyes, wash the affected area for 15 minutes with cool water and notify your teacher.

Procedure

SKILLS HANDBOOK A3.3, A3.4, A3.5

1. Put on your chemical safety goggles and lab apron.
2. Obtain about 25 mL of vinegar in a clean, dry beaker.
3. Pipette a 3.0 mL sample of vinegar into a clean Erlenmeyer flask.
4. Add about 40 mL of water to the flask.
5. Add 3 to 5 drops of phenolphthalein indicator to the flask.
6. Rinse the burette with sodium hydroxide solution. Fill the burette with the sodium hydroxide solution and clear any air bubbles from the valve.
7. Record the initial volume of sodium hydroxide in the burette to the nearest 0.02 mL.
8. Titrate the vinegar sample to the phenolphthalein endpoint (**Figure 1**).
9. Record the final volume of sodium hydroxide in the burette to the nearest 0.02 mL.
10. Discard the waste solution as directed by your teacher.
11. Repeat the titration until you have three trials with results that are within 0.2 mL.
12. Wash your hands thoroughly.

Figure 1

Analyze and Evaluate

SKILLS HANDBOOK A6

(a) Use your three best trials to calculate the average concentration of ethanoic acid in vinegar. T/I

(b) What effect, if any, would the errors below have on the calculated concentration of ethanoic acid? Justify your answer in each case. T/I

(i) forgetting to initially rinse a wet burette with sodium hydroxide

(ii) not clearing a large air bubble in the tip of the burette

(iii) adding 10 mL of rinse water to the flask

(c) Why did the addition of water not affect the determination of the concentration of ethanoic acid in vinegar? T/I

Apply and Extend

(d) Repeat the titration using a pH meter rather than an acid–base indicator to monitor the progress of neutralization. T/I

(e) Prepare a graph of pH against volume of base used. T/I C

(f) Compare this graph with the graph obtained titrating a strong acid with a strong base (Figure 3 in Section 10.3). T/I

CHAPTER 10 SUMMARY

Summary Questions

1. Create a study guide based on the points listed in the margin on page 462. For each point, create three or four sub-points that provide further information, relevant examples, explanatory diagrams, or general equations.
2. Look back at the Starting Points questions on page 462. Answer these questions using what you have learned in this chapter. Compare your latest answers with those that you wrote at the beginning of the chapter. Note how your answers have changed.
3. (a) What are the most significant changes to your understanding of acids and bases as a result of completing this chapter?
   (b) What aspects of acid–base chemistry remain unclear?

Vocabulary

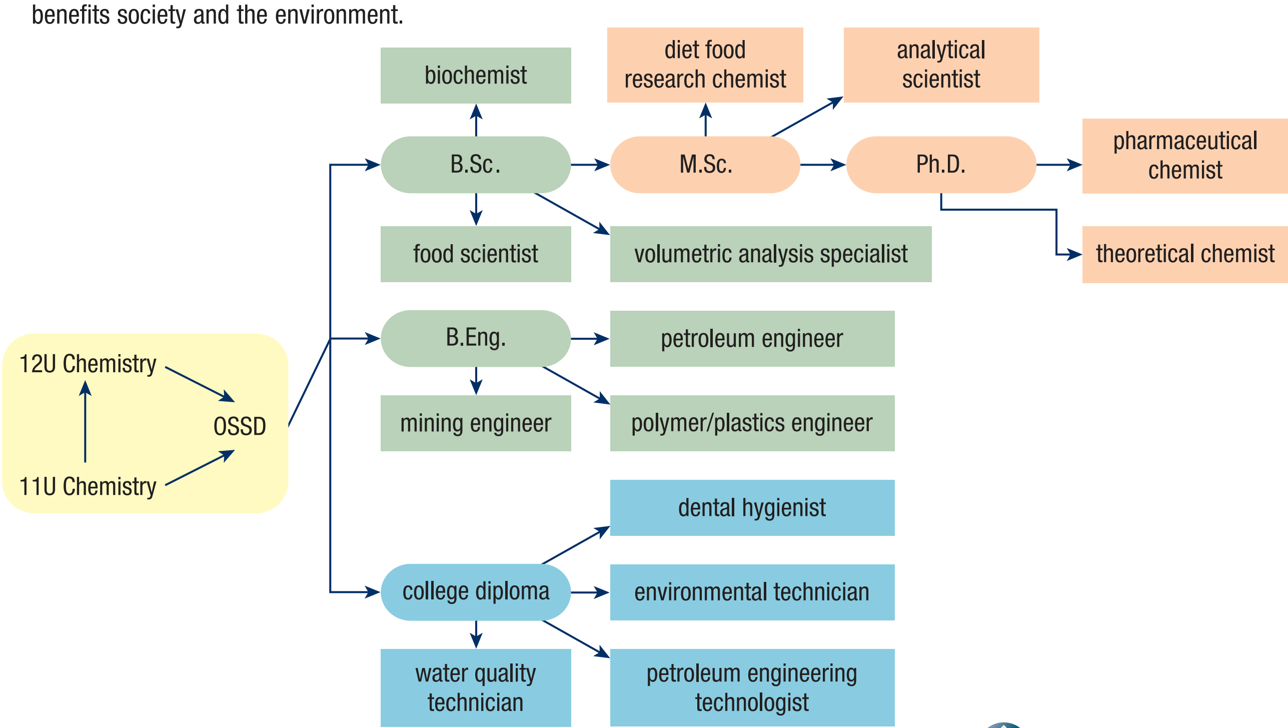
oxyacid (p. 466)
ionization (p. 471)
strong acid (p. 472)
weak acid (p. 472)
titration (p. 476)
titrant (p. 476)
burette (p. 476)
equivalence point (p. 477)
endpoint (p. 477)
primary standard (p. 481)

CAREER PATHWAYS

SKILLS HANDBOOK A7

Grade 11 Chemistry can lead to a wide range of careers. Some require a college diploma or a B.Sc. degree. Others require specialized or post-graduate degrees. This graphic organizer shows a few pathways to careers mentioned in this chapter.

1. Select two careers, related to acids and bases, that you find interesting. Research the educational pathways that you would need to follow to pursue these careers. What is involved in the required educational programs? Prepare a brief report of your findings.
2. For one of the two careers that you chose above, describe the career, main duties and responsibilities, working conditions, and setting. Also outline how the career benefits society and the environment.



GO TO NELSON SCIENCE

CHAPTER 10 SELF-QUIZ

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

For each question, select the best answer from the four alternatives.

1. Which of the following is a property of a base? (10.1) K/U
   (a) tastes sour
   (b) feels slippery
   (c) demineralizes calcium compounds
   (d) has a pH less than 7
2. Which of the following gases forms when an acid reacts with a metal? (10.1) K/U
   (a) carbon dioxide
   (b) oxygen
   (c) hydrogen
   (d) It depends on which acid and metal are used.
3. Which of the following acids contains no oxygen? (10.1) K/U
   (a) chloric acid
   (b) chlorous acid
   (c) hypochloric acid
   (d) hydrochloric acid
4. According to Arrhenius, which of the following ions are produced when an acid is added to water? (10.2) K/U
   (a) hydrogen ions
   (b) hydroxide ions
   (c) metal ions
   (d) halide ions
5. The strength of an acid depends upon which of the following? (10.2) K/U
   (a) the degree of dissociation of acid molecules
   (b) the degree of ionization of acid molecules
   (c) how well the acid conducts an electric current
   (d) whether the acid is contained in foods
6. Stomach fluid has a lower pH than milk. What does this statement tell us about these two liquids? (10.2) K/U
   (a) Both milk and stomach fluids are acids.
   (b) Both milk and stomach fluids are bases.
   (c) Stomach fluid is more acidic than milk.
   (d) Milk is more acidic than stomach fluid.
7. When is a titration most likely to give accurate results? (10.3) K/U
   (a) when its endpoint occurs before its equivalence point
   (b) when its endpoint occurs after its equivalence point
   (c) when its endpoint coincides with its equivalence point
   (d) when it has an endpoint but no equivalence point
8. Which of the following terms describes a sudden change in an observable property that occurs during a titration? (10.3) K/U
   (a) endpoint
   (b) overshot reading
   (c) equivalence point
   (d) stoichiometry
9. A 20 mL volume of a hydrochloric acid solution, $HCl(aq)$, was placed into a beaker containing an appropriate acid–base indicator. It was titrated with a solution of sodium hydroxide, $NaOH(aq)$, until the endpoint was reached. If 25 mL of sodium hydroxide solution with a concentration of 0.20 mol/L was used, what is the concentration of the acid? (10.3) K/U
   (a) 0.20 mol/L
   (b) 0.25 mol/L
   (c) 2.0 mol/L
   (d) 2.5 mol/L

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

10. A common cause of erosion of tooth enamel and sensitive teeth is the ingestion of acidic foods and beverages. (10.1) K/U
11. Carbon dioxide reacts with acids to form carbonate compounds. (10.1) K/U
12. The name of an oxyacid is based on the name of the related oxyanion whose name ends in *-ite*. (10.1) K/U
13. The separation of individual ions from an ionic compound as it dissolves in water is called ionization. (10.2) K/U
14. Neutralization reactions produce water. (10.2) K/U
15. The pH of a strong acid will be higher than the pH of a weak acid of the same concentration. (10.2) K/U
16. A primary standard is a chemical that is highly pure and chemically unstable. (10.3) K/U
17. To find the volume of standard solution used in a titration, subtract the initial reading on the burette from the final reading on the burette. (10.3) K/U
18. Volume readings in a burette should be taken at the bottom of the meniscus. (10.3) K/U

To do an online self-quiz,
GO TO NELSON SCIENCE

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

Knowledge

For each question, select the best answer from the four alternatives.

1. Which of the following is a property of an acid? (10.1) K/U
   (a) It has a pH greater than 7.
   (b) It tastes bitter.
   (c) It turns phenolphthalein pink.
   (d) It reacts with carbonates.
2. A student uses a drinking straw to gently blow exhaled air through a solution. The solution turns cloudy as the result of the formation of a precipitate. Which of the following types of compounds is most likely present in the initial solution? (10.1) K/U
   (a) a carbonate
   (b) an acid
   (c) a base
   (d) a salt
3. Which of the following acids has a name that begins with the prefix *hydro-*? (10.1) K/U
   (a) $HNO_2(aq)$
   (b) $H_2Te(aq)$
   (c) $H_3PO_4(aq)$
   (d) $HClO_2(aq)$
4. According to Arrhenius, which of the following ions are released when a base is added to water? (10.2) K/U
   (a) hydrogen ions
   (b) hydroxide ions
   (c) metal ions
   (d) halide ions
5. Arrhenius based his theories on acids and bases on which of the following concepts? (10.2) K/U
   (a) Some acids and bases are stronger than other acids and bases.
   (b) Acids and bases react with each other in a neutralization reaction.
   (c) Water is always a product in a neutralization reaction.
   (d) Electrolytes in a solution conduct an electric current because of the ions present.
6. Tomatoes are a popular field crop in parts of Canada. Tomatoes have a pH of approximately 4. How would you classify a tomato? (10.2) K/U
   (a) neutral
   (b) base
   (c) acid
   (d) strong
7. What is the purpose of performing a titration? (10.3) K/U
   (a) to determine the concentration of a standard solution
   (b) to determine the concentration of a solution of unknown concentration
   (c) to measure the volume of the titrant
   (d) to show changes in the colours of the indicator used in the titration
8. During a titration, an endpoint was reached after adding 20 mL of sodium hydroxide solution, $NaOH(aq)$, to 20 mL of sulfuric acid, H_2SO_4. The sodium hydroxide solution had a concentration of 1.50 mol/L. What was the initial concentration of the sulfuric acid? (10.3) K/U
   (a) 0.75 mol/L
   (b) 1.50 mol/L
   (c) 2.25 mol/L
   (d) 3 mol/L
9. A titration was performed. The initial burette reading was 13.64 mL and the final reading was 35.62 mL. What volume of titrant was used during the titration? (10.3) K/U
   (a) 21.98 mL
   (b) 22.02 mL
   (c) 48.26 mL
   (d) 49.26 mL

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

10. Bases react with carbonates to produce carbon dioxide gas. (10.1) K/U
11. The correct name of $HClO_2(aq)$ is chloric acid. (10.1) K/U
12. The correct formula for hydroiodic acid is $HI(aq)$. (10.1) K/U
13. In a strong acid, only a small percentage of molecules ionize in solution. (10.2) K/U
14. A solution with a pH of 3 is 10 times more acidic than a solution with a pH of 4. (10.2) K/U
15. Acidic solutions contain many hydroxide ions. (10.2) K/U
16. A titrant is the solution in the burette during a titration. (10.3) K/U
17. The equivalence point of a titration depends on the choice of indicator. (10.3) K/U
18. In an “overshot” titration, too much titrant was added. (10.3) K/U

Match each term on the left with the most appropriate description on the right.

19. (a) pH of 1 — (i) strong acid
    (b) pH of 4 — (ii) strong base
    (c) pH of 7 — (iii) weak acid
    (d) pH of 8 — (iv) weak base
    (e) pH of 12 — (v) neutral (10.2) K/U

Write a short answer to each question.

20. A metal sample was added to hydrochloric acid. Hydrogen bubbles appeared on the surface of the metal. What do you know about the relative activity of the metal compared to the activity of hydrogen? (10.1) K/U
21. How does an oxyacid differ from an acid such as hydrochloric acid? (10.1) K/U
22. The word "acid" comes from a word that describes a property of acids. Explain this statement. (10.1) K/U
23. How is a base related to the process of saponification? (10.1) K/U
24. What gas is produced when each of the following reactants is dropped into an acid? (10.1) K/U
    (a) calcium carbonate, $CaCO_3(s)$
    (b) magnesium, $Mg(s)$
25. Why does a base feel slippery? (10.1) K/U
26. Calcium chloride, $CaCl_2$, is produced in large quantities in Ontario and Alberta. It is used primarily as road salt and as a road dust-control agent. Describe what happens when calcium chloride dissolves in water. What is this process called? (10.2) K/U
27. What is the difference between a strong acid such as hydrochloric acid, $HCl(aq)$, and a weak acid such as acetic acid, $HC_2H_3O_2(aq)$? (10.2) K/U
28. Both acids and bases form ions in solution. How does the process of ion formation differ between acids and bases? (10.2) K/U
29. Suppose you have one test tube containing a strong acid and another test tube containing a weak acid. You do not know which acid is in which test tube. How might you use two iron nails to determine which solution is the strong acid and which one is the weak acid? (10.2) K/U
30. Weak bases are relatively poor conductors of electricity. Explain why this is so. (10.2) K/U
31. Tartaric acid, $H_2C_4H_4O_6(s)$, is commonly mixed with sodium hydrogen carbonate, $NaHCO_3(s)$, in baking powder. When baking powder is wet, these two compounds react to produce carbon dioxide—which makes baked goods rise. Do you think tartaric acid is a strong acid or a weak acid? Explain your answer. (10.2) K/U
32. Why might several trials be needed to obtain accurate titration results? (10.3) K/U
33. Why is it necessary to write a balanced chemical equation for the reaction that takes place during a titration before any calculations are done? (10.3) K/U
34. Compare and contrast the equivalence point of a titration and its endpoint. (10.3) K/U
35. What is the purpose of titrating a solution against a primary standard? (10.3) K/U
36. Describe how to determine the volume of titrant that was used during a titration. (10.3) K/U
37. When performing a titration, why do you think it is necessary to rinse the burette with both distilled water and the titrant? (10.3) K/U

Understanding

38. (a) What does the term "alkali" mean?
    (b) How does this term relate to parts of the periodic table?
    (c) How does this term relate elements on the periodic table to some of their chemical properties? (10.1) K/U T/I
39. Name the following acids: (10.1) K/U
    (a) $HBr(aq)$
    (b) $H_3PO_3(aq)$
    (c) $HNO_3(aq)$
    (d) $HClO_4(aq)$
40. Name the following bases: (10.1) K/U
    (a) $Ba(OH)_2(aq)$
    (b) $LiOH(aq)$
    (c) $Sr(OH)_2(aq)$
    (d) $Be(OH)_2(aq)$
41. Write a formula for each of the following acids: (10.1) K/U
    (a) phosphoric acid
    (b) nitrous acid
    (c) hydrosulfuric acid
    (d) hydrotelluric acid
42. Write a formula for each of the following bases: (10.1) K/U
    (a) calcium hydroxide
    (b) ammonium hydroxide
    (c) sodium hydroxide
    (d) aluminum hydroxide

43. Use the activity series table in **Figure 1** to determine whether or not each of the following metals will react with hydrochloric acid. If it will react, write a balanced chemical equation for the reaction. If it will not react, write NR, for "no reaction." (10.1) K/U C

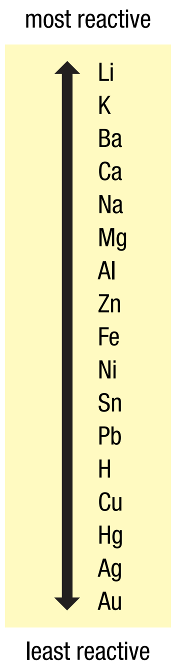


Figure 1

(a) copper
(b) magnesium
(c) potassium
(d) silver
(e) calcium
(f) aluminum

44. HA(aq) is the general formula of an acid. Explain what each part of the general formula means, and why this general formula is effective for representing an acid. (10.1) K/U T/I

45. Pop contains both carbonic acid and phosphoric acid. Write the chemical formulas for these two acids. (10.1) K/U

46. Antacids are over-the-counter medications that relieve discomfort caused by excess stomach acid. Common antacids include compounds such as sodium hydrogen carbonate, $NaHCO_3(s)$, magnesium hydroxide, $Mg(OH)_2(s)$, and calcium carbonate, $CaCO_3(s)$. (10.1, 10.2) K/U C A
(a) What kinds of compounds are these compounds? Why did you draw this conclusion?
(b) Assuming that hydrochloric acid is the main acid in stomach fluids, write a balanced chemical equation for the reaction of each of the compounds listed above when it reaches the stomach.

47. Refer to Figure 6 in Section 10.2 to answer these questions, explaining your reasoning: (10.2) K/U T/I
(a) How much more acidic is battery acid than tomatoes?
(b) How much more basic is milk of magnesia than tomatoes?

48. Seawater is slightly basic: around pH 8. A sample of seawater is diluted with pure water. What effect will this dilution have on the pH of the sample? Explain your answer. (10.2) K/U T/I

49. **Table 1** shows burette readings from three trials in a titration. (10.3) K/U T/I

Table 1 Titration Data

Trial	1	2	3
final burette reading (mL)	36.87	36.06	35.51
initial burette reading (mL)	12.24	11.46	10.59

(a) Determine the volume of titrant used in each of the three trials.
(b) Determine the average volume of titrant used.

50. A student performed three trials of a titration, each time using a 20.00 mL sample of hydrochloric acid, HCl(aq). The titrant was sodium hydroxide solution, NaOH(aq), with a concentration of 0.10 mol/L. The average volume of sodium hydroxide used was 18.54 mL. What is the concentration of the hydrochloric acid? (10.3) K/U T/I

51. A group of students performed a titration of 15.0 mL samples of sulfuric acid, $H_2SO_4(aq)$, with a solution of potassium hydroxide, KOH(aq). Three trials were run. Phenolphthalein was used as an indicator. The concentration of the potassium hydroxide solution was 0.10 mol/L. **Table 2** shows the titration data for three trials. Use these data to answer the following questions. (10.3) K/U T/I C

Table 2 Titration Data

Trial	1	2	3
final burette volume reading (mL)	24.83	16.35	28.95
initial burette volume reading (mL)	8.54	0.03	12.64
volume of base added (mL)			

(a) Copy Table 2 into your notebook. Complete the table with the volume of base added for each trial.
(b) Calculate the average volume of base added.
(c) Write a balanced chemical equation for the reaction that takes place.
(d) Determine the concentration of the sulfuric acid.

52. An 18 mL sample of hydrochloric acid, $HCl(aq)$, in a flask was titrated with a primary standard solution of sodium carbonate, $Na_2CO_3(aq)$. Methyl red was used as an indicator. The primary standard solution was prepared by dissolving 0.53 g of sodium carbonate in enough water to make 100 mL of solution. In a single trial of the titration, the initial volume reading on the burette was 0.21 mL and the final volume reading was 26.23 mL. (10.3) K/U T/I C A
    (a) What volume of primary standard solution was used in this trial?
    (b) What amount of sodium carbonate reacted with the acid, during this trial?
    (c) What was the concentration of the hydrochloric acid solution?

Analysis and Application

53. Certain batteries are known as alkaline batteries. (10.1) K/U T/I A
    (a) What type of compound do you expect to be present in an alkaline battery?
    (b) What specific compound do these batteries contain?
    (c) What does the purpose of an alkaline battery reveal about the strength of this alkaline compound?
54. Suppose that you have hard-water deposits of magnesium carbonate, $MgCO_3(s)$, in your sink. You have one cleaner that contains vinegar, which is dilute ethanoic acid, $HC_2H_3O_2(aq)$, and another cleaner that contains ammonium hydroxide, $NH_4OH(aq)$. Which cleaner do you think would be more effective in removing the hard-water deposit? Explain your answer. (10.1) K/U T/I A
55. Although all rain is slightly acidic, pollutants in rain can make it more acidic. Acid precipitation is a problem in regions downwind of heavy industries and busy streets and highways. The greatest source of acid-causing gases in Canada is refineries that smelt nickel and copper. Suppose the following gases are released into the atmosphere. Write the chemical equation for the reaction that occurs when each of the following gases reacts with water in the air to produce an acid. Name each acid produced. (10.1) K/U T/I C A
    (a) NO_2, nitrogen dioxide (two different acids form from the same reaction)
    (b) SO_2, sulfur dioxide
    (c) SO_3, sulfur trioxide
56. A student correctly names $H_2SO_3(aq)$ as sulfurous acid. She then names $HNO_3(aq)$ as nitrous acid. Is nitrous acid the correct name for the second acid? If not, what is the correct name of $HNO_3(aq)$? Justify your answer(s). (10.1) K/U T/I A
57. Large factories in Alberta and Ontario manufacture nitrogen-based fertilizers for use throughout Canada. Two of the primary compounds produced are nitric acid, $HNO_3(aq)$, and ammonium nitrate, $NH_4NO_3(aq)$. How might one of these compounds be used to make the other compound? Include a chemical equation in your answer. (10.2) K/U T/I C
58. Pure water ionizes very slightly, producing a very low concentration of ions in solution. (10.2) K/U T/I A
    (a) What ions are likely to result from the ionization of water?
    (b) Predict the relative concentrations of these ions in pure water. Explain your prediction.
59. A student was given five unlabelled, identical beakers, each of which contained 250 mL of a clear, colourless liquid. He was told that one liquid was a strong acid and one was a weak acid. The other three liquids were a strong base, a weak base, and pure water. The acids and bases were all at the same concentrations. (10.2) T/I
    (a) How could the student safely identify each liquid?
    (b) Why does the student need to know that the acids and bases have the same concentration?
60. Not all acid–base indicators change colour in a neutral solution. For example, phenolphthalein changes colour in a slightly basic solution at a pH of approximately 8.5. Methyl orange changes colour in an acidic solution around pH 4. Explain why these two indicators would not be a good choice for a titration of hydrochloric acid with sodium hydroxide solution. (10.1, 10.3) T/I A

Evaluation

61. A student has a flask containing a solution and one of the acid–base indicators shown in Table 1 in Section 10.1. The solution is blue. (10.1) K/U T/I A
    (a) Is the solution most likely acidic or basic? Explain your answer.
    (b) The student thinks the indicator is bromothymol blue. Her lab partner thinks it is litmus. Describe how the students could determine which indicator was used.

62. When a laboratory investigation results in a waste solution, your teacher may give you instructions regarding the disposal of the waste solution. Some solutions are safe to wash down the drain with plenty of water. Other solutions should be collected and treated before disposing of them. For each of the following solutions, explain what should be done to the solution so that it can be safely poured down the drain with plenty of water. Include a test that you would perform to ensure that each solution is safe to discard. (10.1, 10.2) T/I A
    (a) a solution that contains nitric acid and has a pH of 4
    (b) a sodium hydroxide solution that has been diluted but still has a pH of 10
63. A student stated that all the steps used in the titration sample problems are not necessary. She said that one could just use the equation $V_A c_A = V_B c_B$ to determine any unknown concentration or volume. Another student pointed out that this equation could be used for some, but not all, titrations. (10.3) K/U T/I A
    (a) Describe a reaction in which the equation could correctly be used. Why can the equation be used for this titration?
    (b) Describe a reaction in which it would not be appropriate to use the student's equation. Why can the equation not be used for this titration?
64. A student measured the mass of a sample of sodium hydroxide, $NaOH(s)$, which had been stored in humid conditions. He then used the solid sample to make a solution. This solution was to be used as a titrant to determine the concentration of a hydrochloric acid solution, $HCl(aq)$. Was the concentration he calculated for the acid higher or lower than the actual concentration? Explain your answer. (10.3) K/U T/I A
65. One student states that, when performing a titration, the meniscus of the titrant must be exactly at the 0 mL mark on the burette. Another student argues that the titrant level can be below the 0 mL mark. Which student do you think is correct? Explain your answer. (10.3) T/I
66. Two students had titration results that did not match the results of the other students in the class. Their situations are described below. For each student, state his or her error. Then, describe what the student should do to obtain better results. (10.3) K/U T/I A
    (a) Thomas had difficulty turning the stopcock on his burette, so he used a graduated cylinder instead, measuring the initial and final volumes.
    (b) Latoya took her initial burette reading, added the titrant until the indicator in the unknown solution first changed colour, then stopped titrating and recorded her final burette reading.

Reflect on Your Learning

67. (a) What acids have you used in the past week?
    (b) What bases have you used in the past week?
    (c) From their properties, how did you know that these materials were acids and bases? K/U T/I
68. (a) What were your thoughts about the safety of using acids before reading this chapter?
    (b) What have you learned that changed your thoughts about the safety of using acids? Give reasons for your changed thoughts. T/I
69. Citric acid is found in foods such as lemons and oranges. Do you think citric acid is a weak acid or a strong acid? Explain your answer. K/U T/I A
70. Field trials in Nova Scotia have shown that ethanoic acid kills Canadian thistles and certain other unwanted plants when it is sprayed on them in concentrations from 5 % to 20 %. Consider the weeds that sometimes grow in your schoolyard. Why might you treat them with ethanoic acid instead of sulfuric acid? T/I A
71. Sometimes, standardized procedures are used for both scientific and everyday purposes. K/U T/I A
    (a) How do you know when the stopcock on the burette is closed and when it is open?
    (b) How is your answer to Part (a) similar to the operation of shut-off valves in the plumbing where you live or at your school?
    (c) Can you think of any other shut-off valves that operate the same way?

Research

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72. In this chapter you learned about two types of acids: oxyacids and binary acids. K/U T/I
    (a) What is a binary acid?
    (b) Research an example of an acid that is neither an oxyacid nor a binary acid.
    (c) Find out more about the acid that you named in Part (b). What are its characteristics and uses?
73. Table 1 in Section 10.1 shows the colour changes observed in four different acid–base indicators. There are other indicators that can be used to show the presence of an acid or a base. K/U T/I A
    (a) Research the following acid–base indicators. For each indicator, give the colour changes it undergoes from acid to base.
        (i) methyl red
        (ii) bromphenol blue
        (iii) bromcresol purple
    (b) What is a universal indicator? What advantage does it have over other indicators?
    (c) What natural materials can act as indicators? List two natural indicators and describe the colour changes they undergo.
74. Part of the boundary between Canada and the United States runs through Lake Ontario. The northwest shore of this lake is the most populated and industrialized area in Ontario. From the start of the twentieth century through the 1980s, the pH of Lake Ontario was of concern. For example, the pH of the lake for part of 1979 averaged 8.15. T/I A
    (a) What environmental factors contributed to any acidity in the lake?
    (b) What factors contributed to making the lake more basic?
    (c) From the pH value given, which factor most affected the water in Lake Ontario during the late 1970s?
    (d) What type of remediation returned Lake Ontario to a pH that was more healthful for the organisms that lived in the lake or depended on it for a water supply?
75. In this chapter you learned about the Arrhenius theory of acids and bases. This theory is just one of several used to define and describe the behaviours of acids and bases. T/I A
    (a) Research the Brønsted–Lowry theory of acids and bases. How are acids and bases defined, according to this theory?
    (b) How does the Brønsted–Lowry theory relate to the Arrhenius theory?
    (c) Research the Lewis theory of acids and bases. How are acids and bases defined, according to this theory?
    (d) For which of these theories is ammonia, $NH_3(g)$, a base? Explain your answer.
76. The reaction between a strong acid and a strong base produces a neutral solution, so an indicator for such a reaction should change colour near pH 7. However, not all acid–base reactions involve a strong acid and a strong base. K/U A
    (a) Research what happens when a strong acid and a weak base react. What general statement can you make about the pH of the solution at the equivalence point?
    (b) Investigate the pH ranges for various acid–base indicators. Choose an indicator that would probably be appropriate for a titration between a strong acid and a weak base.
    (c) Research what happens when a weak acid and a strong base react. What general statement can you make about the pH of the solution at the equivalence point?
    (d) Suggest an indicator that would be appropriate for a titration between a weak acid and a strong base.
77. The compound KHP has been introduced as a primary standard that can be used to titrate a basic solution. K/U T/I
    (a) You know that a primary standard must be pure and chemically stable. Find out what other requirements are desirable in a primary standard.
    (b) In addition to KHP, name another acid that can be used as a primary standard to titrate bases.
    (c) Name two primary standards used to titrate acids.

Hard Water Analysis

Hard water is a nuisance in many Canadian communities. It can lead to "scale" formation in pipes and on heating elements, and form a scum with soap. It is caused by an excess of certain ions in the water, particularly calcium, $Ca^{2+}(aq)$, and magnesium, $Mg^{2+}(aq)$. The concentration of these ions can be determined by titrating a water sample with a solution containing the anion of ethylenediaminetetraacetic acid (abbreviated to EDTA). The chemical formula used for the anion is $EDTA^{4-}(aq)$. Calcium and magnesium ions in water bind with $EDTA^{4-}(aq)$ in a 1:1 ratio to form a larger ion.

Given that it has a −4 ionic charge, you might expect $EDTA^{4-}(aq)$ to bind to 2 metal ions rather than just 1. However, chemists believe that the structure of the metal–EDTA anion prevents this from occurring. Using magnesium, Mg^{2+}, as an example,

$$Mg^{2+}(aq) + EDTA^{4-}(aq) \rightarrow MgEDTA^{2-}(aq)$$

Eriochrome black T (abbreviated to EbT) is the indicator for this titration. At the endpoint the sample turns from dark red to blue. The endpoint is reached when all the $Mg^{2+}(aq)$ and $Ca^{2+}(aq)$ ions have reacted with $EDTA^{4-}(aq)$. In practice, this titration is most effective when the sample is kept basic by adding a small volume of a pH 10 buffer solution.

During the unit, you studied technologies used to soften water. In this task, you will use your understanding and skills to plan a process to soften water (Part A). You will then test the effectiveness of your plan by titrating water samples with $EDTA^{4-}(aq)$ before and after they are softened (Part B).

Purpose

To test the effectiveness of a water-softening method using titration by determining the concentration of the ions responsible for water hardness

Equipment and Materials

SKILLS HANDBOOK A1

- chemical safety goggles
- lab apron
- 25 mL volumetric pipette
- pipette pump
- two 125 mL Erlenmeyer flasks
- three 150 mL beakers
- 100 mL graduated cylinder
- 50 mL burette
- retort stand with burette clamp
- small funnel
- meniscus finder
- wash bottle of distilled water
- 250 mL hard water
- pH 10 buffer solution
- dropper bottle of Eriochrome black T indicator solution
- EDTA solution, $C_{10}H_{14}N_2Na_2O_8(aq)$ (0.01 mol/L)

pH 10 buffer solution is toxic and an irritant. Avoid skin and eye contact. Wash any spills on the skin immediately with cold water. Report any spills to your teacher.

Eriochrome black T indicator is flammable. Keep it away from open flames.

Procedure

SKILLS HANDBOOK A2.4, A3

Part A: Softening the Water Sample

1. Plan a procedure to soften 125 mL of your water sample using the understanding and skills you have developed in this unit.
2. Once you have your teacher's approval, put on your safety goggles and lab apron and proceed with your plan. Keep your softened water and about 125 mL of hard water (in labelled beakers) for analysis in Part B.

Part B: Titrating for Hard Water Ions

1. Use the volumetric pipette to transfer 25.00 mL of the hard water sample to one of the Erlenmeyer flasks.
2. Add 2 mL of pH 10 buffer solution and 4 drops of Eriochrome black T.
3. Prepare the burette for titration with the EDTA solution following titration procedures outlined in Tutorial 1 of Section 10.3.
4. Titrate the hard water sample with EDTA to an Eriochrome black T endpoint (**Figure 1**). The endpoint indicates when all of the calcium and magnesium ions have reacted. Note that the colour change of Eriochrome black T at the endpoint is slower than the colour changes of acid–base indicators. Allow 1 or 2 s for the change to occur. Record your observations carefully.

WEB LINK

This titration is shown in an online video, which also includes many good titration tips. To see the video,

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5. Repeat the titration of your hard water until you have three trials with consistent results.
6. Repeat Steps 1 to 5 using your softened water in place of the hard water. Record your observations carefully.
7. Repeat the titration until you have three trials with consistent results.
8. Dispose of all chemicals as directed by your teacher.

Figure 1 Run the titrant into the sample slowly as you approach the endpoint.

Analyze and Evaluate

SKILLS HANDBOOK A6

(a) Briefly describe how the process you chose softens water. K/U
(b) Use your two or three best trials to calculate the average concentration of hardness ions in the hard water. T/I
(c) Use your two or three best trials to calculate the average concentration of hardness ions in the softened water. T/I
(d) What effect, if any, would the following errors have on the concentration of the hardness ions? Justify your answer in each case. T/I
  (i) Using tap water rather than distilled water to prepare the EDTA solution T/I
  (ii) Not clearing a large air bubble in the tip of the burette T/I
(e) Evaluate the design of Part B of the investigation and your skill at conducting it. Are there any parts of the investigation that might cast doubt on your results? Note any likely sources of error. T/I
(f) Suggest improvements to the experimental design for Part B. T/I
(g) Evaluate your water-softening procedure in Part A. Compare the concentration of metal ions in your softened water to that of other groups. T/I
(h) Suggest improvements to your procedure in Part A. T/I
(i) Why was it important to evaluate Part B of this investigation before Part A? T/I
(j) Why is hard water a nuisance? Research some of the problems that result from hard water, and briefly describe at least three different ways that people use to solve or avoid these problems. T/I A

Apply and Extend

(k) Convert your answer to (b) to parts per million (ppm). T/I

ASSESSMENT CHECKLIST

Your completed Unit Task will be assessed according to these criteria:

Knowledge/Understanding

- ✓ Demonstrate an understanding of methods of softening water.
- ✓ Demonstrate how the presence of hardness ions affects the properties of water.

Thinking/Investigation

- ✓ Plan safe and effective procedures for softening the water samples.
- ✓ Conduct the procedures safely and effectively.
- ✓ Record all observations carefully and in an organized manner.
- ✓ Analyze the results.
- ✓ Evaluate the procedure.

Communication

- ✓ Prepare a suitable lab report that includes the complete procedures, summary of the observations, any necessary analysis, and evaluation.

Application

- ✓ Evaluate the effectiveness of the water softening method using your titration data.

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

For each question, select the best answer from the four alternatives.

1. Water's special qualities depend largely on which of its following characteristics? (8.1) K/U
   (a) very strong covalent bonds between hydrogen and oxygen
   (b) hydrogen bonds that form between water molecules
   (c) weak covalent bonds that form between water molecules
   (d) very strong hydrogen bonds that form between hydrogen and oxygen within water molecules
2. Approximately what fraction of Earth's water supply is available as fresh water in liquid form? (8.1) K/U
   (a) 50 %
   (b) 25 %
   (c) 10 %
   (d) 2.5 %
3. Which of the following statements is true of all solutions? (8.2) K/U
   (a) All solutions are homogeneous and have only one phase.
   (b) All solutions are clear and colourless, and have only one phase.
   (c) All solutions are clear and colourless, and include only one solvent and one solute.
   (d) All solutions are homogeneous and include only one solvent and one solute.
4. Which of the following best explains the cleaning action of soaps and detergents? (8.3) K/U
   (a) The positive cation and negative anion attach to both polar and non-polar substances.
   (b) The polar and non-polar ends of the anion attach to both polar and non-polar substances.
   (c) The positive and negative ends of the anion attach to both polar and non-polar substances.
   (d) The polar and non-polar ends of the cation attach to both polar and non-polar substances.
5. Which part of a solubility curve identifies an unsaturated solution? (8.5) K/U
   (a) the part of the graph above the curve
   (b) the part of the graph on the curve
   (c) the part of the graph below the curve
   (d) the part of the graph just below and just above the curve
6. What volume of 0.750 mol/L potassium chloride solution can you prepare with 50.000 g of solid potassium chloride? (8.6) K/U
   (a) 0.894 L
   (b) 1.250 L
   (c) 1.118 L
   (d) 0.750 L
7. A 1 L sample of a solution was emptied into a 4 L container, which was then filled with water. Which of the following is true of the resulting solution in the 4 L container? (8.7) K/U T/I
   (a) Its concentration increased by a factor of four.
   (b) Its concentration increased by a factor of three.
   (c) Its concentration stayed constant since the amount of solute did not change.
   (d) Its concentration decreased by a factor of four.
8. In the total ionic equation below, which is the complete list of spectator ions?
   $2\ Na^+(aq) + 2\ Cl^-(aq) + 2\ Cu^+(aq) + SO_4^{2-}(aq) \rightarrow$
   $2\ Na^+(aq) + SO_4^{2-}(aq) + 2\ CuCl(s)$ (9.1) K/U
   (a) $Na^+(aq)$
   (b) $Na^+(aq)$; $SO_4^{2-}(aq)$
   (c) $2\ Cl^-(aq)$
   (d) $2\ Cl^-(aq)$; $Cu^{2+}(aq)$
9. Which of the following ions is very soluble when paired with OH^- but only slightly soluble when paired with SO_4^{2-}? (9.1) K/U
   (a) K^+
   (b) Ca^{2+}
   (c) Pb^{2+}
   (d) Ag^+
10. How many people in the world lack basic sanitation facilities? (9.2) K/U
    (a) about 100 million
    (b) less than 1 % of the world's population
    (c) almost 40 % of the world's 6.8 billion people
    (d) more than 80 % of the world's 6.8 billion people
11. In a qualitative analysis to test for the presence of two different cations, two different reagents are added to the sample solution. Why should the investigator use an excess of the first reagent? (9.3) K/U
    (a) to make sure that each cation partially precipitates
    (b) to make sure that both cations fully precipitate
    (c) to make sure that neither cation forms a precipitate
    (d) to make sure that all of one cation precipitates

12. Iron(III) chloride completely dissociates when dissolved in water. In a 0.6 mol/L solution of iron(III) chloride, what will be the concentration of Cl^- ions? (9.5) K/U
    (a) 0.6 mol/L
    (b) 0.2 mol/L
    (c) 1.8 mol/L
    (d) 1.0 mol/L
13. Which of the following properties is shared by both acids and bases? (10.1) K/U
    (a) electrical conductivity
    (b) sour taste
    (c) low pH
    (d) bitter taste
14. What is the correct name for $HBrO_2(aq)$? (10.1) K/U
    (a) bromic acid
    (b) bromous acid
    (c) perbromic acid
    (d) hydrobromide
15. An acid is used to bring the pH of a solution from pH 9 to pH 6. How many times more acidic is the final solution than the initial solution? (10.2) T/I
    (a) 3
    (b) 10
    (c) 100
    (d) 1000
16. A weak acid is dissolved in water to make a solution with a concentration of 1.0 mol/L. What is the concentration of H^+ ions in the solution? (10.2) K/U
    (a) 1.0 mol/L
    (b) 2.0 mol/L
    (c) less than 1.0 mol/L
    (d) greater than 1.0 mol/L
17. Which of the following is true of an endpoint of an acid–base titration but not an equivalence point? (10.3) K/U A
    (a) It is calculated rather than observed.
    (b) It is identified by a colour change.
    (c) It always identifies the precise point at which neutralization occurs.
    (d) It never identifies the precise point at which neutralization occurs.
18. Exactly 1.500 mol of $H_2SO_4(l)$ is dissolved in water. How much NaOH(s) will be needed to neutralize the acid? (10.3) T/I
    (a) 120.0 g NaOH(s)
    (b) 1.5 mol NaOH(s)
    (c) 60.0 g NaOH(s)
    (d) 0.75 mol NaOH(s)

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

19. The main reason that water exists as a liquid at room temperature, not as a gas, is that it forms hydrogen bonds. (8.1) K/U
20. The production of a single hamburger requires about 1000 L of water. (8.1) K/U
21. When one liquid dissolves in another, the liquid with the greater molar mass is considered to be the solvent. (8.2) K/U
22. Alcohol is a polar liquid, so the best way to rinse the taste of alcohol from your mouth would be to drink milk. (8.3) A
23. When you add extra solute to an already saturated solution, the extra solute becomes supersaturated. (8.5) K/U
24. A supersaturated solution contains more solute than the solvent can usually hold at a given temperature. (8.5) K/U
25. An aqueous solution of potassium chloride that contains 40.00 g of the solute in 0.25 L of solution has a concentration of 2.15 mol/L. (8.6) K/U
26. When solute is added to a solution, the concentration of the solution decreases. (8.7) K/U
27. When 2.30 L of a 2.00 mol/L solution is diluted to a final volume of 9.20 L, its final concentration is 1.00 mol/L. (8.7) T/I
28. Potassium ions form insoluble compounds with all anions. (9.1) K/U
29. To obtain a net ionic equation, subtract the spectator ions from the total ionic equation. (9.1) K/U
30. The only kind of contaminants that dissolve in water supplies are chemical contaminants. (9.2) K/U
31. The alkali metals do not readily form precipitates, so they must be identified using a flame test. (9.3) K/U
32. When 1.00 mol of iron(III) oxide reacts with 50.0 g of aluminum, according to the following equation, the compound is the limiting reagent.
    $Fe_2O_3(s) + 2\ Al(s) \rightarrow 2\ Fe(s) + Al_2O_3(s)$ (9.5) K/U
33. Phosphorous acid has the formula $H_3PO_4(aq)$. (10.1) K/U
34. Comparing a strong acid and a weak acid with the same concentration, the weak acid will have a higher pH. (10.2) K/U
35. Bromothymol blue changes colour at pH 7.0. (10.3) K/U
36. The pH of the solution being titrated changes gradually near the equivalence point. (10.3) K/U

Knowledge

For each question, select the best answer from the four alternatives.

1. The fact that water is a liquid at room temperature suggests which of the following? (8.1) K/U
   (a) The forces between atoms in a water molecule are relatively strong.
   (b) The forces between atoms in a water molecule are relatively weak.
   (c) The forces between water molecules are relatively strong.
   (d) The forces between water molecules are relatively weak.
2. What kind of mixture is air? (8.2) K/U
   (a) a homogeneous solution that is completely transparent at all times in all areas
   (b) a mixture that can be homogeneous or heterogeneous depending on local composition
   (c) a heterogeneous mixture that is never transparent
   (d) a heterogeneous mixture in the dark, but a homogeneous mixture in bright light
3. Why do pesticides such as DDT tend to accumulate in the bodies of living organisms? (8.3) K/U
   (a) DDT is non-polar, so it does not get flushed away in urine.
   (b) DDT is polar, so it does not get flushed away in urine.
   (c) DDT is high in London dispersion forces, so it does not get flushed away in urine.
   (d) DDT is soluble in water, so it tends to stay in body tissues.
4. 10 g of a solute is added to 1.00 L of pure water. Half of the added solute fails to dissolve. What can you conclude? (8.5) A K/U
   (a) The solution is saturated.
   (b) The solution is supersaturated.
   (c) The solution reached but did not pass its saturation level.
   (d) The solution was initially unsaturated and then passed its saturation level.
5. Which of the following occurs in the reaction between zinc metal and copper(II) sulfate solution:
   $Zn(s) + CuSO_4(aq) \rightarrow Cu(s) + ZnSO_4(aq)$? (9.1) K/U
   (a) Zinc replaces copper because zinc is more reactive than copper.
   (b) Zinc replaces copper because zinc sulfate is more soluble than copper sulfate.
   (c) Zinc replaces copper because zinc sulfate is less soluble than copper sulfate.
   (d) Zinc replaces copper because zinc has a smaller positive charge than copper.
6. Which kind of contaminants may exist in solution within a community's water supply? (9.2) K/U
   (a) physical contaminants
   (b) biological contaminants
   (c) chemical contaminants
   (d) all of the above
7. Which of the following is a positive test for the presence of nickel? (9.3) K/U
   (a) a characteristic pink colour in the presence of dimethylglyoxime
   (b) skin allergies when jewellery contains nickel
   (c) a bright yellow precipitate in the presence of chloride ions
   (d) a strong odour in the presence of an acid
8. Adding sodium sulfide to a solution that contains a single cation produces a precipitate. In a flame test, the solution produces a light-green flame. Which ion is most likely to be in the solution? (9.1, 9.3) K/U
   (a) silver
   (b) calcium
   (c) zinc
   (d) copper
9. A 0.40 mol/L solution of sodium carbonate, $Na_2CO_3(aq)$, completely dissociates in water. What will be the concentration of sodium ions in the solution? (9.5) K/U
   (a) 0.40 mol/L (c) 0.80 mol/L
   (b) 0.20 mol/L (d) 1.2 mol/L
10. A substance feels slippery on your skin and turns blue in the presence of litmus. Which of the following is most likely to be true? (10.1) K/U
    (a) The substance does not dissolve in water.
    (b) The substance is a poor conductor of electricity.
    (c) The substance has a sour taste.
    (d) The substance reacts with fats.
11. Which of the following statements is true? (10.2) K/U
    (a) Ionic compounds ionize; molecular compounds dissociate.
    (b) Ionic compounds dissociate; molecular compounds ionize.
    (c) Ionic compounds both dissociate and ionize; molecular compounds dissociate only.
    (d) Ionic compounds and molecular compounds both ionize and dissociate.

12. Equal volumes of 1.0 mol/L solutions of hydrochloric acid, HCl(aq), and sodium hydroxide, NaOH(aq), are combined. Which of the following is true of the solution that results? (10.2) K/U
    (a) The solution contains a dissolved ionic compound.
    (b) Both anions are present in the solution.
    (c) The resulting solution will be acidic.
    (d) The resulting solution will be basic.

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

13. Water contracts as it cools to 0 °C, causing ice to float on the surface of liquid water. (8.1) K/U
14. An opaque mixture of two solids could be a homogeneous solution. (8.2) K/U
15. As sodium chloride dissolves in water, each negative chloride ion gets surrounded by the oxygen ends of water molecules, completely hydrating the ion. (8.3) K/U
16. The phrase "like dissolves like" refers to the idea that polar solvents will dissolve non-polar solutes. (8.3) K/U
17. No gas bubbles are visible in a still, sealed bottle of pop because much of the carbon dioxide gas is dissolved. (8.5) K/U
18. When diluting a concentrated acid, always add the water to the acid. (8.6) K/U
19. Ions that do not actively participate in a reaction are called spectator ions. (9.1) K/U
20. Improving water quality in a developing country may increase school attendance. (9.2) K/U
21. Flame tests can reliably identify cations in solution when more than one cation is present. (9.3) K/U
22. Acidic foods soften teeth, so brushing your teeth immediately after eating acidic foods is a good idea. (10.1) K/U
23. Acids readily react with carbon dioxide to produce carbonate compounds. (10.1) K/U
24. Though the equivalence point and the endpoint in a titration are usually different, they can be identical if the correct indicator is chosen. (10.3) K/U

Match each property on the left with each appropriate term on the right.

25.

(a) pH < 7	(i) acid
(b) pH > 7	(ii) base
(c) sour taste	(iii) both acid and base (10.1) K/U
(d) bitter taste	
(e) slippery feel	
(f) conductivity varies	
(g) turns litmus red	

Match each titration term on the left with its definition on the right.

26.

(a) burette	(i) the standardized solution of known concentration
(b) titrant	(ii) the point at which indicator colour changes permanently
(c) titration	(iii) the theoretical point at which neutralization is complete
(d) receiving flask	(iv) the calibrated tube that is used measure titrant
(e) endpoint	(v) the procedure used to determine the concentration of a solution
(f) equivalence point	(vi) the container that holds a known volume of a solution of unknown concentration (10.3) K/U

Write a short answer to each question.

27. Which kind of bonding is responsible for most of the special characteristics of water? (8.1) K/U
28. A solution contains 51 mL of ethanol and 49 mL of propanone. Which substance is the solvent? Explain. (8.2) K/U
29. What process do ionic compounds undergo as they dissolve in water? (8.3) K/U
30. A liquid solvent at 20 °C holds more solute than the same solvent at 30 °C. What can you conclude about the solute? (8.5) A
31. Name two types of substances that tend not to dissolve in water. (8.3, 9.1) K/U
32. Lead(II) ions will precipitate with most anions listed in Table 1 in Section 9.1. Which anion does not form a precipitate with lead(II) ions? (9.1) K/U
33. A solution contains potassium, silver, and sodium ions. Which of these cations forms the least soluble compound when mixed with sulfate ions, $SO_4^{2-}(aq)$? (9.1) K/U
34. Which kind of equation identifies highly soluble compounds written in their dissociated ionic form? (9.1) K/U
35. Which kind of equation identifies only the entities that actually react and change in an equation? (9.1) K/U
36. List any spectator ions in this equation.
    $$Fe(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Cu(s) + Fe^{2+}(aq) + SO_4^{2-}(aq)$$ (9.1) K/U
37. Shampoo manufacturers want to avoid precipitates forming when people wash their hair in hard water. What type of compound is included in most shampoos so that precipitates do not form? (9.2) K/U

38. The goal of water treatment is not to purify water, but rather to make it potable. What does "potable" mean? (9.2) K/U

39. Spectroscopy is essentially a sophisticated form of what kind of analytic test? (9.3) K/U

40. In the reaction represented below, 2.0 mol of reactant A was combined with 3.0 mol of reactant B.

$A + 2\,B \rightarrow AB_2$

Which reagent was the limiting reagent? (9.5) K/U

41. Soap is made using a formula discovered by many ancient civilizations. Which two ingredients were combined to make soap? (10.1) K/U

42. In the presence of phenolphthalein indicator, what colours do you expect to see with acids and bases? (10.1) K/U

43. Write the name of each of the following acids: (10.1) K/U C
 (a) $H_2CO_3(aq)$
 (b) $HF(aq)$
 (c) $H_3PO_4(aq)$
 (d) $H_2S(aq)$
 (e) $HClO_4(aq)$

44. Write the formula for each of the following acids: (10.1) K/U C
 (a) phosphorous acid
 (b) hydrochloric acid
 (c) hydroiodic acid
 (d) chlorous acid
 (e) chloric acid

45. A metal and an acid react in a test tube. What observation would you expect to make? (10.1) K/U

46. (a) What process describes a compound separating into its ions as it dissolves in water?
 (b) What process describes uncharged particles reacting with water and becoming ions? (10.2) K/U

47. Write a net ionic equation that applies to all acid–base neutralization reactions. (10.2) K/U C

48. Which would ionize more completely in water, a strong acid like hydrochloric acid or a weak acid like phosphoric acid? (10.2) K/U

49. Which would conduct electricity better, a strong acid like sulfuric acid or a weak acid like ethanoic acid? (10.2) K/U

50. What analytical process is used to find the concentration of an unknown solute in a solution? (10.3) K/U

51. Two events must coincide in order for a titration to produce accurate results. What are these two events? (10.3) K/U

Understanding

52. What special kind of molecular structure does water have? List three unique characteristics of water and explain how each characteristic is related to water's special structure. (8.1) K/U

53. Is air a homogeneous solution or a heterogeneous mixture? Explain. (8.2) K/U

54. Ice water has two phases—solid ice and liquid water. Is ice water a heterogeneous mixture or a homogeneous mixture? Explain. (8.2) K/U A

55. Will hexane, $C_6H1_4(l)$, dissolve in water? Explain why or why not. (8.3) K/U

56. Why do non-polar toxins tend to be more harmful to biological organisms than polar toxins? (8.3) K/U A

57. Soaps and detergents are categorized as surfactants. (8.3) K/U C A
 (a) What property do soaps and detergents have that make them surfactants?
 (b) What structure do soaps and detergents have? Include a diagram to illustrate your answer.
 (c) How does the structure of soaps and detergents help them clean?

58. (a) How would you make a supersaturated solution of glucose in water?
 (b) How would you make this supersaturated solution into a saturated solution, without adding water? (8.5) T/I A

59. A student dissolved 1.00 mol of solid potassium chloride in 1.00 L of water. The student then labelled the solution to have an amount concentration of 1.00 mol/L. What mistake did the student make? (8.6) T/I

60. Aluminum metal, $Al(s)$, reacts with hydrochloric acid to produce aluminum chloride, $AlCl_3(aq)$, and hydrogen gas. (9.1) T/I C
 (a) Write the balanced formula equation for the reaction.
 (b) Identify any solids in the equation.
 (c) Write the total ionic equation.
 (d) Identify the spectator ions in the equation.
 (e) Write the net ionic equation.

61. Write a formula equation, a total ionic equation, and a net ionic equation for the reaction of iron and silver nitrate solution. (9.1) T/I C

62. Due to sewage contamination, a pond begins to lose many of its fish species. How might installing a fountain in the pond water with oxygen improve the situation? (9.2) A

63. Arrange the steps in the water treatment process in the correct sequence. (9.2) K/U
    - filtration
    - fluoridation
    - aeration
    - disinfection
    - ammoniation
    - collection
    - removal of debris
    - softening
    - post-chlorination
    - coagulation, flocculation, and sedimentation
64. Astronomers are able to identify chemicals in distant stars by analyzing what? (9.3) T/I
65. A 2.00 mol/L solution of iron(III) sulfate $Fe_2(SO_4)_3(aq)$, is prepared in a beaker. (9.5) T/I
    (a) What is the concentration of iron(III) sulfate?
    (b) What is the concentration of iron(III) ions?
    (c) What is the concentration of sulfate ions?
66. A piece of metal is dropped into hydrochloric acid. (10.1) K/U
    (a) Will a reaction necessarily take place? Explain your answer.
    (b) If a reaction occurred, predict what the product(s) would be.
    (c) What visible evidence would there be that the reaction is taking place?
67. Calcium oxide reacts with hydrochloric acid to form water and calcium chloride. Write a balanced equation for this reaction. (10.1) K/U T/I C
68. Iron(II) sulfide is the compound that causes the greenish colour around the yolks of hard-boiled eggs. When pure iron(II) sulfide reacts with hydrochloric acid, a very unpleasant smell results. (10.1) K/U T/I C
    (a) What are the products of this reaction?
    (b) Write a balanced equation for this reaction.
69. A chemist blows carbon dioxide into a solution in a test tube. This is sufficient to produce a precipitate in the test tube that turns out to be a carbonate. (10.1) K/U T/I
    (a) What kind of substance was in the test tube?
    (b) What property of bases does this example illustrate?
70. Acids in their undissolved form can be represented by a general formula. (10.1) K/U C
    (a) Write the general undissolved acid formula.
    (b) Write a general equation to explain what happens when the acid dissolves in water.
    (c) Write an equation to describe what happens when an acid ionizes in water. Choose an acid to use as an example.
71. A strong acid, $H_2SO_4(aq)$, is completely neutralized with a strong base, $KOH(aq)$. (10.2) K/U T/I C
    (a) Predict the products of the neutralization reaction.
    (b) Write a balanced formula equation for the neutralization reaction.
    (c) Write a total ionic equation for the neutralization reaction.
    (d) Write a net ionic equation for the neutralization reaction.
    (e) What conclusion can you draw about acid–base neutralization from the net ionic equation that you wrote?
72. Describe the difference between a strong acid and a weak acid. (10.2) K/U C
73. Explain why strong acids conduct electricity better than weak acids, assuming that the two acids are at equal concentrations. (10.2) K/U C
74. Explain what is wrong with this statement: "A solution of pH 10 is 100 times more acidic than a solution of pH 9." Then write a better statement to replace it. (10.2) K/U C

Analysis and Application

75. Write a dissociation equation for each of the following compounds: (8.3) K/U T/I
    (a) Li_2S
    (b) magnesium chloride
    (c) $Al_2(SO_4)_3$

76. In an investigation, 50 g of potassium chloride is added to 100 g of water at 60 °C. The mixture is stirred until no more potassium chloride dissolves. Use the graph in **Figure 1** to answer the following questions. (8.5) K/U T/I

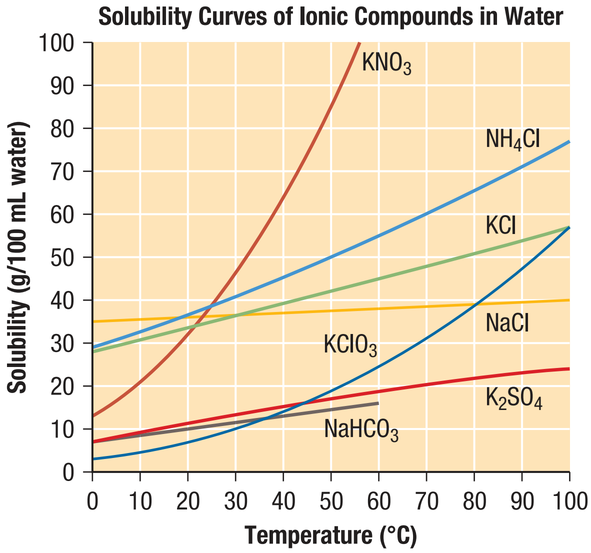


Figure 1

(a) Is the solution saturated or not? Explain how you know.

(b) At what temperature would this solution become unsaturated? Explain your answer.

77. Use Figure 1, above, to consider the solubilities of potassium nitrate and potassium chloride. (8.5) K/U T/I

(a) Which solute is more soluble at 15 °C, potassium nitrate or potassium chloride? Explain your answer.

(b) At what temperature are potassium nitrate and potassium chloride equally soluble?

(c) At what temperature(s) is potassium nitrate more soluble than potassium chloride?

78. What volume of a 0.25 mol/L solution of sodium chloride can be made with 50.00 g of sodium chloride? (8.6) K/U T/I

79. Water is added to 0.500 L of 2.500 mol/L hydrochloric acid to make a final volume of 4.250 L. What is the concentration of the resulting solution? (8.7) K/U T/I

80. Write the formula equation, the total ionic equation, and the net ionic equation for the reaction of silver metal and copper(II) sulfate solution. Then, identify the spectator ions. (9.1) K/U T/I C

81. Which parts of the water treatment process remove each of the following types of contaminants? (9.2) K/U

(a) physical contaminants

(b) biological contaminants

(c) chemical contaminants

82. After an influx of fecal material into a pond, the water suddenly blooms with aquatic organisms. (9.2) T/I C

(a) This influx increases the nutrient content of the water. Why, then, is it considered harmful?

(b) Identify the sequence of events that occurs as a result of the higher nutrient content of the water.

(c) Identify organisms that may not be harmed by this influx, and explain why they are unharmed.

83. Water treatment plants use sodium carbonate to "soften" water by removing calcium. Explain how this process works. Include a chemical equation in your explanation. (9.2) K/U C

84. A solution is suspected of containing barium ions, lead(II) ions, or both. According to an analysis flow chart, potassium iodide is added. (9.3) K/U T/I

(a) What would you expect to see if the solution contains lead?

(b) After filtering, sodium sulfate is added. What would you expect to see if the sample contains barium?

(c) The sample then undergoes further analysis and is found to still contain lead. How can this be possible?

85. A solution is suspected of containing strontium ions, $Sr^{2+}(aq)$, manganese ions, $Mn^{2+}(aq)$, or both. (9.3) K/U T/I

(a) Plan a qualitative analysis for identifying the ions. Describe your plan.

(b) Draw a flow chart to represent your qualitative analysis.

86. Determine the minimum volume of 0.25 mol/L sodium carbonate needed to precipitate all of the calcium ions in 200 mL of 0.20 mol/L calcium chloride by answering the following: (9.5) K/U T/I

(a) Write a balanced equation for the reaction.

(b) What amount of calcium chloride is present?

(c) What is the mole ratio of sodium carbonate to calcium chloride?

(d) What amount of sodium carbonate, $n_{Na_2CO_3}$ is needed to precipitate all of the calcium ions?

(e) What volume of sodium carbonate is needed?

87. A reaction features 20.0 g of substance A and only 1.0 g of substance B. Jeff determines that substance A is the limiting reagent. Malia insists that Jeff must be wrong because, by mass, 20 times as much substance A is used as substance B. Can Jeff's identification of the limiting reagent be correct? Explain your answer. (9.5) T/I C A

88. A 0.25 L sample of 2.20 mol/L $FeCl_3(aq)$ reacts with 37.50 g of dissolved $H_2S(aq)$. The products are $HCl(aq)$ and $Fe_2S_3(s)$. (9.5) K/U T/I C
   (a) Write a balanced equation for the reaction.
   (b) What amounts of $FeCl_3$ and H_2S are present as reactants?
   (c) What is the mole ratio of H_2S to $FeCl_3$ in the balanced equation?
   (d) What is the amount ratio of H_2S to $FeCl_3$ present?
   (e) Which reactant is the limiting reagent?
   (f) What mass of Fe_2S_3 will be produced by this reaction?
89. Draw two diagrams that compare the ionization of strong and weak acids. (10.2) K/U T/I C
90. Equal volumes of solutions of pH 5 and pH 2 are compared. The solutions have equal amount concentrations. (10.2) K/U T/I
   (a) Which solution is more acidic?
   (b) How much more acidic is one solution than the other?
   (c) Which solution has a higher concentration of H^+ ions?
91. A 1.0 mL sample of 1.0 mol/L hydrochloric acid is diluted with water to make a solution whose volume is 1.0 L. (10.2) K/U T/I C
   (a) Write the ionization equation for hydrogen chloride in water to form hydrochloric acid.
   (b) What is the concentration of H^+ ions in the original solution?
   (c) What is the concentration of H^+ ions in the diluted solution?
   (d) How much less acidic is the diluted solution than the original solution?
   (e) Relate the change in acidity to the change in hydrogen ion concentration.
92. **Figure 2** shows a titration curve involving sodium hydroxide and hydrochloric acid. (10.3) K/U T/I A

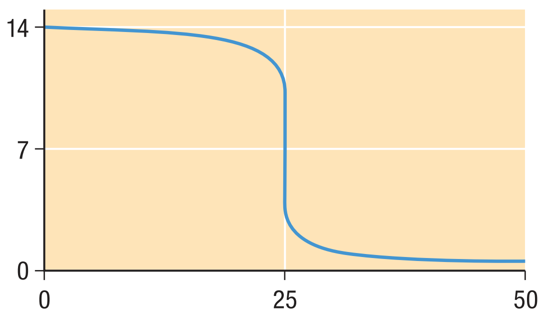


Figure 2

   (a) Suggest a title and axis labels for the graph.
   (b) Explain the shape of the graph.
93. Three 0.0100 L samples of sulfuric acid of unknown concentration are titrated with 0.20 mol/L potassium hydroxide solution. A phenolphthalein indicator is used to show the endpoint. The three trials used 15.40 mL, 15.36 mL, and 15.44 mL of the $KOH(aq)$ titrant. (10.3) K/U T/I C
   (a) Calculate the average volume of KOH titrant used, in litres.
   (b) Write a balanced equation for the reaction.
   (c) Determine the amount of $KOH(aq)$ used to neutralize the acid.
   (d) Determine the amount of acid that was neutralized.
   (e) Find the concentration of the acid.
94. A 0.600 g sample of potassium hydrogen phthalate, $KHC_8H_4O_4$, is dissolved and titrated with 16.50 mL of a sodium hydroxide solution of unknown concentration. Find the concentration of the sodium hydroxide solution. (10.3) K/U T/I C

Evaluation

95. What makes water ideal for development of life on Earth? (8.1) K/U A
96. Think about what you have learned about the demand for water in today's world. (8.1) K/U T/I A
   (a) How do you assess the global supply of fresh water compared to the overall global demand for the future?
   (b) How much impact can consumer behaviour have on water use in the future?
   (c) Where must improvements in water use be made in order for the world to meet its demand for water?
   (d) How do agricultural practices in North America affect water quality? What changes do you think need to be make in agricultural practices to limit the negative effects?
97. Two bottles of pop are opened. Bottle A has a temperature of 2 °C; bottle B has a temperature of 30 °C. After 15 min, which bottle is likely to have more fizz? Explain your answer. (8.5) T/I A
98. Student A took 1.00 L of a 1.00 mol/L sodium chloride solution and added water to make a final volume of 5.00 L of solution in his container. Student B had a 0.25 mol/L sodium chloride solution in her container. Both students claim that their solution is more concentrated. Which student is correct? How do you know? (8.7) K/U T/I

99. Students in a class evaluated different equations that show the reaction between sodium hydroxide, $NaOH(aq)$, and hydrochloric acid, $HCl(aq)$. Each student chose the type of equation—unbalanced formula equation, balanced formula equation, total ionic equation, or net ionic equation—that they thought best explains what happened during the reaction. Which type of equation do you think most students chose? Explain your answer. (9.1) K/U

100. Two students suspect that a solution contains lead(II) ions. Student A adds hydrochloric acid to a sample of the solution. No precipitate forms, so she concludes that no lead ions are present. Student B adds sulfuric acid to a different sample of the same solution, and a precipitate does form. He uses the results as confirmation that lead ions are present. Which student is correct? Or is neither student correct? Explain your answer. (9.3) K/U A

101. (a) Use what you know about naming acids to create a concept map that can be used to write the formula of an acid from its name.
(b) Use your concept map to write the formula for sulfurous acid. (10.1) K/U C A

102. A paleontologist finds material that may once have been a part of an ancient seabed. If it was an ancient seabed, it is likely to contain fragments of shells made of calcium carbonate, $CaCO_3(s)$, from marine life. To test for the presence of calcium carbonate a lab assistant treats a sample of the material with acid. Another lab assistant disagrees with this test and treats a sample with a base. Which lab assistant is performing a valid test for calcium carbonate? Explain your answer, including a prediction of what would be observed during a positive test. (10.1) K/U A

103. A compound is dissolved in water. (10.2) K/U A
(a) How could you test to see if ions formed as a result of dissolving?
(b) Suggest how you could determine whether dissociation or ionization occurred.

104. The compound ammonia, NH_3 has some unusual characteristics. (10.2) K/U C
(a) Why is ammonia sometimes classified as a base?
(b) How is ammonia different from other bases that you have learned about?
(c) What do you conclude from the behaviour of ammonia?

105. Evaluate three common acid–base indicators for use in titration of a strong acid with a strong base. (10.3) A

Reflect on Your Learning

106. Consider what you have learned about the structure and cleaning action of detergents. Then imagine you have been commissioned to develop a new detergent. Outline a plan for developing the new detergent. Describe how your detergent might be more effective than current formulas. (8.3) K/U A

107. Look back on what you have learned about highly soluble and slightly soluble ionic compounds. How can you account for the fact that a compound like copper(I) chloride is slightly soluble in water while copper(I) sulfate is highly soluble? What must be the difference in the two compounds? Include a diagram to help explain your answer. (9.1) K/U C A

108. Using water wisely and minimizing waste has an obvious consequence for water resources. How can efficient water conservation also help Canada solve its energy problems? (9.2) T/I A

109. Explain the strategy you use when planning a qualitative analysis to determine the presence of two cations, in the same solution, based on solubility. (9.3) K/U T/I

110. Consider how you could apply the concept of a limiting reagent to the situation of cars and tires. (9.5) K/U A
(a) With exactly 30 tires and 7 cars, which would be the "limiting reagent": cars or tires? Explain your answer.
(b) Describe a situation in which tires would be seen as the "limiting reagent."
(c) Comment on how appropriate the analogy of cars and tires is to the chemistry of limiting reagents. Does the analogy work? Explain your answer.

111. A chemist suggests the following changes to writing the formulas of acids. (10.1) T/I A
- Start the formulas of all strong acids with an "H" for hydrogen.
- For weak acids, position the "H" for hydrogen at the end of the formula.

(a) What advantages do you see to making these changes?
(b) What disadvantages do you see to making these changes?
(c) Do you think the changes should be implemented? Explain your answer.

112. Use what you have learned in this unit to solve the following puzzle. (10.2) K/U T/I A
- In a remote region of northern Canada, forest rangers found evidence that lightning strikes were killing wildlife such as waterfowl on small ponds and lakes at an unusually high rate.
- The rangers found no increase in lightning storms or the frequency or intensity of lightning that hit the general area.
- Forests in the nearby area have been severely damaged by acid rain.

(a) How might you explain the situation?
(b) What would you do to test your explanation?

113. A friend who has a summer job in a lab needs to perform a titration to find the concentration of an unknown solution. Write a general step-by-step procedure for your friend to follow. (10.3) K/U T/I C

Research

GO TO NELSON SCIENCE

114. In the 1950s James Watson and Francis Crick discovered the structure of DNA. They proposed a mechanism for how this complex molecule functions as the genetic material by which information is passed from generation to generation. Watson and Crick's theory depended on the idea of hydrogen bonding. T/I A
(a) What role do hydrogen bonds play in creating the structure of DNA?
(b) How did the identification of base pairing lead to the cracking of the genetic code?

115. The global water crisis is gaining increasing attention in the news media. T/I C A
(a) Research the water crisis in general. List what you feel to be the two most urgent concerns in global water use.
(b) Research a particular situation in a geographical location of your choosing. Create a report, in a format of your choice, on the conflicts and problems with water in that area.

116. Research the prospect of switchable solvents as illustrated by the work of Dr. Phillip G. Jessop at Queen's University in Ontario. T/I C A
(a) What problems do conventional (non-switchable) solvents now pose? Give a specific example.
(b) How do switchable solvents improve the situation?
(c) Draw a diagram to show how the switchable solvent system works.

117. The idea of "green" chemistry is intriguing to people who are interested in improving the environment. Research a company that is attempting to engage in green chemistry. T/I A
(a) Describe the company's basic philosophy.
(b) Describe some of the green projects that this company is currently undertaking.

118. Research the waste-water treatment situation in your community. T/I A
(a) What system does your community have for treating wastewater treatment system function?
(b) How well does the current water treatment system function?
(c) What plans are there for future improvements in the system?

119. The H^+ designation is not the only way to represent a hydrogen ion. T/I A C
(a) Research the hydronium ion and its relationship to the standard "H^+" ion used in this textbook.
(b) What are the advantages or disadvantages of each way to represent hydrogen ions?
(c) Would you recommend changing the way hydrogen ions are written?

120. Chromatography is a classic technique in chemistry for identifying unknown solutes. Research chromatography and describe a simple experiment that uses paper chromatography to separate the different components of a coloured ink. In your report, include a diagram of the experimental set-up. A C

121. The presence of drugs in drinking water is now a worldwide problem. For example, nine different drugs were found when 20 different drinking-water treatment plants in Ontario were tested by a national research institute. Research this problem and report your answers to the following questions: A C
(a) What did you find out about the sources of drug contamination?
(b) How concerned are scientists about the situation? In your answer, include reasons why scientists are concerned or are not concerned.
(c) What forms of remediation or prevention were proposed?

UNIT 5

Gases and Atmospheric Chemistry

OVERALL EXPECTATIONS

- analyze the cumulative effects of human activities and technologies on air quality, and describe some Canadian initiatives to reduce air pollution, including ways to reduce their own carbon footprint
- investigate gas laws that explain the behaviour of gases, and solve related problems
- demonstrate an understanding of the laws that explain the behaviour of gases

BIG IDEAS

- Properties of gases can be described qualitatively and quantitatively, and can be predicted.
- Air quality can be affected by human activities and technology.
- People have a responsibility to protect the integrity of Earth's atmosphere.

UNIT TASK **PREVIEW**

You will investigate a variety of methods of carbon capture and compare them through a cost–benefit analysis.

The Unit Task is described in detail on page 616. As you work through the unit look for Unit Task Bookmarks to see how information in the section relates to the Unit Task.

FOCUS ON **STSE**

GASES AND THEIR APPLICATIONS

We do not have to look very far to find an example of a gas. We live in a fluid mixture of gases, moisture, and suspended solids that we call the atmosphere. The gas molecules in this mixture are in constant motion, energized by the Sun. Only Earth's gravity prevents the atmosphere from escaping into space. Gas molecules in the atmosphere also push on Earth's surface. The pressure that this force exerts is critical for keeping water in the liquid state. Without it, the water molecules in liquid water could easily overcome their hydrogen bonding and evaporate or boil. In other words, without the atmosphere, liquid water could not exist.

There are many important applications of the properties of gases. We know, for example, that gases expand rapidly when heated. We rely on this property to move our vehicles and generate some of our electricity. Combustion of gasoline inside an automobile engine releases a mixture of extremely hot gases. The expansion of these gases pushes on a piston in each cylinder in the engine. This motion is ultimately responsible for making the wheels of the car turn.

The expansion of gases has important applications. Bread, for example, expands or "rises" because of carbon dioxide generated by yeast in the bread dough. Gases are also used to expand many of the foam products that we routinely use. Styrofoam, for example, is made from a hard, brittle plastic called polystyrene. During the manufacturing process, a gas is blown into the hot, liquid polystyrene. Tiny bubbles of gas expand the polystyrene much like rising bread dough. Until recently, the gases used to expand polystyrene were chlorofluorocarbons (CFCs). However, CFCs are damaging to Earth's ozone layer, which protects us from the harmful solar ultraviolet radiation. A major Styrofoam manufacturer has now developed a greener way to expand polystyrene using carbon dioxide. Carbon dioxide is used because it is inexpensive and does not damage the ozone layer. Carbon dioxide, however, is a greenhouse gas and contributes to climate change. To avoid adding more carbon dioxide to the atmosphere, waste carbon dioxide from other industrial processes is used to expand polystyrene. The production of Styrofoam illustrates how green alternatives provide a more sustainable way to manufacture products involving gases.

Questions

1. Use the concept of attractive forces between entities to explain why there are three common states of matter.
2. Weather reports often describe changes in atmospheric pressure. How might a change in pressure affect a gas? What effect might pressure changes have on weather?
3. How do changes in air pressure affect you?
4. What property of gases applies when you inflate a bicycle tire with a hand pump?
5. Describe two applications of gases—one consumer related and the other industrial. On what properties of gases are these applications based?
6. Give two examples of human activities that have a negative impact on the composition of the atmosphere. How might these activities be made greener?

CONCEPTS

- the kinetic molecular theory
- states of matter
- intermolecular forces
- the mole concept
- stoichiometry
- volume
- density

SKILLS

- write balanced equations for reactions involving gases
- calculate molar masses
- interpret graphed data
- identify relationships between variables
- use correct units and the correct number of significant digits when expressing the answer to a calculation
- plan and safely conduct investigations involving gases
- critically evaluate the results of investigations

Concepts Review

1. In which state of matter are entities usually closest together? Explain your answer. K/U
2. In which state of matter do entities exhibit the greatest motion? Explain your answer. K/U
3. Which state of matter is easiest to compress? Explain your answer. K/U
4. Describe what happens to entities during the
   (a) melting of a solid.
   (b) freezing of a liquid. K/U
5. Why do one mole of solid carbon dioxide and one mole of gaseous carbon dioxide have the same mass but different densities? K/U
6. List three different units of volume. K/U T/I
7. **Figure 1** shows the boiling points of hydrogen compounds of four of the Group 16 elements. The boiling point of each compound is plotted against the period on the periodic table in which the element is located. T/I A
   (a) Compare the boiling point of water, H_2O, to the boiling points of the other compounds.
   (b) If water followed the trend of the other compounds, what might its boiling point be?
   (c) Why does water not follow the trend in boiling points?

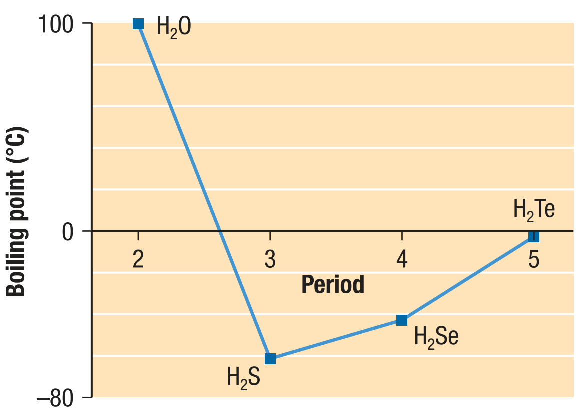


Figure 1

8. Which of the graphs in **Figure 2** shows a direct relationship and which shows an inverse relationship? Explain your answers. T/I

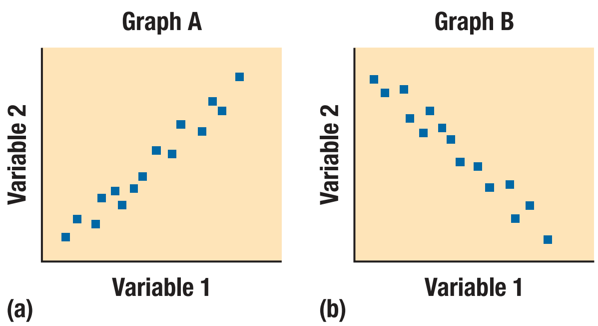


Figure 2

9. Look closely at the graph in **Figure 3**. T/I
   (a) Which gas expands the most when heated? Which gas expands the least? Justify your choices.
   (b) Use the graph to estimate the volumes of each gas at
      (i) 50 °C
      (ii) −100 °C

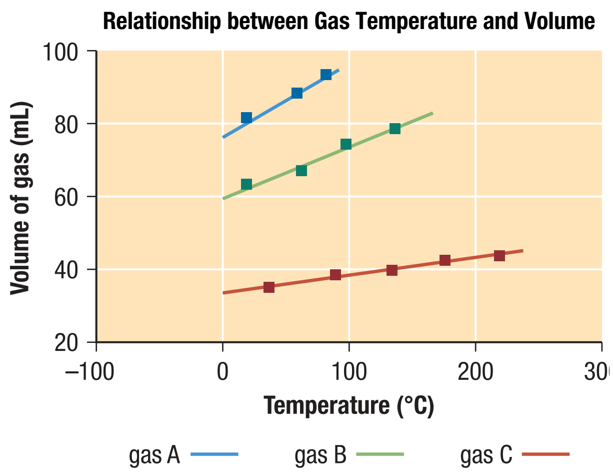


Figure 3

10. You blow up balloons to decorate the gym for a dance. Several of the balloons escape outside where it is −10 °C. Why does the volume of the balloons outside differ from those that stay inside the gym? A

11. Why do you think water does not rush into the open mouth of the swimmer in **Figure 4**?

Figure 4 If you open your mouth under water, water does not fill your mouth and lungs.

Skills Review

12. **Figure 5** summarizes the 2006 and estimated 2020 emissions of greenhouse gases (GHGs) for Canada's provinces and territories. Emissions of GHGs contribute to climate change. T/I
    (a) Identify the top two GHG emitters in 2006. What is the difference in their emissions, in megatonnes?
    (b) Why do you think these two provinces produce the most GHGs?
    (c) Which province is estimated to increase emissions the most by 2020? Justify your choice.

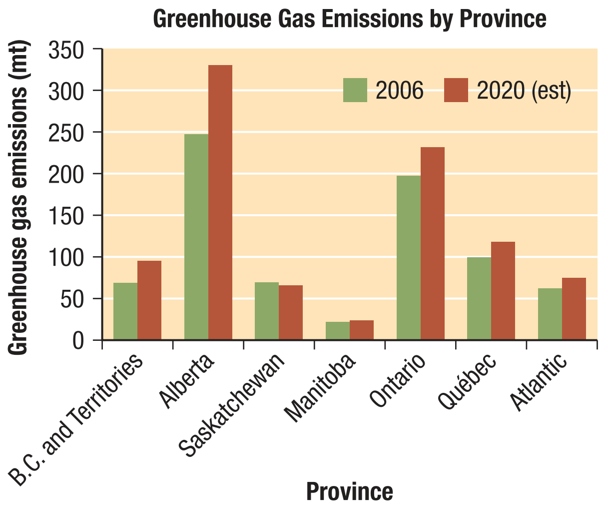


Figure 5

13. Gases are often stored and transported under pressure. They may be flammable and they may be oxidizers. Which WHMIS symbol indicates each of these two hazards? T/I C A

14. Three large balloons are filled with equal volumes of hydrogen, helium, and carbon dioxide, respectively. Use the densities provided in **Table 1** to predict what will happen when the balloons are released into the air. Justify your answer. A

Table 1 Densities of Gases

Gas	Density (g/mL)
balloon 1: hydrogen	0.000 089
balloon 2: helium	0.000 18
balloon 3: carbon dioxide	0.001 98
air	0.001 28

15. Sulfur dioxide is a by-product of fossil fuel combustion and a significant contributor to acid precipitation. K/U T/I C
    (a) Sulfur dioxide reacts with atmospheric oxygen to form sulfur trioxide. Write a balanced chemical equation for the reaction of sulfur dioxide with oxygen.
    (b) Sulfur trioxide then combines with atmospheric water vapour, forming sulfuric acid. Write a balanced chemical equation describing the formation of sulfuric acid.
    (c) What is the molar mass of sulfur dioxide?
    (d) A smelter emits 1.0 t of sulfur dioxide gas. Convert this mass to an amount, in moles.
    (e) What is the maximum amount of sulfur trioxide that can be formed from 1.0 t of sulfur dioxide?
    (f) If all the sulfur trioxide reacts with water to form sulfuric acid, what amount of sulfuric acid is produced?
    (g) Convert the amount of sulfuric acid calculated in (f) from moles to mass.

16. Given the following equation, rearrange the variables to solve for each of these variables: T/I

$$\frac{AB}{C} = \frac{DE}{F}$$

    (a) B
    (b) D
    (c) F

CAREER PATHWAYS **PREVIEW**

Throughout this unit you will see Career Links in the margins. These links mention careers that are relevant to Gases and Atmospheric Chemistry. On the Chapter Summary page at the end of each chapter you will find a Career Pathways feature that shows you the educational requirements of the careers. There are also some career-related questions for you to research.

CHAPTER 11

The Gas State and Gas Laws

KEY CONCEPTS

After completing this chapter you will be able to

- explain the importance of the components of the atmosphere
- analyze the effects of some technologies and human activities on air quality
- describe the Air Quality Health Index and assess air quality based on this index
- outline some effects of poor air quality on human health and the environment
- propose actions to reduce your personal carbon footprint
- describe and evaluate some government initiatives designed to improve air quality
- use appropriate terms and units to describe gases and atmospheric chemistry
- use the kinetic molecular theory to explain the properties and behaviour of gases
- apply the gas laws to solve problems related to the behaviour of gases
- use the gas laws in investigations involving gases

Why Is the Atmosphere So Important?

All that separates us from outer space is a thin layer of gases that we call the atmosphere. Earth's atmosphere extends to more than 100 km above Earth's surface but, compared to the radius of Earth, the atmosphere is quite thin—like the skin of an onion.

Life, as we know it, could not exist on Earth without the atmosphere. The atmosphere traps enough solar energy during each day to prevent us from freezing at night. It protects us from much of the harmful ultraviolet radiation emitted by the Sun. The atmosphere also contains the oxygen necessary to support terrestrial life. Oxygen makes up only about 21 % of the air you breathe. Much of the rest consists of nitrogen.

Earth's atmosphere consists mostly of gases, but also includes some liquids and solids. The composition of the atmosphere has changed significantly during Earth's history. About three billion years ago, there was hardly any oxygen. Instead, frequent volcanic eruptions filled the atmosphere with a toxic mixture rich in methane and sulfur compounds. Later, as volcanic activity on Earth subsided, the first micro-organisms capable of photosynthesis appeared. Recall that photosynthesis is the process in which a plant uses solar energy to transform water and carbon dioxide into large organic compounds such as glucose. Oxygen, the waste product of this reaction, is released into the atmosphere. As the populations of photosynthetic organisms grew, so did the concentration of oxygen in the atmosphere. Over billions of years, the activity of these micro-organisms and, subsequently, plants helped to increase the concentration of atmospheric oxygen to its current level. This made it possible for oxygen-dependent organisms, such as humans, to evolve and flourish. Some of this oxygen also reacted to produce the ozone that is essential for our survival.

Evidence suggests that another significant change in the composition of the atmosphere is currently under way. This change is only a few hundred years old. The cause is air pollutants generated by human activities. These pollutants are changing our climate and affecting human health. Since air is constantly in motion, pollutants generated in one part of the world migrate to other parts. As a result, the efforts of one country to curb its air pollution may be futile without the cooperation of its neighbours.

In this chapter you will examine the characteristics of the atmosphere as well as the properties of gases. You will use this understanding to explore the impact of human activity on air quality and ways in which that impact can be reduced.

STARTING **POINTS**

Answer the following questions using your current knowledge. You will have a chance to revisit these questions later, applying concepts and skills from the chapter.

1. Carbon dioxide and sulfur dioxide have been in the atmosphere for millions of years. Why do you think they have become a significant environmental problem only recently?
2. Give three examples of human activities that affect air quality.
3. Why do you think that initiatives to improve air quality often require the cooperation of many countries before they can be effective?
4. Why do you think poor air quality is bad for both our physical health and the "health" of our economy?

Mini **Investigation**

Inflating Experience

Skills: Predicting, Planning, Performing, Observing, Analyzing

SKILLS HANDBOOK A1.2, A2.1

The air that surrounds us is mostly a mixture of gases. Gases are fluids, just like liquids. This means that gases flow from one place to another in the same way that liquids do. Gases also exert a force as their entities strike a surface. In this activity you will investigate how restrictions to airflow influence how easy it is to inflate a balloon.

Equipment and Materials: balloon; 500 mL plastic bottle

1. Stretch the balloon a few times to ensure that it is flexible.
2. Designate one student to blow in the balloon. Inflate the balloon and then allow the gas to escape.
3. Insert the closed end of the balloon into the bottle. Stretch the opening of the balloon over the mouth of the bottle (**Figure 1**).
4. Have the same student attempt to inflate the balloon.
5. Discuss with your partner what alterations could be made to the bottle to make it easier to inflate the balloon.
6. Proceed, using the additional materials supplied by your teacher.

A. Why was it more difficult to inflate the balloon in Step 4 than it was in Step 1? T/I

B. Explain how your alteration to the bottle helped. T/I

Figure 1

11.1 States of Matter and the Kinetic Molecular Theory

Figure 1 Liquid nitrogen can be used to remove warts.

Have you ever had a wart removed? A wart is a non-cancerous skin growth caused by a viral infection. A wart can be "burned off" professionally by applying a small quantity of liquid nitrogen directly to the wart (**Figure 1**). Nitrogen is normally a gas at room temperature, having a boiling point of −196 °C. However, it can be liquefied if it is cooled below this temperature. As a result, liquid nitrogen is extremely cold—cold enough to instantly freeze the water in each cell of the wart. As the water freezes, it expands, causing the cells in the tissue of the wart to burst—much like a pop can that was left too long in the freezer. As the tissue is destroyed, the patient experiences a burning sensation in the affected area. The liquid nitrogen that was used in the procedure absorbs thermal energy from the wart and harmlessly evaporates to become nitrogen gas. This example illustrates how the three common states of matter are affected by the temperature of their surroundings.

In this section we will review the properties of these states and examine a useful theory that helps explain these properties.

States of Matter

There are three common states of matter: solid, liquid, and gas. Many of the differences between these states can be explained by considering the forces of attraction between entities (**Table 1**). Solids, for example, maintain their volume and shape. This suggests that the forces of attraction between the entities in solids are quite strong. Strong attractive forces also keep the entities in solids close together. This helps to explain why solids are difficult to compress.

Table 1 Three States of Matter

State	Properties	Model	Example
solid	• has definite shape and volume • is virtually incompressible • does not flow easily		
liquid	• takes the shape of the container but has a definite volume • is slightly compressible • flows readily		
gas	• takes the shape and volume of the container • is highly compressible • flows readily		

Liquids also maintain their volume. Unlike solids, however, liquids flow to take the shape of their container. Their ability to flow suggests that the forces of attraction in liquids are not as strong as in solids. Weaker attractive forces imply that the entities in a typical liquid are slightly farther apart than the entities in solids. The combination of weaker attractions and more space makes it easier for the entities to slide past each other as the liquid flows. This explains why liquids can flow and take the shape of their container. It also explains why liquids are slightly more compressible than solids. In Unit 1 you learned that the intermolecular forces of attraction in molecular substances are London forces, dipole–dipole attractions, and hydrogen bonds. Although these attractive forces are weak, their presence prevents liquids from vaporizing into gases.

Gases take the shape and volume of their container. For example, the stench of a partially decomposed banana in a locker quickly fills the school hallway when the locker is opened. Observations like these suggest that the attractive forces are very weak between entities in gases, and the distances between entities are great. This explains why gases are the easiest of the three states to compress (**Figure 2**).

In summary, many of the properties of solids, liquids, and gases can be explained by referring to the strength of the attractive forces between their entities. However, this explanation has its limitations. For example, it cannot explain why a fully inflated balloon shrivels when placed into liquid nitrogen and then expands when it is removed. It also cannot explain why the warmth of a room causes liquid nitrogen to evaporate. To complete our theoretical understanding of the states of matter we must also consider the motion of the entities in matter.

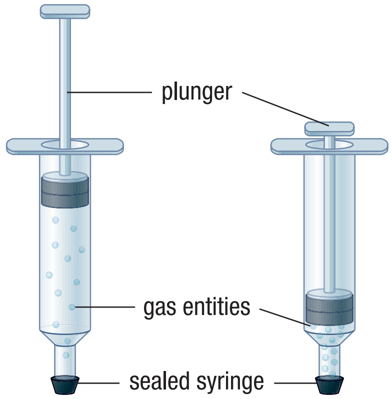


Figure 2 Gases are compressible: they can be squeezed so that they occupy a smaller volume. These two syringes contain the same number of gas entities.

Kinetic Molecular Theory

The aroma of cookies baking in the kitchen quickly fills an entire home. Why does this happen? One of the earliest clues to explain observations such as this came from the work of the Scottish scientist Robert Brown (1773–1858). Brown had been examining small specks of pollen suspended in water under the microscope. He observed that the pollen specks were moving in a random, zig-zag motion, as if they were being continually hit by invisible particles. This observed random movement eventually became known as **Brownian motion** (**Figure 3**).

Brownian motion the random movement of microscopic particles suspended in a liquid or a gas

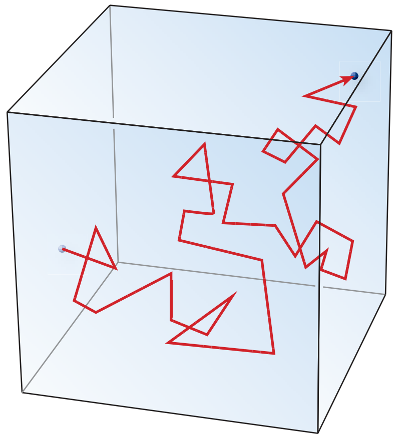


Figure 3 Brownian motion is the continuous random motion first described by Robert Brown.

Scientific interpretations of Brown's observations led to the development of the kinetic molecular theory. The main idea of the **kinetic molecular theory (KMT)** is that the entities in solids, liquids, and gases are in constant, random motion. As entities move about, they collide with one another and any other object in their path. The word "kinetic" comes from the Greek word *kinema*, meaning motion. The motion of these entities explains how the smell of baking cookies reaches your bedroom, some distance away from the kitchen. **Kinetic energy** is the energy of an entity due to its motion.

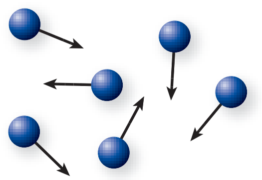
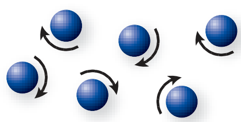
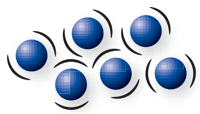
kinetic molecular theory the idea that all substances are composed of entities that are in constant, random motion

kinetic energy energy possessed by moving objects

How Entities Move

There are three ways in which the entities in matter can move: translational motion, rotational motion, and vibrational motion (**Table 2**). The type of motion in a sample of matter depends on the strength of the forces of attraction present.

Table 2 The Three Types of Motion

Translational motion	Rotational motion	Vibrational motion
• the movement of an entity through space along a linear (straight-line) path • in gases and liquids 	• the spinning of an entity in place • in gases and liquids • very limited in solids 	• the back-and-forth vibration of entities • in gases, liquids, and solids 

Strong attractions limit motion. The entities in most solids are therefore limited to vibrational motion. Strong attractions also result in a solid being the most ordered of the states of matter. Liquids exhibit all three forms of motion to some extent because their attractive forces are weaker than those in solids. That is also why the liquid state is less ordered than the solid state. For most practical applications, we may assume that there are no attractions between entities in a gas. As a result, the entities in gases exhibit all three forms of motion. Translational motion, however, is the most significant. Translational motion results in frequent, random collisions between entities in gases. Gas is the least ordered or most jumbled of the three common states of matter. **Table 3** compares the types of motion, strength of attractions, and degree of organization in the three states of matter.

Table 3 Types of Motion, Forces, and Organization of Entities in the Three States of Matter

	Solids	**Liquids**	**Gases**
Types of motion	vibrational	vibrational, rotational, and translational	vibrational, rotational, and translational
Strength of attraction	strongest	intermediate	weakest
Organization of entities	highly organized	intermediate level of organization	least organized

Kinetic Energy and Temperature

You now know that the entities in any substance are in constant motion and that the energy possessed by any moving object is called kinetic energy. When a substance is warmed, its entities move more rapidly. The entities in a solid, for example, vibrate back and forth because of their kinetic energy. As the solid is warmed, its entities remain fixed in their position but vibrate more rapidly. The faster these entities move, the greater their kinetic energy. The kinetic energy of the individual entities may vary but, on average, their kinetic energy increases as the solid is warmed. This increase in kinetic energy makes the substance feel warmer to the touch. **Temperature** is a measure of the average kinetic energy of the entities in a substance. When you measure the temperature of a substance with a thermometer, the entities of the substance collide with the glass of the thermometer. As the substance is warmed, these collisions become more energetic. This additional energy is then transferred to the liquid in the thermometer, causing the liquid to expand (**Figure 4**).

temperature a measure of the average kinetic energy of the entities of a substance

WEB **LINK**

There are online simulations that show molecular motion and its dependence on temperature. To investigate one of these simulations,

GO TO NELSON SCIENCE

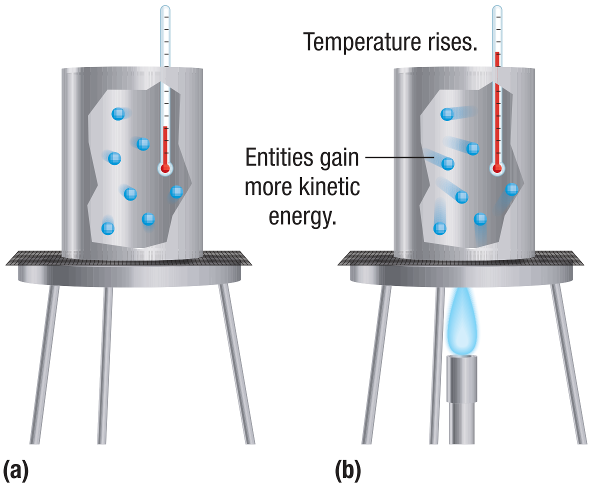


Figure 4 (a) Before heating, the entities have only a little kinetic energy. (b) After heating, they have much more kinetic energy and the thermometer indicates a higher temperature.

As more energy is transferred to a solid, its entities vibrate even faster. Eventually, the attractive forces between the entities are broken, allowing the entities to flow past each other. As a result, a change in state occurs in which the solid turns into liquid.

Warming a liquid also increases the kinetic energy of its entities. If sufficient energy is applied, the attractive forces between the entities in the liquid are completely overcome, allowing the liquid to boil and turn into a gas.

11.1 Summary

- The three common states of matter are solid, liquid, and gas. These three states have different properties, such as the ability to flow or be compressed.
- The kinetic molecular theory states that all substances contain entities that are in constant, random motion.
- Entities in a sample of matter may have translational motion, rotational motion, and/or vibrational motion, depending on their state.
- Entities in solid substances have limited motion, entities in liquids have an intermediate level of motion, and entities in gases have the most motion.
- Temperature is a measure of the average kinetic energy of the entities in a substance.
- As a substance is warmed, the kinetic energy of its entities increases, raising the temperature of the substance. If warming continues, some of the absorbed energy helps to overcome the attractive forces between the entities, resulting in a change of state.

11.1 Questions

1. Explain why solids have a definite shape and volume. K/U
2. Why are gases relatively easy to compress while solids are virtually incompressible? K/U
3. Rank the three common states of matter in order from least ordered to most ordered. How is this ranking related to the attractive forces between the entities in each state? K/U T/I
4. Why do gases and liquids flow, while solids do not? K/U
5. (a) What is the kinetic molecular theory (KMT)?
   (b) Use the KMT to explain the change illustrated in **Figure 5**.
   (c) Use the KMT to explain why the scent of perfume spreads more slowly outdoors in winter than it does in summer. K/U A

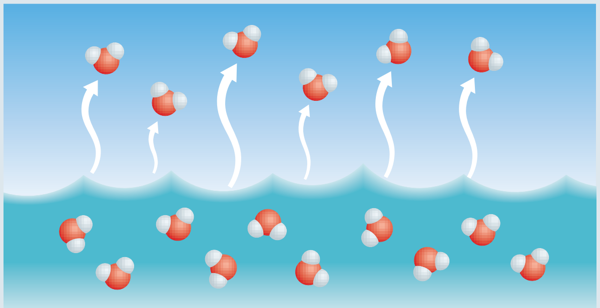


Figure 5

6. (a) How is temperature related to the motion of entities?
   (b) How does heating change the motion and temperature of the entities of a particular substance? K/U
7. Liquid nitrogen can be prepared by cooling and compressing air. Explain what happens to the molecules of nitrogen as it condenses. A
8. Propane is sold as a liquid in pressurized tanks. Use your knowledge of forces of attraction to explain what happens to liquid propane when it is released into the air in the burner of a propane barbecue. A
9. Dry ice is solid carbon dioxide. When exposed to room temperature, dry ice sublimates from a solid directly to a gas. Explain this change of state in terms of energy changes, molecular motion, and forces of attraction. T/I
10. Smoke is a mixture of tiny particles that are suspended in air. T/I A
    (a) Describe the motion of the particles of smoke moving from a smokestack through the atmosphere on a still day.
    (b) Would you expect particles of smoke to disperse faster in the summer or the winter? Explain your answer.
11. Why must car mechanics ensure that there are no air bubbles in the liquid in a vehicle's brake lines? To answer this question, you may need to research hydraulic brake systems. K/U A
12. Research one medical or industrial use of liquefied gases. T/I A
    (a) Briefly explain how the gas is liquefied.
    (b) Explain how the gas is used in this application.

GO TO NELSON SCIENCE

11.2 The Atmosphere and Its Components

The atmosphere is a thin blanket of moisture and gases that protects us from the hostile environment of outer space. Most of the gases in the atmosphere are concentrated in the first 100 km above Earth's surface. Beyond that height, their concentration is quite low. Weather systems in the lower atmosphere help to circulate warm and cold air and distribute precipitation. The atmosphere stabilizes the surface temperature, making Earth habitable. Without it, Earth would experience extremes in daytime and night-time temperatures like those on planets or moons with a little or no atmosphere.

Layers of the Atmosphere

The atmosphere consists of four layers: troposphere, stratosphere, mesosphere, and thermosphere (**Figure 1**). The temperature of these layers varies. The troposphere is the layer closest to Earth's surface and is warmest near Earth's surface. Here, solar energy absorbed by Earth is transferred to nearby air molecules as thermal energy. Air movements in the troposphere help determine our weather.

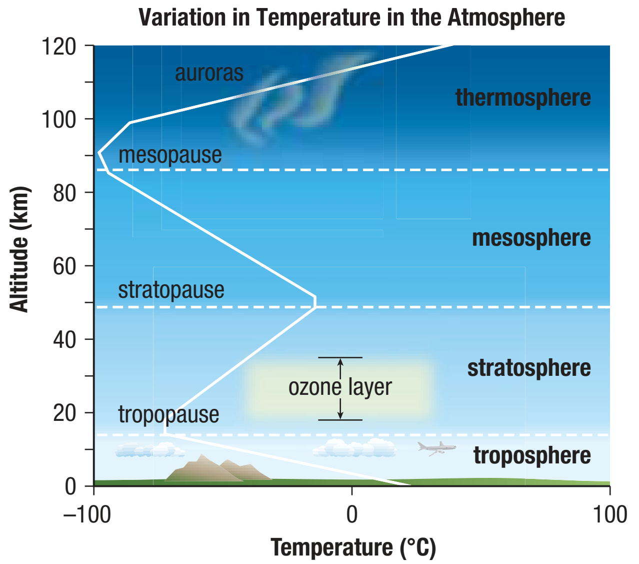


Figure 1 Notice that the temperature does not decrease steadily with increasing altitude.

Next to the troposphere is the stratosphere. This is a dry region that contains a slightly higher than normal concentration of ozone. Ozone has a unique ability to trap energy in the form of ultraviolet (UV) radiation from the Sun. This trapped energy is released to nearby gas molecules, increasing their kinetic energy. This helps to explain why the temperature in the stratosphere increases with altitude rather than decreases.

The middle layer, or mesosphere, extends from about 50 km up to about 85 km. Gas concentrations in this layer are quite low. Since the mesosphere contains little ozone, its temperature decreases with increasing altitude.

The upper layer of the atmosphere—the thermosphere or ionosphere—extends beyond 85 km. Gas molecules in this region absorb high-energy solar radiation and become ionized. In the process, they emit visible radiation in the form of spectacular displays called auroras (**Figure 2**). This release of energy also helps to warm the thermosphere. The auroras occur particularly over polar regions and are know as the aurora borealis, or northern lights, and the aurora australis, or southern lights.

Figure 2 Auroras are natural light displays that appear over polar regions. Interactions with oxygen entities form a yellow-green colour, or occasionally red, whereas collisions with nitrogen molecules result in blue or purple.

The atmosphere is most dense near the surface of Earth. Here, the molecules in air are pushed together by the atmosphere above them. The atmosphere becomes less dense as the height above sea level, or altitude, increases. Decreasing air density implies that the oxygen concentration also decreases with increasing altitude. The lack of oxygen makes strenuous high-altitude activities, such as mountain climbing, very tiring. Commercial aircraft flying at high altitudes have to be pressurized so that the interior air pressure and oxygen concentration are similar to those on Earth's surface.

Composition of the Atmosphere

The two atmospheric layers closest to Earth, the troposphere and the stratosphere, have the biggest influence on us. Nitrogen and oxygen are by far the most common gases in these layers, followed by argon, carbon dioxide, and other trace gases (**Table 1**).

Table 1 Components of Dry Air at Sea Level

Gas	Percentage composition by volume	Concentration (ppm)
$N_2(g)$	78.08	7.808×10^5
$O_2(g)$	20.95	2.095×10^5
Ar(g)	0.934	9.34×10^3
$CO_2(g)$	0.036	3.6×10^2
Ne(g)	0.001 818	18.18
He(g)	0.000 524	5.24
$CH_4(g)$	0.000 2	2
Kr(g)	0.000 114	1.14

Nitrogen

The most abundant gas in the atmosphere is nitrogen, $N_2(g)$. The element nitrogen is important in all biological systems. It is found in the proteins and DNA of plants and animals. Nitrogen cycles through Earth in the nitrogen cycle. Nitrogen-fixing bacteria convert nitrogen gas into soluble nitrates—a form that plants can easily absorb through their roots. Nitrogen is released back into the atmosphere from the bodies of decaying plants and animals or their wastes (**Figure 3**).

Figure 3 All life is dependent on nitrogen and oxygen.

Oxygen and Ozone

There are two forms of elemental oxygen in the atmosphere: the familiar diatomic form, $O_2(g)$, and the much rarer form with 3 atoms, $O_3(g)$, known as ozone.

Oxygen makes up 21 %, by volume, of the atmosphere. Our cells require oxygen to conduct cellular respiration—a chemical process that releases the energy stored in large organic compounds like glucose. This energy is used to fuel numerous processes that are essential for life.

Most of the ozone in the atmosphere is concentrated in the stratosphere. Even there, though, its concentration may reach only 10 ppm. Ozone absorbs UV light quite effectively. Because of this property, ozone acts as a radiation shield, preventing excessive UV radiation from reaching Earth's surface. UV exposure can damage DNA and increase the risk of cancer. Some ozone is also found in the troposphere where it is an air pollutant and a key ingredient in smog (Section 11.4).

Other Gases

Argon, water vapour, and carbon dioxide make up the majority of the remaining gases in the atmosphere. Argon is colourless, odourless, and chemically inert. Water vapour is approximately 1 % to 3 % of the atmosphere by volume. Over 99 % of all water vapour is in the troposphere where it is a component of the water cycle. Carbon dioxide is only 0.036 % of the volume of the lower atmosphere but is essential for life on Earth. Green plants and algae use carbon dioxide to manufacture energy-rich molecules, such as glucose, during photosynthesis. Carbon dioxide is returned to the atmosphere by living organisms as a waste product of cellular respiration.

The Greenhouse Effect and Climate Change

Earth's atmosphere provides more than just the gases necessary to support life. It also helps warm Earth's surface by trapping thermal energy radiated by the ground. Scientists estimate that, without the atmosphere, Earth's average temperature would be a frosty −18 °C. Tucked within its atmospheric "blanket," Earth's average temperature is 15 °C—and rising! We can take a closer look at how this happens. Earth's atmosphere is transparent to much of the higher-energy radiation from the Sun (such as visible light). Much of this radiation is absorbed by Earth's surface and transformed into thermal energy. Warm objects on Earth's surface then give off this energy as lower-energy infrared (IR) radiation. The warmth of a fireplace is an example of IR radiation.

A greenhouse gas (GHG) is a gas in the atmosphere that traps infrared radiation, contributing to the greenhouse effect. Carbon dioxide, methane, and water vapour are the most important GHGs. Other, less significant GHGs include nitrous oxide and fluorinated gases such as sulfur hexafluoride. Greenhouse gases are particularly effective at absorbing certain energies of IR radiation. Just as there are certain frequencies of sound that can cause a glass to vibrate, specific frequencies of IR radiation cause greenhouse gases to vibrate. As they do, they re-radiate this energy. Some of it is sent back to Earth's surface where it warms the ground. The process of trapping infrared radiation within the atmosphere is called the **greenhouse effect** (**Figure 4**).

greenhouse effect a natural process whereby gases and clouds absorb infrared radiation emitted from Earth's surface and radiate it, heating the atmosphere and Earth's surface

UNIT TASK **BOOKMARK**

The information about atmospheric carbon dioxide in this section will be useful as you work on the Unit Task (page 616).

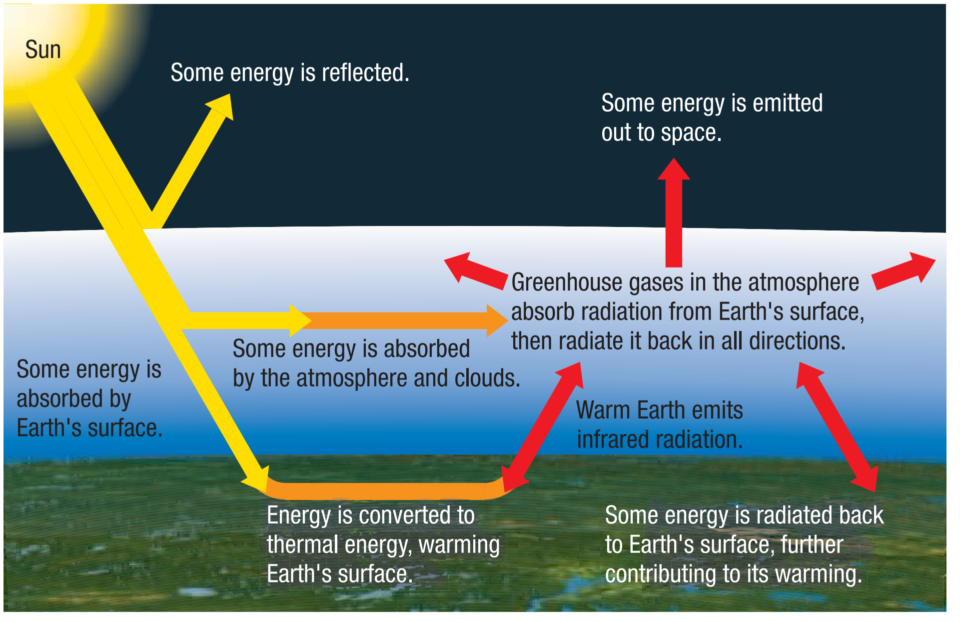


Figure 4 The greenhouse effect is a natural phenomenon that makes life on Earth possible.

There have been GHGs in Earth's atmosphere for millions of years. Furthermore, the concentrations of these gases have fluctuated over time. Scientists investigating these natural fluctuations have found that higher concentrations of GHGs tend to correspond to periods of higher global temperatures (**Figure 5**). Events such as volcanic emissions, continental drift, changes in solar energy or ocean circulation, and strikes by large meteors are associated with changes in global temperature.

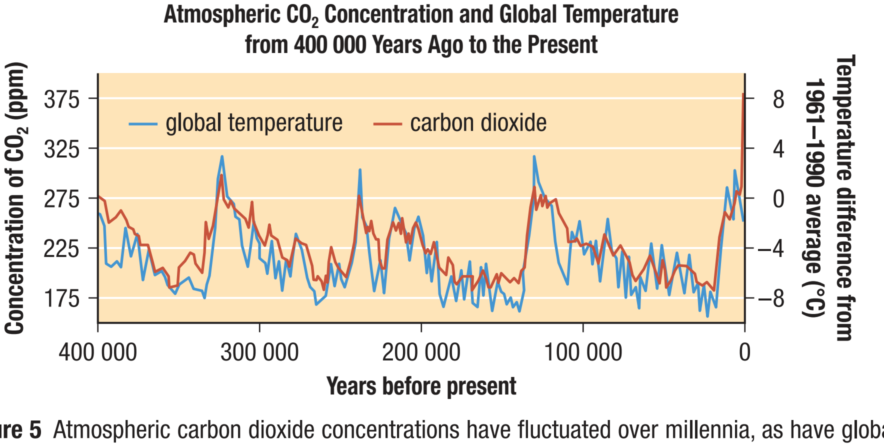


Figure 5 Atmospheric carbon dioxide concentrations have fluctuated over millennia, as have global temperatures.

In the last 200 years, human activity has significantly increased the concentrations of GHGs in the atmosphere. Eighty percent of Canada's total GHG emissions result from the production or use of fossil fuels. Canada produces 2 % of the total annual GHG emissions worldwide. The greater the concentration of these gases, the greater the greenhouse effect.

In the past decades, climate change has begun to occur because the increased levels of greenhouse gases are enhancing the natural greenhouse effect. Climate change is caused by both natural and human processes that change the atmosphere's chemical composition. Studies show that the increase in human-produced GHGs is the main cause of climate change.

Climate change will have a significant effect in Canada. Scientists predict that water levels in the Great Lakes could drop 30 to 80 cm. This would also adversely affect the recreation and shipping industries. It is predicted that Canada's temperate vegetation could spread 160 to 640 km north into the boreal forest over the next 100 years. The Arctic could experience a decrease in sea ice and degradation of permafrost. Many of these changes are already occurring.

Reducing Greenhouse Gas Emissions

What can be done about GHGs? Government, the business sector, and individuals are all concerned about GHG emissions. The Government of Canada is committed to reducing greenhouse gas emissions by 17 % between 2005 and 2020. Some provincial governments enforce emission standards for vehicles, and provide funds for farmers and businesses to invest in environmentally friendly practices. Businesses are reducing packaging and transporting their products more efficiently.

Each of us can begin in our own homes to reduce carbon dioxide emissions. Recycling and reusing products decreases the quantity of carbon dioxide produced in manufacturing and transportation. Driving less, using public transport, and carpooling all reduce the consumption of fossil fuels. Insulating, caulking, and weatherstripping around windows and doors can reduce the use of heating fuel.

Capturing the carbon dioxide is another strategy to limit the concentration of carbon dioxide in the atmosphere. The process of removing carbon dioxide from the atmosphere and storing it is called **carbon sequestration**. Two common types of sequestration are biological and geological (**Figure 6**). In biological sequestration, plants are used to naturally sequester or remove carbon dioxide. Planting trees, for example, is a simple and familiar carbon sequestration strategy.

WEB **LINK**

To find out more about climate change and what the Canadian government is proposing to do about it,

GO TO NELSON SCIENCE

carbon sequestration the process of removing carbon dioxide and other forms of carbon from the atmosphere, and then storing it

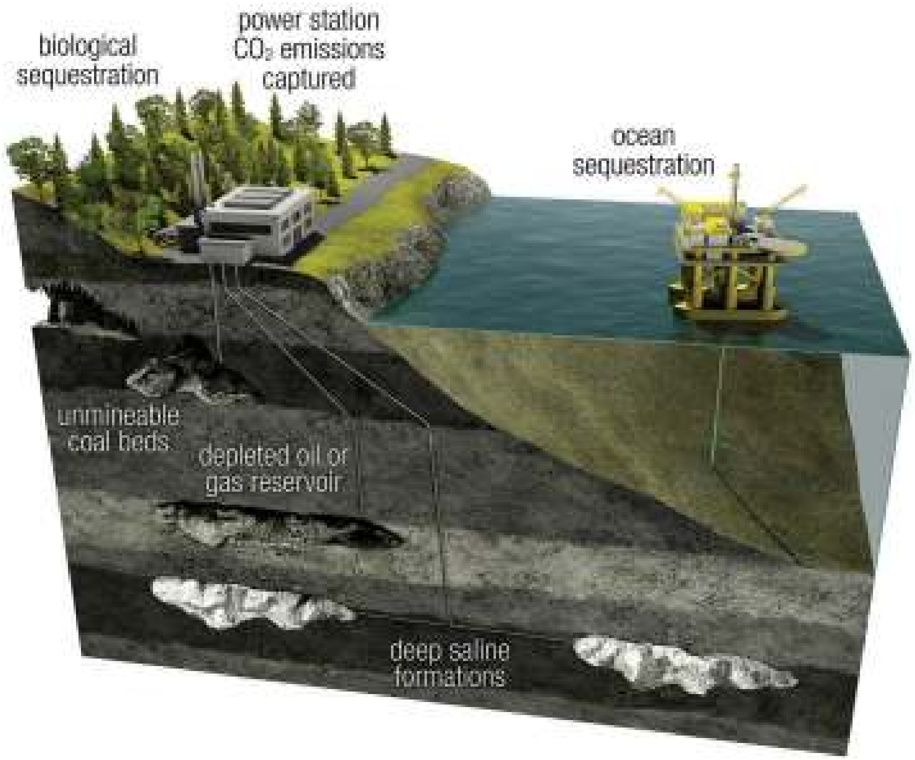


Figure 6 Carbon sequestration technologies include biological sequestration, such as planting trees, and geological sequestration, such as pumping carbon dioxide into depleted reservoirs.

UNIT TASK **BOOKMARK**

The information on carbon sequestration will be useful as you work on the Unit Task (page 616).

Geological sequestration involves capturing and storing carbon dioxide in underground deposits. Some forms of geological sequestrations have costs associated with them. Carbon dioxide may leak back into the atmosphere or into groundwater. In groundwater, carbon combines with water to form acids which then release minerals into drinking water. In addition, sequestration is costly because the carbon dioxide must be compressed into a liquid, which requires energy.

F.K. Hare—A Pioneer in Climate Change

Figure 7 Frederick Kenneth Hare (1919–2002)

Frederick Kenneth Hare was born in England in 1919 (**Figure 7**). He immigrated to Canada after working as a meteorologist during World War II. He completed his Ph.D. at the Université de Montréal, studying Arctic climatology.

By the 1960s, Hare had already expressed concern over the link between GHG emissions and climate change. During his career, Hare wrote and spoke extensively to raise awareness of issues related to climate change, winning numerous awards for his efforts. He was a Companion of the Order of Canada (1987) and a recipient of the Order of Ontario (1989). In 1989, he was awarded the International Meteorological Organization Prize from the United Nations.

CAREER LINK

To find out more about the work of a climatologist,

GO TO NELSON SCIENCE

Research This

Better Lawn Care

Skills: Researching, Analyzing, Evaluating, Communicating, Identifying Alternatives — SKILLS HANDBOOK A5.1

Gasoline-powered lawn mowers groom lawns efficiently, but they emit greenhouse gases. There are alternatives to using gasoline-powered lawn mowers, including using push mowers and replacing lawns with naturalized gardens (**Figure 8**).

Figure 8 Using a push mower can reduce greenhouse gas emissions.

1. Research the regulations regarding lawn care in your municipality.

A. Briefly summarize the regulations. A

B. Suggest at least two alternatives for urban lawn care. T/I A

C. List at least one potential benefit and one potential problem with each of these options. T/I A

D. Gasoline-powered lawn mowers emit a great deal of air pollution. In your opinion, how feasible would it be to impose a bylaw prohibiting the use of gasoline-powered lawn mowers? Would such a bylaw be fair? Justify your opinion. C A

GO TO NELSON SCIENCE

11.2 Summary

- The atmosphere is a mixture of gases and moisture that surrounds Earth.
- The atmosphere has four layers: troposphere, stratosphere, mesosphere, and thermosphere. The density, temperature, and composition of each layer vary.
- The atmosphere stabilizes the temperature of Earth, preventing large extremes in daytime and nighttime temperatures from occurring.
- The greenhouse effect is a naturally occurring process in the atmosphere that warms Earth by preventing some infrared radiation from escaping into space.
- Climate change is caused by natural and human processes that alter the atmosphere's composition. The increase in emissions of greenhouse gases is the primary cause of climate change.

11.2 Questions

1. Why are oxygen masks immediately deployed on passenger aircraft if the cabin pressure drops? A
2. Why does the temperature of the atmosphere vary as altitude increases? K/U T/I
3. How is carbon dioxide naturally recycled? K/U
4. (a) Describe the energy flow that occurs during the greenhouse effect.
   (b) What role do greenhouse gases play in reducing the loss of energy radiated from Earth's surface?
   (c) What effect has human activity had on the natural greenhouse effect? Why? K/U A
5. Suggest three ways you can reduce your carbon dioxide emissions in your daily living. A
6. How do you think buying and eating locally grown food might help to reduce greenhouse gas emissions? A
7. Refer to a map to locate the following places. Predict what effect climate change might have on the people who live in these cities. Justify your prediction. K/U T/I A
   (a) Thunder Bay, Ontario
   (b) Iqaluit, Nunavut
   (c) Texel, the Netherlands
   (d) Zhengzhou, China
8. The shape of snowflakes is affected by the temperature, water vapour content, and other atmospheric conditions (**Figure 9**). Research the possible link between snowflake crystal formation and tropospheric ozone levels. T/I

Figure 9

9. (a) What is carbon sequestration?
   (b) Research and briefly report on an example of geological carbon sequestration. Based on your research, identify one question or concern you have about the use of this technology. K/U T/I
10. Scientists studying greenhouse gases have recently discovered trifluoromethyl sulfur pentafluoride, CF_3SF_5. Although is it not intentionally produced, concentrations of CF_3SF_5 have been steadily increasing since 2000. Research trifluoromethyl sulfur pentafluoride. Find out how its energy-trapping effect compares with that of carbon dioxide, how long it persists in the atmosphere, and how it is produced. Summarize your findings in a poster or a web page entitled "Greenhouse Supergas." T/I C
11. In Canada, the transportation sector is the largest source of greenhouse gas emissions. For every litre of fuel your car burns, approximately 2.4 kg of carbon dioxide gas is produced. The average engine burns 3.5 L of gasoline per hour while idling. T/I A
   (a) Research current gasoline prices.
   (b) Determine the "cost" in dollars and in CO_2 production of idling your car for 10 minutes.
   (c) What is the cost of idling for 10 minutes, 5 times per week, over an entire year?
   (d) In your opinion, which is the most significant: the "dollar cost" or the "environmental cost" of idling a vehicle? Why?
   (e) Most car idling occurs in cities. Does that mean that any environmental impacts of idling are restricted to cities? Explain your answer, including examples.

GO TO NELSON SCIENCE

The Road to Discovering Argon

ABSTRACT

Argon is the third most abundant gas in our atmosphere, but its identity remained elusive until almost 1900. How do you detect something that is invisible, odourless, and rarely reacts? The first clue to the existence of argon could easily have been dismissed as experimental error. Lord John Rayleigh had been trying to determine the density of nitrogen. He used two different sources for his nitrogen: air and nitrogen-containing compounds. He expected both batches of nitrogen to have the same mass. Instead, he found that nitrogen from the atmosphere was always slightly heavier than nitrogen made from compounds. He collaborated with Sir William Ramsay and the two scientists realized that the extra mass was a new element. They had discovered argon!

Argon is a colourless, odourless, noble gas. Because of its inertness it has many industrial and medical uses. It is used in neon light tubes, where it glows purplish-blue, as well as in fluorescent and incandescent bulbs. Argon lasers are used to treat eye conditions such as glaucoma or retinal detachment, skin conditions such as port-wine stains (hemangioma), and tumours (**Figure 1**). Argon is also used to preserve important documents as it displaces oxygen- and moisture-laden air.

Figure 1 Argon can be used in laser surgery to treat port-wine stains, a type of birthmark.

The First Clues

The road to the discovery of argon began in the early 1880s in Britain where Lord John Rayleigh was conducting research to determine the densities of hydrogen and oxygen. To determine density, he needed to find the mass of a fixed volume of gas. Determining volume was easy because gases take the volume of their container. Determining the mass of a gas is far more difficult. You cannot just put a sample of gas on a scale because gases are buoyant in air. To overcome buoyancy and other complications, Rayleigh had to develop new procedures and specially designed scales in order to make very precise mass measurements. After three years of painstaking work, he finally published the densities of oxygen and hydrogen in 1888.

Producing Nitrogen

Rayleigh then used this expertise to measure the density of nitrogen. He chose ammonia, NH_3, as his source of nitrogen. He bubbled pure oxygen through liquid ammonia and then through a red-hot tube. In the process, oxygen reacted with ammonia:

$$3\ O_2(g) + 4\ NH_3(l) \rightarrow 6\ H_2O(g) + 2\ N_2(g)$$

After chemically removing the water vapour, Rayleigh had a highly pure sample of nitrogen gas. He then determined the density of this gas. Rather than publishing his results immediately, Rayleigh thought it would be best to verify his results by using another method of producing nitrogen. He could have generated nitrogen from another chemical reaction. Instead, Rayleigh chose to extract nitrogen from the air using a well-documented procedure.

At the time, air was thought to consist mainly of nitrogen, oxygen, and tiny concentrations of other gases such as carbon dioxide and water vapour. Rayleigh believed that if you could chemically remove oxygen from air, the gas that remained would be essentially nitrogen. To remove oxygen from air, he passed air over hot copper metal to produce copper(II) oxide:

$$O_2(g) + Cu(s) \rightarrow 2\ CuO(s)$$

Since the volume of gas he collected using this method was the same as in the earlier ammonia method, Rayleigh expected this sample of "nitrogen" to have the same mass as well. Instead,

> "[The gas] proved to be one-thousandth part heavier than that obtained by the ammonia method; and on repetition, that difference was only brought out more clearly." (Rayleigh, 1895, p. 702)

The Discrepancy

Although "one-thousandth part heavier" may not sound like much, Rayleigh was so confident in his work and his equipment that he knew the discrepancy was not a result of experimental error. His inability to explain this extra mass clearly frustrated Rayleigh, who said,

"[It] puzzled me very much, and which, at that time, I regarded it only with disgust and impatience." (Rayleigh, 1895, p. 702)

Further Testing

Rayleigh then wondered whether the methods he used to make nitrogen were the source of the problem. To test this, he extracted nitrogen from the air using three different methods. He also produced nitrogen from four different chemical compounds. The nitrogen extracted from the atmosphere was consistently heavier (**Table 1**).

Table 1 Mass of Identical Volumes of Nitrogen

Average mass of "atmospheric nitrogen" (g)	Average mass of "chemical nitrogen" (g)
2.3102	2.2990

The difference between the averages is about 0.0112 g or 11.2 mg—enough to prove that atmospheric nitrogen contained an unknown gas.

Isolating the Mystery Gas

Rayleigh had frequently discussed his work with the British chemist Sir William Ramsay, who had also been studying gases. In an attempt to isolate the mystery gas, Ramsay allowed "atmospheric nitrogen" to react with hot magnesium to produce magnesium nitride:

$$3\ Mg(s) + N_2(g) \rightarrow Mg_3N_2(s)$$

When all the nitrogen was removed, a tiny quantity of gas remained (**Figure 2**). Further analysis suggested that this gas was an element that had not previously been described. Ramsay determined that the atomic mass of this mystery element was 39.9 u. This placed the gas partway between chlorine and potassium on the periodic table. Rayleigh and Ramsay named this element argon. The name "argon" means "the inactive one".

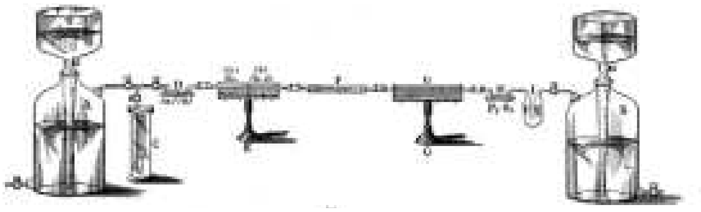


Figure 2 The apparatus for isolating argon from air, as published in Ramsay's book, *The Gases of the Atmosphere*

The First Noble Gas—Argon

The discovery of argon posed a serious problem for chemists—they had no place to put it on the periodic table. Ramsay suggested that another column be added to accommodate it. Many chemists, including Dmitri Mendeleev, were initially furious with this idea. Mendeleev, if you recall, had already published his periodic table in 1869. But Ramsay's idea soon caught on. Within five years the remaining elements of the noble gas family were discovered and joined argon on the periodic table.

In recognition for their research on gases and the discovery of argon, Rayleigh was given the Nobel Prize for Physics in 1904, and Ramsay was given the Nobel Prize for Chemistry.

Further Reading

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Cobb, C., & Goldwhite, H. (1995). *Creations of fire: Chemistry's lively history from alchemy to the atomic age*. Cambridge, MA: Perseus Publishing.

Lord Rayleigh, & Ramsay, W. (1895). "Argon, a new constituent of the atmosphere." *Philosophical Transactions of the Royal Society of London*. A 186: 187–241.

GO TO NELSON SCIENCE

11.3 Questions

1. Why did Rayleigh choose to make nitrogen using a different method before publishing his data? T/I
2. What evidence suggested that the nitrogen Ramsay extracted from air was not pure? K/U
3. Why do you think Rayleigh was confident that the discrepancy in mass observed was not experimental error? T/I
4. What evidence indicated that the larger mass of nitrogen derived from the atmosphere was not due to an experimental error? K/U T/I
5. Why do you think Ramsay's idea to extend the periodic table met with so much opposition? A

11.4 Air Quality

Imagine living at a time when the air was so corrosive it hurt to breathe; when simply inhaling was hazardous to your health. This is not a scene from a science-fiction movie. There have been many times in our history when air quality posed a serious threat to human health. For example, in 1952, a thick layer of smog blanketed London, England, killing more than 4000 people (**Figure 1(a)**). Smog is a complex mixture of air pollutants resulting mostly from human activity. The London smog of 1952 contained a great deal of sulfur dioxide and fine particles. These pollutants came from coal with a high sulfur content that was burned as a home-heating fuel. Smog over Beijing, China, was a concern during the 2008 Olympics (**Figure 1(b)**). Prior to the games, the Chinese government spent billions of dollars to improve air quality. Some of these efforts included closing or relocating factories that were major sources of air pollution and banning car traffic in some parts of the city.

Figure 1 (a) Killer smog in London, 1952. (b) Air pollution caused problems at the 2008 Olympic Games in Beijing. Beijing is the 13th most polluted city in the world. Sixteen of the world's 20 most polluted cities are in China.

The smog experienced in Beijing and many cities in North America is sometimes called photochemical smog. **Photochemical smog** contains ground-level ozone, gases, and fine particles. These pollutants are produced by the reaction of vehicle and factory emissions with sunlight. The prefix *photo-* in "photochemical" means "light." Sunlight is important in the formation of photochemical smog, which often appears as a brown haze over a city.

photochemical smog a hazy cloud of air pollutants formed by the reaction of emissions of factories and vehicles with sunlight

Exposure to smog can cause eye and throat irritation and aggravate a pre-existing respiratory problem like asthma. Poor air quality is a significant health concern in Ontario. In 2008, the Ontario Medical Association (OMA) reported that poor air quality was a contributing factor in more than 9000 premature deaths in Ontario. Thousands more people become sick each year with respiratory conditions aggravated by poor air quality. Furthermore, the OMA estimates that poor air quality costs the Ontario economy more than a billion dollars annually in hospital admissions, medication, and absenteeism from work.

Many substances in the atmosphere contribute to poor air quality. Most of these substances can be classified as particulate matter, gases, or volatile organic compounds (VOCs).

Particulate Matter

When a forest fire occurs, more deaths are caused by smoke inhalation than by burns (**Figure 2**). Smoke inhalation occurs when you breathe in combustion products. These are usually gases and particulate matter. **Particulate matter** is a mixture of solid and liquid particles found in the atmosphere. The size of these particles ranges from just a few molecules to clearly visible dust particles. Although the most obvious sources are smoky fires and poorly tuned vehicle engines, almost all combustion reactions produce some particulate matter. Industrial processes such as the smelting of mineral ores also produce particulate matter (**Figure 3**). About half of the fine particulate matter (<2.5 μm in diameter) in Ontario originates in the United States.

Figure 2 Wildfires produce particulate matter and toxic gases.

particulate matter the mixture of very small solid and liquid particles found in the atmosphere

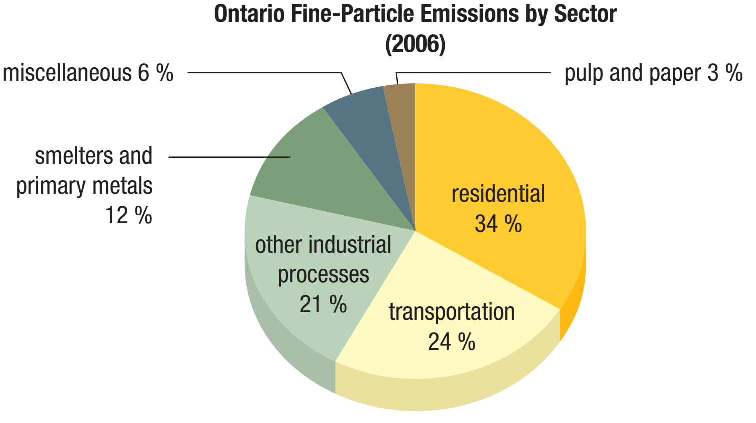

Figure 3 Ontario sources of fine-particle emissions

Particulate matter can damage the respiratory system. Microscopic dust particles, for example, inhaled deep into the lungs can irritate sensitive tissue and damage the lungs.

Particulate matter also reduces visibility and can threaten air travel. In April of 2010, air travel over much of Europe was halted as volcanic ash drifted over U.K. airspace. The source of this volcanic ash was the Eyjafjallajökull volcano in Iceland, nearly 2000 km away (**Figure 4**). Winds carried this ash across Europe and disrupted the plans of nearly 7 million travellers worldwide. Air traffic in the U.K. ceased for nine days and was interrupted in many other countries. Volcanic ash is abrasive and can damage engine parts. Intense heat can fuse ash to the interior of the engine, causing the engine to fail. The instrumentation used to determine air speed can also be disrupted. Radar can detect the ash only if it has a significant moisture component. Pilots cannot avoid volcanic ash if they cannot detect it.

Figure 4 The Eyjafjallajökull volcano in Iceland erupted in 2010.

Pollutant Gases

Numerous gases contribute to atmospheric pollution. Most of the worst gas pollutants are produced through human activities. Once emitted, gases may react with other compounds in the air to produce secondary products.

Sulfur Dioxide

Sulfur dioxide, SO_2, is a clear, colourless gas with a strong, choking odour. More than 2×10^6 t of sulfur dioxide was released in Canada in 2005 by industrial sources. Much of this gas is produced as a result of the combustion of fossil fuels, such as coal and natural gas, that contain sulfur impurities (**Figure 5**). As you learned in Chapter 5, sulfur dioxide emissions contribute to acid precipitation.

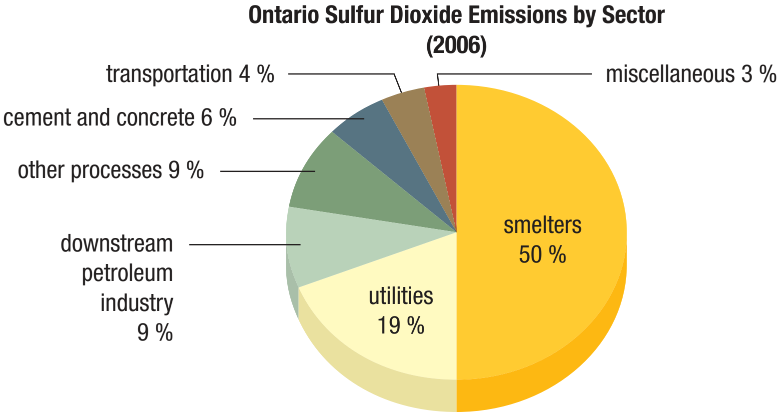

Figure 5 The mining and smelting industries and electrical power generation are major contributors to sulfur dioxide emissions in Canada.

Sulfur dioxide can cause respiratory irritation, stimulate mucus production, and lead to wheezing and shortness of breath. People with respiratory difficulties, including asthma, are especially susceptible to the effects of sulfur dioxide.

Nitrogen Oxides

While driving into an urban centre, you may see brownish smog hovering over the city. This haze is the result of nitrogen dioxide, NO_2. Nitrogen dioxide is one of several nitrogen oxides that pollute air. Others include nitrogen monoxide, NO, and dinitrogen tetroxide, N_2O_4. In Chapter 5 you learned that most nitrogen oxide emissions result from vehicles that use fossil fuels. Like sulfur dioxide, nitrogen oxides contribute to acid precipitation.

Carbon Monoxide

Carbon monoxide, CO, is a colourless, odourless gas that is toxic in high concentrations. Carbon monoxide results from natural events including volcanic activity and forest fires. Many human activities result in carbon monoxide production. As you learned in Chapter 5, incomplete combustion reactions produce both carbon monoxide and pure carbon. In Ontario, the majority of carbon monoxide released into the atmosphere (approximately 85 %) comes from the transportation sector (**Figure 6**).

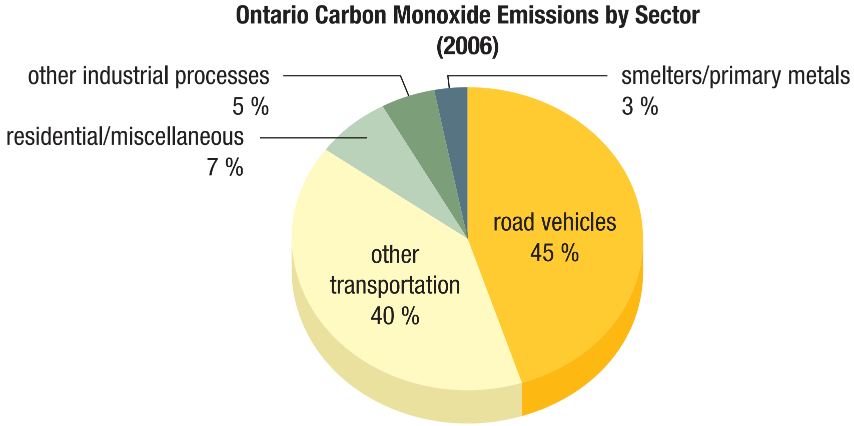

Figure 6 Transportation produces most of the carbon monoxide emissions in Ontario.

The concentration of carbon monoxide can be particularly high in urban areas. During rush hour on windless days, the concentration can reach 50 to 100 ppm. Any level above about 25 ppm can affect health in the long term, even though there may be no noticeable symptoms. However, outdoor carbon monoxide concentrations rarely reach dangerous levels. As you will learn in the next section, indoor carbon monoxide is a greater risk to human health.

Volatile Organic Compounds (VOCs)

Figure 7 The VOCs from automotive paint can be toxic. An airtight mask and respirator prevent this worker from inhaling toxic fumes.

Volatile organic compounds (VOCs) are another class of air pollutants. VOCs are solid or liquid carbon-containing compounds that vaporize readily. Gasoline is a common example of a VOC. When you fill up a car, you may notice ripples in the air above the opening to the gas tank. This observation plus the odour of gasoline indicate that gasoline vapour is leaving the tank. Many VOCs are naturally occurring. The distinctive odour of a pine tree, for example, is due to a volatile hydrocarbon called alpha-pinene. Consumer products such as aerosol sprays, paints, air fresheners, dry cleaned clothing, wood preservatives, and cleaning solvents are also important sources of VOCs (**Figure 7**). The symptoms of VOC exposure vary greatly because so many compounds are VOCs. However, some symptoms include eye irritation, headaches, and rashes. In larger doses VOCs can depress the central nervous system and cause cancer. As you will see later in this section, VOCs also play an important role in the formation of smog.

Ozone

Ozone can be both harmful and beneficial. In Section 11.2 you learned that ozone in the stratosphere absorbs harmful ultraviolet (UV) radiation from the Sun. This prevents much of the UV radiation from reaching Earth's surface. Ground-level ozone, however, is a significant health hazard. Ozone is a powerful bleach. It is so corrosive to biological tissue that some municipalities use it to disinfect their drinking water.

Ground-level ozone is produced from the reaction of nitrogen oxides and volatile organic compounds (VOCs) in the presence of sunlight. Nitrogen oxides, if you recall, are produced when fuel is burned in air at the high temperatures of an internal combustion engine. Under these conditions, nitrogen first reacts with oxygen:

$$N_2(g) + O_2(g) \rightarrow 2\ NO(g)$$

Once released into the atmosphere, colourless nitrogen monoxide reacts with more oxygen to form the reddish-brown pollutant nitrogen dioxide:

$$\underset{\text{(colourless)}}{2\ NO(g)} + O_2(g) \rightarrow \underset{\text{(reddish brown)}}{2\ NO_2(g)}$$

Nitrogen dioxide itself is quite toxic but, with the help of UV radiation and oxygen, it reacts to form ozone in a two-step process. First, nitrogen dioxide absorbs UV radiation to decompose, releasing an oxygen atom:

$$NO_2(g) \xrightarrow{\text{UV radiation}} NO(g) + O(g)$$

A single oxygen atom is highly reactive. It quickly reacts with a nearby oxygen molecule to form ozone:

$$O(g) + O_2(g) \rightarrow O_3(g)$$

VOCs are not directly involved in the formation of ground-level ozone. However, current research suggests that VOCs help to produce nitrogen dioxide in the atmosphere, which then reacts to form ozone. The chemical reactions involved are quite complex, but most are powered by the energy of sunlight.

As you can see, solar energy plays a key role in the production of ground-level ozone. Not surprisingly, the concentration of ground-level ozone tends to peak during a sunny day around noon, when the sunlight is most intense. Ground-level ozone, together with particulate matter and other pollutants, form the brownish haze that we call smog.

Research This

The National Commuter Challenge

Skills: Researching, Analyzing, Evaluating, Communicating, Defining the Issue — SKILLS HANDBOOK A5.1

Using public transportation instead of private vehicles is one way to reduce air pollution (**Figure 8**). Ontario has a number of initiatives to promote the use of public transit. You will now have an opportunity to explore and evaluate some of these initiatives.

Figure 8

1. Research the Commuter Challenge or another Canadian initiative to reduce greenhouse gases. Choose an initiative for which results are available online.
2. Review the results for Ontario.

A. Describe the “who, what, where, when, and why” of the initiative you have chosen. Also include how it is scored. K/U

B. List as many benefits of sustainable transport as you can. Group the benefits in the following categories: Health and Personal; Economic; and Environmental. K/U

C. Describe the success of the initiative, considering questions such as the following: K/U T/I
   - Does the Commuter Challenge attract mostly workplace participants or individual participants? Try to explain why.
   - Which three communities in Ontario had the highest/lowest participation rates?

D. Do you think the program is successful? Do you think that it helps to improve air quality? Explain. T/I A

E. Is the initiative running where you live? If so, how could you and your family participate? If not, outline how you could set it up in your location. T/I A

F. What, if anything, prevents you and your family from using public transit more often? A

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The Air Quality Health Index

It is a beautiful summer day outside but you have been told to stay indoors and close the windows because the air is too polluted. Who decides that the air is too polluted for outdoor activities? How is this decision made?

Every year, thousands of deaths in Canada are attributed to air pollution. At-risk populations include people with existing cardiovascular or respiratory problems, those active outdoors, the elderly, and the very young. Environment Canada developed the **Air Quality Health Index (AQHI)** to help Canadians assess the risk of health effects resulting from air pollution. The scale is based on the health risk of levels of common air pollutants including ground-level ozone, particulate matter, and nitrogen dioxide. The concentrations of these pollutants are measured and combined into a single number representing air quality. The scale goes from 1 to 10+. The values are grouped into categories representing low, moderate, high, and very high health risks (**Figure 9**).

Air Quality Health Index (AQHI) a numerical scale used to indicate overall air quality based on concentrations of air pollutants including ozone, particulate matter, and nitrogen dioxide

WEB **LINK**

To find out more about the Air Quality Health Index (AQHI), to check out today's forecast in your region, and to learn what you can do to help improve the air quality,

GO TO NELSON SCIENCE

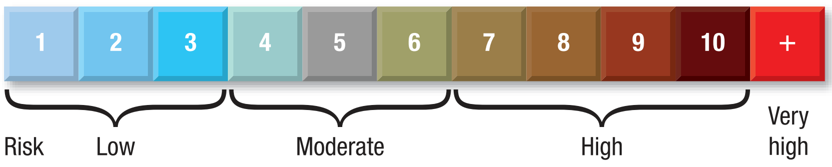

Figure 9 The Air Quality Health Index scale

In addition to the numerical values, the AQHI provides health messages to help individuals determine if they are at risk (**Table 1**).

LEARNING **TIP**

AQHI versus AQI?
The AQHI is a new index that reflects the health risks associated with the mixture of air pollutants on a given day. The Air Quality Index (AQI) reports the level of the worst individual pollutant on that day and does not consider health risks.

Table 1 AQHI Targeted Health Messages

Health risk	AQHI	Health messages	
		At-risk population	**General population**
low	1–3	Enjoy your usual outdoor activities.	Air quality is ideal for outdoor activities.
moderate	4–6	Consider reducing or rescheduling strenuous activities outdoors if you are experiencing symptoms.	There is no need to modify your usual outdoor activities unless you experience symptoms such as coughing and throat irritation.
high	7–10	Reduce or reschedule strenuous activities outdoors. Children and the elderly should also take it easy.	Consider reducing or rescheduling strenuous activities outdoors if you experience symptoms such as coughing and throat irritation.
very high	10+	Avoid strenuous activities outdoors. Children and the elderly should also avoid outdoor physical exertion.	Reduce or reschedule strenuous activities outdoors, especially if you experience symptoms such as coughing and throat irritation.

11.4 Summary

- Photochemical smog is a mixture of pollutants produced by the reaction of vehicle and factory emissions with sunlight.
- Particulate matter is produced during combustion reactions.
- Sulfur dioxide is a pollutant gas produced by the combustion of fossil fuels containing sulfur impurities. It reacts to cause acid precipitation.
- Nitrogen oxides, produced when vehicles consume fossil fuels, contribute to smog and acid precipitation.
- Carbon monoxide is a colourless, odourless gas mostly produced during incomplete combustion.
- Volatile organic compounds (VOCs) are carbon compounds that vaporize readily. VOCs may be toxic and contribute to the formation of ground-level ozone.
- Ground-level ozone is a component of smog that can cause irritation of the eyes and respiratory system.
- The Air Quality Health Index (AQHI) is designed to help Canadians assess the risk that the air quality may pose to their health.

Investigation 11.4.1

Comparing Air Quality in Several Locations (p. 563)
You will have an opportunity to explore the information contained in the Air Quality Health Index and apply it to a variety of situations.

11.4 Questions

1. (a) What is photochemical smog?
   (b) Why do you think photochemical smog is more of a problem in summer than in winter? K/U
2. (a) List the two greatest sources of particulate matter emissions in Ontario.
   (b) Outline three ways in which Canadian households can reduce their particulate matter emissions.
   (c) Why does particulate matter pose a health risk? K/U A
3. Identify two gaseous air pollutants that contribute to acid precipitation and list their main sources of emissions. K/U
4. Describe why carbon monoxide concentrations in the air may be higher than normal during a traffic jam. K/U T/I
5. (a) What are VOCs?
   (b) List three examples of consumer products that release VOCs.
   (c) What precautions should be taken when working with products that emit VOCs? K/U A
6. What effect might each of the following factors have on the formation of ground-level ozone and smog near a large city? A
   (a) wind
   (b) cloud cover
   (c) season
   (d) a public holiday
7. (a) What does AQHI stand for?
   (b) What is this scale intended to accomplish?
   (c) What segment of our population would find the AQHI particularly useful? K/U A
8. **Figure 10** shows the variation of urban nitrogen dioxide and ozone concentrations over a two-day period. T/I
   (a) At approximately what time of the day do these concentrations reach their maximum and minimum?
   (b) Suggest an explanation for your answer in (a).

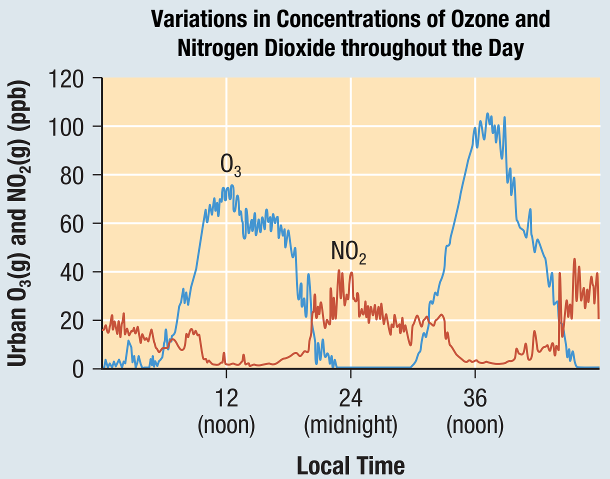

Figure 10

9. When it comes to improving air quality, politicians often argue that "Canada cannot do it alone." The cooperation of our neighbours is essential. Do you agree or disagree with this statement? Why? C A
10. Research the phenomenon of global dimming. Write a brief report that describes global dimming and outlines the effect it may have on climate change. C A

GO TO NELSON SCIENCE

11.5 Indoor Air Quality

Health Canada estimates that we spend 90 % of our time indoors. The air quality of our indoor spaces is therefore critical for good health. We think of our homes and other indoor spaces as being safe places, free from environmental pollutants. After all, we do not run cars inside and factories do not emit gases into our homes. However, indoor air is commonly two to five times more polluted than outdoor air. In this section we discuss the sources of indoor air pollutants and health problems associated with them.

Indoor air pollutants can be from chemical or biological sources. Gases and particles are chemical pollutants, whereas biological pollutants are microscopic organisms and contaminants produced by living organisms.

Chemical Pollutants

Figure 1 An open fire and candles may be cozy, but they can also be sources of particulate matter.

off-gassing the release of one or more gases from a substance or product at normal temperatures and pressures

Indoor chemical pollutants are mostly gases, but some are in the form of particulates. Particulate matter is produced in home offices by printers and copiers. It also results from combustion in fireplaces and wood-burning stoves, and from lit candles and cigarettes (**Figure 1**).

Some indoor air pollutants are intentionally released into the air. Walk into the perfume or candle section of a department store and you may be overwhelmed by the mixture of scents. While pleasant to many of us, these scents can trigger allergic reactions in some people.

Pollutants can also get into indoor air through an unintentional phenomenon called off-gassing. **Off-gassing** is the release of gases from a substance at room temperature. You might notice that a new shower curtain has a smell for a few days. This is an example of off-gassing. Off-gassing also causes that "new car" smell and the odour of a freshly painted room. Off-gassed chemicals are usually volatile organic compounds (VOCs) that evaporate into the air at room temperature (Section 11.4). Off-gassing from a particular product decreases over time. It may take months or even years for the off-gassing to cease. New buildings often have considerable levels of off-gassing (**Figure 2**).

Figure 2 Rooms should be well ventilated to allow the gases from new furnishings to escape.

We will consider three examples of chemical gas pollutants more closely: methanal (formaldehyde), carbon monoxide, and radon.

Methanal

The best known application of methanal, CH_2O, or formaldehyde, is as a preservative and embalming fluid to preserve organisms. However, it has many industrial uses and is in a variety of materials used in our homes. Methanal, a VOC, is a colourless, flammable gas with a distinctive sharp odour. Low levels of this compound are quite common in our indoor air. It can enter the air by two methods: from off-gassing and from combustion—including cigarettes.

Methanal can be off-gassed from a variety of products, including

- pressed wood products (such as particleboard, laminate flooring, plywood) made with certain glues or adhesives
- paints, coatings, and cosmetics containing preservatives
- wallpaper, cardboard, and some paper products
- fabrics and draperies coated with a "permanent press" finish

The health risks associated with methanal are related to length of exposure and individual sensitivity. Very high concentrations of methanal are linked to an increased risk of cancer in the nasal cavity. This rare cancer mostly affects individuals who were exposed to high levels of methanal at work. The risk of developing cancer due to methanal exposure in most Canadian homes is extremely small. However, it is still wise to keep the methanal levels in our homes as low as possible in order to avoid other, less serious, health effects. Short-term exposure to methanal can cause eye, throat, and nose irritation and may cause allergic sensitivity. Longer exposure may lead to respiratory problems.

How can you minimize methanal concentrations?

- Do not allow smoking inside your home or outside near doors and windows.
- Do not idle cars or other gas-powered equipment in the garage or near open doors and windows.
- Ensure that fireplaces, wood stoves, and their chimneys are clean and in good working order.
- Avoid bringing products containing methanal into your home. Find out whether there are "low-formaldehyde" alternatives to conventional products.
- If you use paints or other products containing methanal, ensure that there is good ventilation.
- Air out products containing methanal before bringing them into the house.

Carbon Monoxide

Carbon monoxide, CO, is a colourless, odourless, tasteless gas at room temperature. These qualities make carbon monoxide difficult to detect. Carbon monoxide poisoning kills more than 400 people every year in Canada.

Low-level exposure to carbon monoxide may cause flu-like symptoms such as headaches, nausea, and shortness of breath. High levels of exposure or extended periods of low-level exposure can lead to chest pain, exhaustion, poor vision, dizziness, and impaired mental function. At very high levels, symptoms include coma, convulsions, and death.

A good indication of carbon monoxide poisoning is that the person's lips, cheeks, and fingernails become very red. How does carbon monoxide poisoning occur? Oxygen is carried in the blood by hemoglobin. Hemoglobin picks up oxygen in the lungs and releases it in the tissue of the body where it is needed (**Figure 3(a)**). The waste product carbon dioxide is then carried from body tissues to the lungs where it is exhaled. Carbon monoxide, however, binds to hemoglobin over 200 times more strongly than does oxygen, producing a bright red substance called carboxyhemoglobin. The hemoglobin is therefore unavailable to carry oxygen to the tissues (**Figure 3(b)**). Without oxygen, cellular function cannot continue. The cells begin to die. This can eventually lead to death. The red carboxyhemoglobin causes the victims of carbon monoxide poisoning to have very red lips, cheeks, and fingernails.

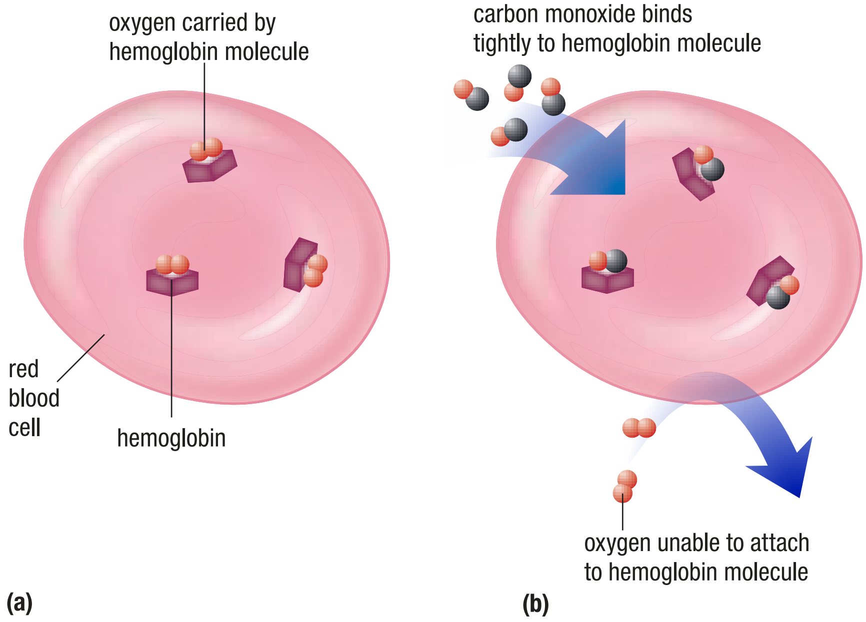

Figure 3 Carbon monoxide prevents oxygen from bonding to hemoglobin.

Figure 4 Every house should have at least one carbon monoxide detector.

Carbon monoxide is released as a product of incomplete combustion and can come from several sources. Car exhaust can enter homes from attached garages. Furnaces, fireplaces, water heaters, and other fuel-burning appliances can also release carbon monoxide. It is also a component in tobacco smoke. Carbon monoxide can also be emitted from barbecues, grills, and space heaters. This is why you should never use a barbecue indoors or in a poorly ventilated area.

The air in most Canadian homes contains less than 2 ppm of carbon monoxide. It is important to install carbon monoxide detectors in the hallways near all sleeping areas, and regularly check that they work (**Figure 4**). There are several ways to keep the level of indoor carbon monoxide low:

- Keep all fuel-burning appliances well maintained.
- Do not allow smoking inside your home.
- Do not use fuel-burning lamps or heaters designed for outdoor use inside your home.
- After a snowstorm, check that the exhaust vents for your dryer, furnace, fireplace, and wood stove are unobstructed.
- Do not idle a car or other gas-powered equipment in a garage attached to the house.

Radon

Radon is a colourless, odourless gas. As you learned in Section 1.4, radon is also a radioisotope that is produced by the radioactive decay of naturally occurring uranium in soil and rock. Almost all homes contain low levels of radon because radon can enter a building through contact with soil (Section 1.4, Figure 8). Concentrations tend to be higher in regions with uranium-rich soil. Building codes often require measures to reduce the entry of radon into homes.

According to Health Canada, radon is the leading cause of lung cancer among non-smokers. They estimate that the probability of a non-smoker getting lung cancer from long-term exposure to high levels of radon is 1 in 20. For smokers, the risk increases to 1 in 3.

The following suggestions can help minimize your exposure to radon:

- Seal any cracks and gaps in external walls, especially around pipes and drains.
- Ensure that all basement floors are sealed and well maintained.
- Ventilate basement subfloors to the outside to prevent radon from being trapped.
- Periodically measure radon levels using a radon detector.

Figure 5 (a) Dust mites and (b) mould are biological pollutants.

Biological Pollutants

Biological pollutants are living organisms or the products of living organisms. Many of these organisms thrive in the warm environments of our homes. Bacteria and viruses that cause infectious diseases, such as chicken pox and influenza, are biological pollutants. Other biological pollutants include mites, mould, dust, dander, and pollen. Fecal matter from pets, cockroaches, and rodents can also dry and become airborne.

Even humans are a source of biological pollutants. We each shed thousands of skin cells every minute. These cells are a major component of dust. Our pets contribute their own skin cells and hair, known as dander. This organic matter is a food source for dust mites (**Figure 5(a)**). Dust mites and their waste are among the most common household allergens. Dust mites live in warm, humid environments such as linens, mattresses, pillows, upholstered furniture, and carpets.

The term "mould" commonly refers to various species of fungi that grow indoors (**Figure 5(b)**). Mould fragments and spores can become airborne. Some people experience respiratory difficulties, skin rashes, or other allergic reactions when they inhale these particles.

Since most biological contaminants thrive in high humidity, we can minimize them by keeping the moisture content of the air low. Ways to reduce biological pollutants include the following strategies:

- Prevent leaks and excessive condensation and use ventilation or dehumidifiers to keep the humidity low.
- Clean up floods and spills quickly.
- Clean away any mould found indoors.
- Keep air filters clean.
- Clean the house regularly.

Research This

Telecommuting

Skills: Researching, Analyzing, Evaluating, Communicating

SKILLS HANDBOOK **A5.1, A5.2**

More and more people are working from home offices these days. Because of computers and instant electronic communication, many workers do not have to be in the same location as their colleagues and managers. They can place orders, prepare documents and spreadsheets, plan events, hold meetings, and even teach from their own homes (**Figure 6**). Working from home has many benefits. These workers do not have to travel to work, so their carbon footprint is reduced. They may work a flexible schedule, fitting the same number of hours in around other tasks and family responsibilities. On the downside, though, some people find it hard to focus on work when they are at home. The home office may not be equipped with all the required furniture and technology. The air in the home may not meet the legal requirements of clean air in a workplace.

1. Research a list of five careers or businesses that can be operated from a home office.
2. Research the average commuting time for Ontario workers. If possible, find out how most people travel to work.
3. Research the legislation that covers an official workplace, and find out whether there is any similar legislation for a home workplace.
4. Research the trend toward working at home, and people's attitudes to it.
5. Research other pros and cons of telecommuting.

A. Create a list of considerations (for example, save on gas) that affect whether or not telecommuting would be a good choice for you. T/I

B. Organize your considerations into a t-chart, with "For Telecommuting" and "Against Telecommuting" as the column headings. Include reasons for your placement of each consideration. T/I C

C. Write a one-paragraph answer to the following question: How effectively does the use of digital communications for business reduce our carbon footprint? T/I C

D. Suggest three ways to improve the air quality in a home office. T/I A

GO TO NELSON SCIENCE

Figure 6 Working from home has its pros and cons.

11.5 Summary

- Good indoor air quality is critical for good health.
- Indoor air pollutants can be classified as chemical or biological pollutants.
- Chemical pollutants include methanal, carbon monoxide, and radon. There are many ways to reduce exposure to indoor gas pollutants.
- Biological pollutants include mould, bacteria, and dust mites, and can lead to respiratory problems, asthma attacks, and allergic reactions.

11.5 Questions

1. Describe "off-gassing," giving two examples. K/U C
2. (a) List three products that can release methanal into the indoor environment.
   (b) How can we minimize the concentration of methanal in our indoor environments? List at least three suggestions. K/U
3. (a) Explain how carbon monoxide exposure can affect oxygen levels in blood.
   (b) What are the symptoms that appear with exposure to low levels of carbon monoxide?
   (c) List four examples of how to minimize your exposure to carbon monoxide at home. K/U
4. (a) Why is radon harmful?
   (b) Compare the risk of lung cancer for smokers to the risk for non-smokers.
   (c) How does radon enter our homes? K/U T/I
5. You are offered a part-time job as a cashier in an underground parking garage. You have some concerns about the air quality in this job. K/U T/I A
   (a) What pollutants should you be concerned about?
   (b) What are the symptoms that you should look for to indicate elevated levels of exposure?
6. With efforts to make homes energy efficient, we have made our homes almost airtight. Research how an airtight home affects the air quality inside the home. Explain your findings. T/I A
7. In the aftermath of Hurricane Katrina, hundreds of families moved into trailers provided by the government. Many of these people started complaining of health problems, including headaches and breathing difficulties. Research the cause of these health problems and what was done to solve them. In your opinion, how could the problems have been avoided in the first place? T/I
8. Indoor air pollution is a serious challenge in developing countries. Research indoor air pollution using the World Health Organization (WHO) website. How do the challenges of indoor air pollution differ in developing countries compared to conditions in Canada? T/I A
9. Phytoremediation is the use of plants to remove toxins from our environment. Research three indoor plants that have the ability to remove pollutants from the air. Provide a brief description of each plant including which toxins they remove. T/I C A
10. Ionizing air cleaners are designed to reduce indoor air pollution. *Consumer Reports* magazine investigated these devices. Read the report of the investigation and summarize your findings in a table. T/I C
11. Many people suffer from respiratory problems when they smell perfumes or other scented products. In Canada, a worker can request a scent-free environment under the *Human Rights Act*. Research this legislation. In your opinion, are the rights of people to wear perfume unfairly restricted by this legislation? Present your opinion and your reasons for it, in a short oral or audiovisual presentation. T/I C
12. Electronic equipment used in offices and home offices can contribute to poor indoor air quality. Research the connection between air quality and electronics and write a paragraph summarizing your findings. T/I
13. In 2010, Health Canada banned 11 hair smoothing products that were found to contain more than the allowable level of methanal. Research how much methanal was contained in these products. What is the allowable level of methanal in hair products? Do you think methanal should be allowed in products such as these? T/I C

GO TO NELSON SCIENCE

11.6 Explore an Issue in Atmospheric Chemistry

Car Idling

SKILLS MENU

- Defining the Issue
- Researching
- Identifying Alternatives
- Analyzing
- Defending a Decision
- Communicating
- Evaluating

Take a moment to consider the person serving you at a drive-through. While you are receiving your super-sized meal, he or she may be exposed to a super-sized dose of car exhaust. Exhaust from idling vehicles contributes to poor air quality in the immediate area (**Figure 1**).

Figure 1 Idling cars at a drive-through increase the levels of carbon monoxide and other gases.

Air quality is an important environmental issue in our homes, schools, and workplaces. Many factors affect air quality. Car exhaust contains a number of pollutants. When vehicles burn fuel they produce a variety of airborne chemicals, including carbon monoxide, nitrogen oxides, VOCs, and particulate matter (Section 11.4). Carbon monoxide, when inhaled, prevents the blood from transporting oxygen around the body. Nitrogen oxides lead to the production of ground-level ozone, which can cause respiratory problems. VOCs cause a range of health problems, from irritated eyes to nausea. Particulate matter can irritate the lungs, making existing respiratory problems worse. Idling car and bus engines are the primary source of these pollutants. An idling engine operates at lower temperatures than normal, resulting in more incomplete combustion and more of the resulting pollutants.

Children are at particularly high risk from the health effects caused by air pollution because they breathe more frequently than adults, inhaling more air for each kilogram of their body weight. In addition, students often socialize outside or participate in outdoor activities. People who are active outdoors may also be particularly sensitive to air pollution.

The air quality around schools can be poor, particularly when students are being dropped off or picked up. Car idling affects the air quality in the school environment.

Figure 2 Some regions already have bylaws to limit car idling.

The Issue

Maintaining good air quality around schools is in everyone's best interest. Does car idling affect the air quality near your school? How effective are the policies regarding idling (**Figure 2**)? What, in your opinion, can be done to improve air quality?

Goal

To develop and present a plan that, when implemented, will improve the air quality near your school

Research

SKILLS HANDBOOK A5.1

Items to consider in your research include

- the effects of car emissions on air quality and the health of young people.
- whether your municipality has a bylaw for car idling. If so, what does the bylaw state? How well is it observed? How is it enforced?
- the effectiveness of existing car idling programs at other schools or in other municipalities. Determine whether similar programs would work well in your situation.
- the financial cost of car idling, including wear and tear on the car and the quantity of gasoline used per minute.
- the myths vs. realities of car idling, such as the quantity of gasoline used when a car is stopped and started compared to the quantity used during idling.

WEB LINK

To start your research into the effects, bylaws, and costs associated with vehicle idling,

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Identify Solutions

Consider ways to improve the air quality near your school.

- If there is no bylaw restricting idling, would an anti-idling bylaw improve the situation? Is there another, better way to reduce air pollution?
- If there is a bylaw restricting idling, is it effective? Could it be made more effective? If so, how?

Brainstorm suggestions. Remember that they must be realistic and enforceable.

Make a Decision

Decide which of your suggestions is most likely to be effective in improving the air quality around your school. Develop a plan to put your suggestion into action. Decide whom you will need to convince in order to implement your plan. It might be your principal, the parent council, the school board, your local councillor, or the municipal council.

Communicate

Prepare a written proposal and presentation based on your decision. Your proposal should be targeted at the people who can help you to implement your plan. Your presentation could be in the form of a talk, a movie, a slideshow, or any other appropriate format of your choice. It should include a summary of your research and persuasive arguments in support of your plan.

Do a "dress rehearsal" of your presentation for your class.

Plan for Action

As a class, consider all the suggestions that have been presented. Decide which would be the best one to pursue. Your decision could be based on overall impact, likelihood of success, innovation, community involvement, or any other criteria of your choice. When you have made your class selection, work together to refine the presentation.

If appropriate, make your presentation to the relevant authority. Put your plan into action!

Atmospheric Pressure

Have you ever felt your ears "pop" in an airplane or when driving over a hilly road? What you are experiencing is a change in air pressure. Like everything else near Earth, the atmosphere is under the influence of gravity. Gravity pulls this enormous mass toward the centre of Earth.

Pressure

Pressure is defined as the force per unit area. Like volume and temperature, pressure is a physical property of a gas. As you will soon learn, there are many important applications of gas pressure. Before discussing air pressure specifically, we will consider a more visible example of pressure. **Figure 1** illustrates some of the factors that determine pressure. The gravitational pull of Earth exerts a downward force on everyone, including the person in Figure 1. As a result, he exerts a downward force on the nails. However, to minimize pain, this person would be wise to distribute this force over as large an area as possible. The greater the area, the lower the pressure. He would still exert the same downward force if he were to stand on one foot on the nails. However, the force would then be concentrated into a smaller area. The smaller the area, the greater the pressure.

Mathematically, pressure, P, is expressed as

$$P = \frac{F}{A}$$

The SI unit for pressure is the pascal (1 Pa = 1 N/m^2). Pressure is directly related to the size of force applied. The greater the force, the greater the pressure. Pressure is inversely related to area. A large force applied to a small area will produce a large pressure. If the same force is applied to a large area, the pressure will be less. For example, if a force 100 N were applied to an area of 1.00 cm^2 (0.000100 m^2 or 1.00×10^{-4} m^2), the pressure would be 1.00×10^6 Pa (**Figure 2(a)**). If the same 100 N force were applied to a much larger area of 1.00 m^2, the pressure would be only 100 Pa (**Figure 2(b)**).

Figure 1 The more nails there are, the less painful this experience will be.

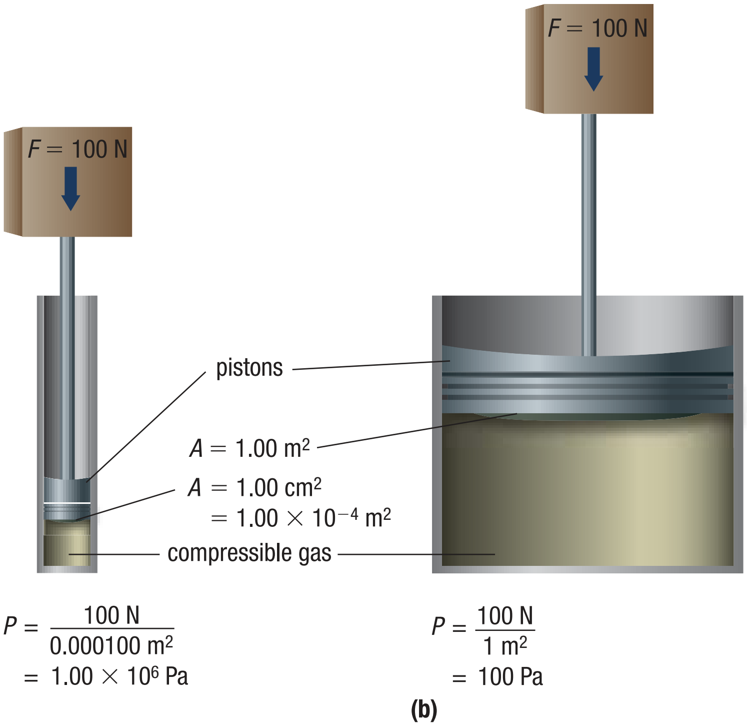

Figure 2 The smaller the surface area on which the mass is resting, the greater the pressure exerted.

When a piston applies a pressure to a trapped sample of gas, as in Figure 2, the gas exerts a pressure on the walls of its container. It is the force exerted by the gas molecules as they collide with the inner walls of the container that results in the observed pressure. These collisions are what keep bicycle tires hard when they are inflated.

Measuring Atmospheric Pressure

atmospheric pressure the force per unit area exerted by air on all objects

standard pressure 101.325 kPa (often rounded to 101 kPa)

standard temperature and pressure (STP) 0 °C and 101.325 kPa

standard ambient temperature and pressure (SATP) 25 °C and 100 kPa

Atmospheric pressure is the force per unit area exerted by air on all objects. It is commonly reported in kilopascals, kPa. At sea level, the pressure exerted by a column of air with a base of one square metre is equal to 101.325 kPa (often rounded to 101 kPa). This pressure is known as **standard pressure** and is the basis for another unit of pressure: the atmosphere. One atmosphere is equal to 101.325 kPa.

Traditionally, chemists defined the standard conditions for work with gases as the temperature 0 °C and pressure 101.325 kPa. A gas sample at these conditions is said to be at **standard temperature and pressure (STP)**. However, since 0 °C is not a convenient temperature at which to conduct laboratory investigations, scientists have recently defined another set of standard conditions. These conditions are called **standard ambient temperature and pressure (SATP)** and are defined as 25 °C and 100 kPa. The SATP standard is more convenient than STP because it more closely represents the conditions in a laboratory.

Evangelista Torricelli (1608–1674) was the first person to devise a method of measuring atmospheric pressure. He was trying to solve a problem. Pump makers in Tuscany could not raise water more than 10 m using a suction pump. Torricelli used mercury, which is denser than water, to investigate the vacuum and atmospheric pressure. He prepared a glass tube similar to an extremely long test tube. He filled the tube with mercury and carefully inverted it, submerging the open end into a dish containing more mercury (**Figure 3**). The mercury in the tube was pulled down by gravity. However, the mercury did not all run out of the tube. Why not? Air pressure pushed on the mercury in the dish, effectively pushing mercury into the tube. A vacuum formed at the top of the tube. The vacuum exerted no downward pressure on the mercury inside the tube.

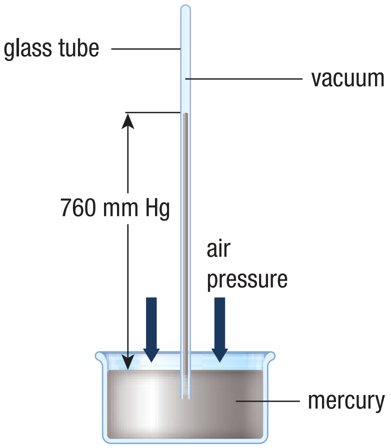

Figure 3 Torricelli's apparatus for measuring atmospheric pressure was based on the work of an earlier scientist: Galileo.

Torricelli noticed that the mercury level in the tube changed slightly from day to day. The fluctuating mercury level was due to changes in air pressure. This device for measuring atmospheric pressure became known as a barometer. At one time, the standard pressure was defined as 760 mm Hg or 760 Torr in honour of Torricelli.

Scientists had been investigating gases for many years before there was a standardized unit for pressure. Some scientists developed their own ways of measuring pressure. This is one reason we now have so many units for pressure. Some of these units are used in a specific situation. For example, medical professionals use mm Hg for measuring blood pressure. In Canada we still commonly measure tire pressure in psi (pounds per square inch), even though we use the metric system for many other quantities. **Table 1** shows the conversion of several SI and non-SI pressure units.

Table 1 SI and Non-SI Units of Pressure

Unit name	Unit symbol	Definition/conversion
pascal	Pa	1 Pa = 1 N/m^2
millimetres mercury	mm Hg	760 mm Hg = 1 atm = 101.325 kPa
torr	Torr	1 Torr = 1 mm Hg
atmosphere	atm	1 atm = 101.325 kPa (exactly)
pounds per square inch	psi	1 psi = 6895 Pa

Tutorial 1 Converting between Units of Pressure

Sometimes you are given a measurement of pressure in one unit, such as millimetres of mercury (mm Hg), and you need to convert it into a different unit, such as pascals (Pa). This is a fairly simple mathematical procedure. You can use the definitions in Table 1 to write conversion factors that allow you to switch from one unit to another.

Sample Problem 1: Converting from kPa to mm Hg

The average atmospheric pressure on Mars is 0.60 kPa. What is this value in mm Hg?

Given: $P = 0.60$ kPa

Required: pressure in mm Hg

Solution:

Step 1. Find the relationship between kPa and mm Hg (Table 1).

$760 \text{ mm Hg} = 101.325 \text{ kPa}$

Step 2. Write the relationship as a fraction, with the unit you want to find as the numerator.

$$\frac{760 \text{ mm Hg}}{101.325 \text{ kPa}}$$

Step 3. Multiply the given value by the conversion factor developed in Step 2.

$$0.60 \cancel{\text{kPa}} \times \frac{760 \text{ mm Hg}}{101.325 \cancel{\text{kPa}}} = 4.5 \text{ mm Hg}$$

Statement: The average atmospheric pressure on Mars expressed in mm Hg is 4.5 mm Hg.

Practice

1. Convert each of the following measurements of pressure to the units indicated: K/U
   (a) 203 kPa to mm Hg [ans: 1520 mm Hg]
   (b) 40.0 kPa to Torr [ans: 3.00×10^2 torr]
   (c) 717 mm Hg to Pa [ans: 9.56×10^4 Pa]

Atmospheric Pressure and Altitude

More than 8000 climbers have attempted to scale Mount Everest, the highest mountain on Earth (elevation 8850 m). More than 2000 of these climbers have been successful, while more than 200 people have perished in the attempt. At least 120 bodies remain lost on Mount Everest. Reaching the summit is a gruelling journey. The cold and the rough terrain are certainly difficult, but the atmosphere presents its own challenge. The air density is low at this altitude, with less oxygen per unit volume than air at sea level. As a result, most climbers require oxygen tanks at high altitudes (**Figure 4**).

The density of gases in the atmosphere changes with altitude, decreasing as altitude increases (**Figure 5**). Greater gas density means more collisions per unit area and, therefore, higher pressure at Earth's surface. Like air density, atmospheric pressure decreases as altitude increases (**Table 2**).

Figure 4 Most climbers need oxygen tanks to complete the final summit of Mount Everest.

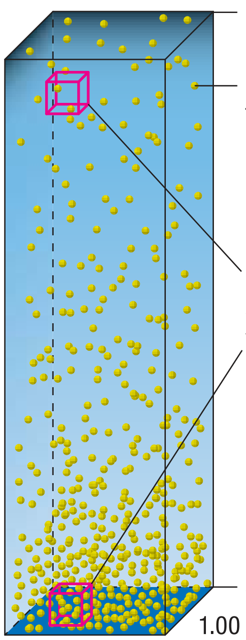

Figure 5 There are more entities in any given volume of air at sea level than in the same volume high in the atmosphere. The density of air is therefore greater at sea level than higher up.

Table 2 Typical Air Pressure at Various Altitudes above Sea Level

Location	Altitude (in m)	Air pressure (kPa)
Mount Everest (highest location on Earth)	8850	34
Mount McKinley (highest location in North America)	6194	48
Mexico City	2240	78
Calgary, AB	1049	90
Toronto, ON	112	100
Montreal, QC	57	101
Victoria, BC	49	101
Death Valley (lowest location in North America)	−86 (below sea level)	102
Dead Sea (lowest land location on Earth)	−413 (below sea level)	106

The change in atmospheric pressure is the reason why your ears sometimes hurt when you change altitude quickly. Your ear is a complicated organ designed to detect sound waves. The middle ear is an air-filled chamber that is isolated from the outside air by the eardrum and is connected to a channel called the Eustachian tube which vents into your throat (**Figure 6(a)**).

When a plane takes off and climbs higher in the sky, the atmospheric pressure in the cabin decreases. With less pressure on the eardrum, the volume of gas in the ear increases and presses the eardrum out. This gives the uncomfortable feeling of fullness in the ears. Fortunately, you can stop the discomfort by making the Eustachian tubes open up, allowing air to flow from the middle ear into the throat. This equalizes the pressure of the air in the middle ear with the atmospheric pressure in the plane. This venting is the "popping" sensation that you feel as your ears clear. Chewing gum, yawning, and swallowing all tend to make it easier for the Eustachian tube to vent air into your throat (**Figure 6(b)**). Aircraft designers have to factor this huge pressure change into their engineering plans.

CAREER **LINK**

To find out more about the work of an aircraft designer,

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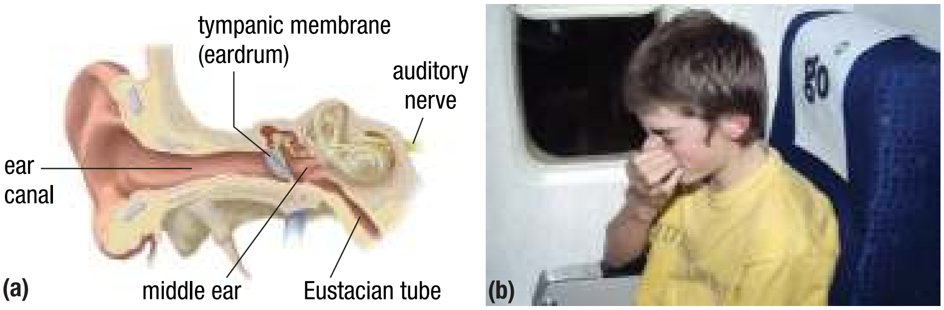

Figure 6 (a) A diagram of the ear (b) This is one way to clear your ears during air travel.

WEB **LINK**

To see a video of this tank being crushed,

GO TO NELSON SCIENCE

Atmospheric pressure can cause damage on a much larger scale. Railway tank cars are designed to withstand pressures that are higher inside than outside. If the pressure inside the tank were suddenly made much lower than the pressure outside, the atmospheric pressure would crush the tanks (**Figure 7**).

Figure 7 What happens when you reduce the internal pressure in a sealed vessel? (a) The interior of the steel tanker was heated. This expanded the air inside the tanker. The tanker was then sealed. (b) As the air inside cooled and contracted, the outward pressure it exerted could no longer match the pressure of the atmosphere—the tanker collapsed!

Mini Investigation

How Strong Is Your Pop Can? (Teacher Demonstration)

Skills: Questioning, Planning, Performing, Observing, Analyzing, Communicating SKILLS HANDBOOK A1.2, A2.1

In this investigation you will explore the effect of atmospheric pressure on a pop can. You will boil water inside the can to produce water vapour. Then you will cool the can rapidly by placing it in a pail of water.

Equipment and Materials: chemical safety goggles; lab apron; plastic pail; graduated cylinder; empty aluminum pop can; beaker tongs; heat source (hot plate or Bunsen burner clamped to a retort stand); tap water

This activity may involve open flames and boiling water. Tie back long hair and secure loose clothing and jewellery.

1. Wearing chemical safety goggles and a lab apron, your teacher will fill a bucket three-quarters full with cold water.
2. An aluminum pop can will be filled with about 10 mL of water.
3. Your teacher will then hold the can over a heat source until the water boils.
4. The can will then be placed over the pail of water, inverted, and submerged in the cold water.
5. Observe any changes in the aluminum can.

A. How does heating the water in the can change the conditions inside the can? T/I
B. What effect does the water in the pail have on the conditions inside the can? T/I
C. Explain the changes that you observed when you inverted the can in the water. T/I

High-Altitude Training

Many endurance athletes train at high altitudes in an attempt to improve their performance. Some research shows benefits to this type of training, but other studies do not. When athletes train at high altitudes, they generally go to elevations above 2000 m where the air pressure is 77 to 80 kPa (Table 2, page 543). At this altitude there is still 21 % oxygen in the air but all atmospheric gases are at lower density. A lower density means that each breath contains less oxygen than it would at sea level. After three or four weeks the body compensates for lower oxygen levels by making more red blood cells to carry oxygen and producing more enzymes to utilize oxygen.

When athletes return to lower altitudes they may feel energized, having an increased ability to use oxygen. Not all athletes perform better after altitude training. Disappointing performance may result because the athlete cannot train as rigorously while the body adjusts to the higher altitude and lower oxygen level. Some endurance athletes live at high altitudes but train at low altitudes. They believe that this way they obtain the physiological advantages of high altitudes, but they can still train intensively.

WEB LINK

To check out an interactive graph that compares the concentration of atmospheric oxygen at various altitudes,

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CAREER LINK

To find out more about being an athletic trainer,

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11.7 Summary

- Pressure (P) is defined as the force (F) exerted per unit area (A).

$$P = \frac{F}{A}$$

- The SI unit for pressure is pascal (Pa). 1 Pa = 1 N/m^2
- Atmospheric pressure is the force per unit area exerted by air on all objects. Standard pressure (the pressure exerted at sea level by a column of air with a base of 1 m^2) is 101.325 kPa.
- Standard temperature and pressure (STP) conditions are 0 °C and 101.325 kPa.
- Standard ambient temperature and pressure (SATP) conditions are 25 °C and 100 kPa.
- The density of atmospheric gases is greatest at sea level.

11.7 Questions

1. Copy and complete **Table 3** in your notebook. K/U

Table 3 Equivalent Values of Pressure in Different Units

	Pressure (kPa)	Pressure (mm Hg)
(a)	58	
(b)		125
(c)	130	
(d)		950

2. (a) Define STP and SATP.
   (b) Explain why scientists have defined two sets of standard conditions.
   (c) Which standard is most frequently used, and why? K/U
3. Skydiving is a sport that requires specific equipment for the demands of the environment (**Figure 8**). According to the Canadian Sport Parachuting Association, oxygen is mandatory for jumps that exceed 4572 m. Explain why this is necessary. T/I A

Figure 8 At high altitudes, skydivers must wear oxygen masks.

4. Some athletes sleep in low-pressure tents prior to participating in an athletic event at a high altitude. How is this likely to improve athletic performance? Do you think this is likely to be more or less effective than high-altitude training? Explain. K/U A
5. Modern aircraft have pressurized cabins. T/I A
   (a) What is meant by "pressurized cabin"?
   (b) What can occur if pressure is lost in the cabin?
6. The Summer Olympic Games were held in Mexico City in 1968. They produced some interesting results. The performance of many endurance athletes fell short of expectations. Many events that involved jumping and throwing, however, produced better than expected results. Use the data in Table 2 to help explain these results. A
7. (a) Create a graph of air pressure against altitude, using the data in Table 2. Label each of the locations on your graph.
   (b) Describe the trend in the data. Suggest an explanation for this trend.
   (c) Research the altitude and typical air pressure where you live, and add it to the graph. K/U C A
8. Research altitude sickness. What are the symptoms? Explain this phenomenon. T/I A
9. Research how air is used in pneumatic nail guns. T/I C A
10. High-pressure injectors, sometimes called jet injectors or hypospravs, are used to inject vaccines and other medicines under the skin of patients without using a needle. Research how these devices work and the benefits and drawbacks associated with their use. T/I A

GO TO NELSON SCIENCE

The Gas Laws—Absolute Temperature and Charles' Law

Everyone loves the light, fluffy taste of popcorn while watching a good movie. Popcorn is thought to have been discovered by the Native Americans thousands of years ago (**Figure 1**). Archaeologists believe that popcorn was originally popped by placing it on extremely hot stones. As it popped, it shot off this way and that. Popcorn behaves as it does because of the behaviour of gases. The study of gases and their application has a long history in chemistry.

Figure 1 Popcorn is a tasty application of the gas laws.

While you may not be aware of it, you have quite a bit of experience with the properties of gases. When the weather turns cool in the fall, you might notice that you need to pump up your bicycle tires. During air travel or when diving deep into a swimming pool you have probably felt a strange sensation in your ears. While we can easily see and feel the effects of gases in such examples as popcorn and painful ears, studying gases is a bit more of a challenge. We cannot see most gases. They are not easy to touch and feel. Too often, we think of a container being empty if it "only" has air in it.

Curious people often make observations of phenomena, and then search for patterns and explanations. Observations of how gases behave preceded the laws that now help us predict and explain gas behaviour. Many of the gas laws are named after the scientists who made the observations. Before discussing the properties of gases and the laws that describe these properties, we will consider the two temperature scales that scientists most often use when conducting investigations with gases.

Absolute Zero and Temperature Scales

How can a temperature have a negative value? Why does water boil at 100 °C? Is 30 °C twice as hot as 15 °C? The Celsius scale, developed by the Swedish scientist Anders Celsius in 1742, is convenient for most everyday situations. Celsius devised the scale by taking a thin, closed glass tube of a pure liquid, such as mercury, and recording the height of the liquid when the tube was placed in ice water. He called this height "100 degrees." He then repeated the process placing the tube in boiling water to find the height that he called "zero degrees." Celsius then divided the distance between the two marks evenly into 100 divisions. Each of these divisions is one degree Celsius. The scale was later reversed to make it a more practical unit of measure (**Figure 2**). (The Celsius scale is also known as the centigrade scale. Centigrade simply means "divided into 100 degrees.")

Figure 2 The Celsius scale is convenient for daily applications.

WEB **LINK**

Anders Celsius was primarily an astronomer. To find out more about the life and achievements of Anders Celsius,

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Later, another temperature scale was developed that proved to be more useful. Examine **Figure 3**, which is a volume versus temperature graph for several different gases. We can extend the lines to the left of the measured values to find the theoretical volumes of gases as their temperatures decrease.

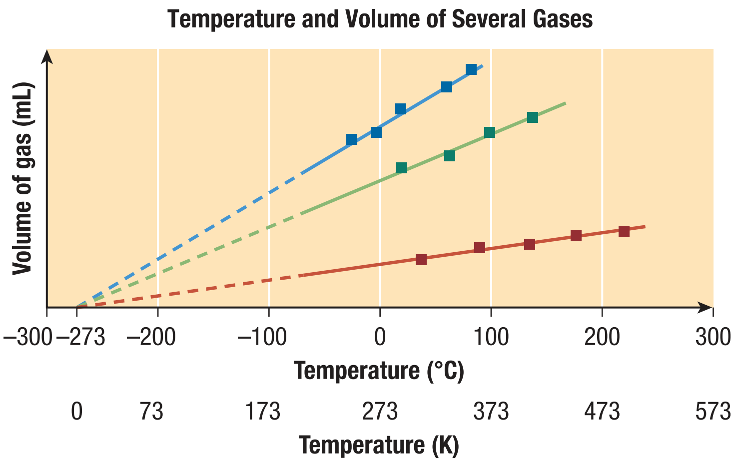

Figure 3 The solid lines represent actual measurements. When the graphs of several volume–temperature experiments are extrapolated (dashed lines), all the lines meet at absolute zero, –273.15 °C or 0 K.

What happens when we reach a theoretical volume of zero? If we extrapolate the data shown on the graph, we see that the lines will all intersect the temperature axis at −273.15 °C. Lord Kelvin was the first scientist to notice this. He experimented with a variety of gases, graphing their volumes at various temperatures, and noticed that the graphs of all gases intercept the temperature axis at the same point. This implied that the volume of every gas would become zero at a temperature of −273.15 °C.

This temperature, at which a gas theoretically has no volume, is now known as **absolute zero**. It is the lowest possible temperature. There is some controversy over which scientist is responsible for this term. (See Section 12.3, The Science of Cold.)

absolute zero the theoretical temperature at which the entities of a material contain no kinetic energy and therefore transmit no thermal energy; equal to −273.15 °C

Lord Kelvin developed a new temperature scale based on the value of −273.15 °C. He decided that this measurement would be equivalent to 0 on the new scale. By setting a value of zero for the lowest temperature that matter can achieve, Lord Kelvin ensured that the new scale would have no negative values. This is the scale that we now call the **Kelvin temperature scale** or the **absolute temperature** scale. It has the unit symbol "K." Chemists and others studying gases routinely use this scale for their measurements and calculations. We will also use it throughout this unit. **Figure 4** shows some typical values of Celsius temperatures and the corresponding values in kelvins. In theory, at 0 K, entities should have no motion and therefore no kinetic energy. Scientists ever since have been striving to slow the motion of entities, in a race to achieve absolute zero.

Kelvin temperature scale a temperature scale that includes absolute zero and the same-sized unit divisions as the Celsius temperature scale

absolute temperature a measurement of the average kinetic energy of the entities in a substance; unit symbol K

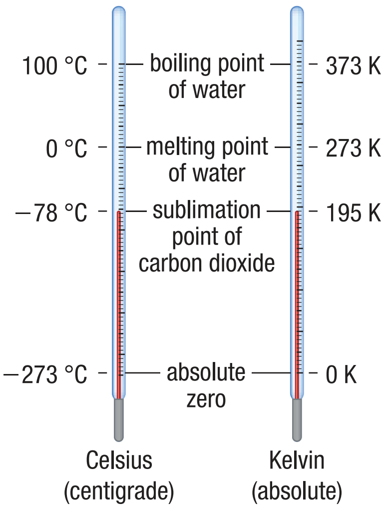

Figure 4 Various freezing and boiling points on the Celsius scale and the Kelvin scale

Tutorial 1 Temperature Conversions

To convert temperatures from degrees Celsius to kelvins (absolute temperature) or vice versa, use the following conversion equation, in which t represents the temperature in degrees Celsius and T represents the temperature in kelvins:

$$T = t + 273.15$$

You will notice that most temperature readings are taken in degrees Celsius and are whole numbers. After converting these values to kelvins you will always round to the nearest whole number. For this reason, the value of 273 is adequate for temperature conversions involving whole numbers. When converting temperatures, use your common sense. You have a great deal of experience with temperatures between –20 °C and 40 °C.

Sample Problem 1: Converting from Celsius to Kelvin

Standard ambient temperature is defined as 25 °C. Convert this temperature to absolute temperature.

Given: $t = 25\ °C$

Required: T

Analysis: $T = t + 273$

Solution: $T = 25 + 273$

$T = 298\ K$

Statement: Standard ambient temperature is 298 K.

Sample Problem 2: Converting from Kelvin to Celsius

Convert 150 K to degrees Celsius.

Given: $T = 150\ K$

Required: t

Analysis: $T = t + 273$

$t = T - 273$

Solution: $t = 150 - 273$

$t = -123\ °C$

Statement: A temperature of 150 K is equivalent to −123 °C.

LEARNING TIP

°C vs. K

Note that the unit for centigrade is degree Celsius (°C) whereas the unit for absolute temperature is kelvin (K). There is no degree symbol.

Practice

SKILLS HANDBOOK A6.1, A6.3

1. Convert the following temperatures from degrees Celsius to absolute temperature: K/U
   (a) Dry ice sublimes into a gas at −78 °C. [ans: 195 K]
   (b) The hottest temperature recorded in Canada is 45 °C (in Saskatchewan, 1937). [ans: 318 K]
2. Convert the following temperatures from absolute temperature to degrees Celsius: K/U
   (a) Pure gold melts at 1337 K. [ans: 1064 °C]
   (b) The coldest temperature recorded in Canada is 210 K (in Snag, Yukon, 1947). [ans: –63 °C]

Charles' Law—The Relationship between Volume and Temperature

What causes a kernel of popping corn to pop? Popping corn has two unusual features: a rigid, non-porous shell and a small quantity of water (8 % to 20 % by mass) trapped inside. As the kernel is heated the water becomes a gas: water vapour. This vapour expands as much as it can within the non-porous shell. Eventually the expanding water vapour bursts the shell, causing the kernel to "pop." This relationship between the temperature and volume of the water vapour in popcorn is an illustration of a gas law.

An understanding of the relationship between the volume of gas and temperature has been attributed to French scientist Jacques Charles (1746–1823). Charles investigated the expansion of a variety of gases by placing a sample of gas in a closed, expandable container. For example, a container might be sealed by a piston that was free to move. In such a container, the amount of gas (in moles) and the pressure were kept constant.

Charles found that the volume of a gas increases as its temperature increases. Graphing volume against Celsius temperature produces a straight line. Figure 3 on page 547 shows this type of relationship for three gases. As each gas is cooled, it eventually condenses into a liquid. Hence, we cannot measure the temperatures of gases below this point.

The relationship between the volume and temperature of a gas is called **Charles' law.** This law states that the volume of a gas is directly proportional to its temperature, in kelvins, provided that pressure and the amount of gas are constant. Mathematically, we can represent Charles' law by the equation

$$V = aT$$

In this equation a is a constant and T is the absolute temperature (in kelvins). The equation for Charles' law suggests that doubling the absolute temperature, for example from 273 K to 546 K, results in a doubling of the volume (**Table 1**.) Note, however, that doubling the Celsius temperature, for example from 20 °C to 40 °C, does not result in a doubling of volume. This is why absolute temperature is used in mathematical problems.

Table 1 The Effect of Increasing Temperature on the Volume of a Gas

Temperature, t (°C)	Temperature, T (K)	Volume, V (mL)
0	273	100
20	293	107
40	313	115
80	353	129
273	546	200

WEB LINK

To find out more about the science of popcorn,

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Charles' law the statement that as the temperature of a gas is increased, the volume of the gas increases proportionally, provided that the pressure and amount of gas remain constant; the volume and temperature of a gas are directly proportional

LEARNING TIP

Temperature in Kelvin—Absolutely Always!

The Celsius scale is convenient for everyday use. When we confront mathematical problems with multiplication and division, however, it becomes problematic. A temperature rise from 20 °C to 40 °C is not a 100 % increase in temperature because the absolute temperature is only rising from 293 K to 313 K. If temperature is a variable in a calculation, the Kelvin scale must be used.

LEARNING **TIP**

Direct Relationship

When one variable is changed (increased or decreased) and the other responds in a similar manner, the two variables are said to be directly related (or directly proportional) to each other. For example, if everyone in your class increased the time they spent preparing for a test and the marks for the class increased, the two variables would be directly proportional.

Since $V = aT$, then $\frac{V}{T} = a$ (a is a constant)

This suggests that any two sets of volume and temperature values for the same gas under the same experimental conditions, should give the same ratio. We can check this prediction with two sets of values from Table 1, such as (273, 100) and (353, 129):

$$\frac{100 \text{ mL}}{273 \text{ K}} = 0.37 \text{ mL/K} \qquad \frac{129 \text{ mL}}{353 \text{ K}} = 0.37 \text{ mL/K}$$

Using the fact that the ratios are the same, we can also express Charles's law in terms of an initial set of measurements V_1, T_1 and a final set of measurements V_2, T_2:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

We can therefore express Charles' law both mathematically and in words:

Charles' Law

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}, \quad \text{or} \quad \frac{V}{T} = a \text{ (constant)}$$

The volume of a gas is directly proportional to its temperature in kelvins, provided the pressure and the amount of gas remain constant.

To explain Charles' law, consider what happens at the molecular level (**Figure 5**). As the temperature of the gas is increased, the kinetic energy of the gas entities also increases. Entities move more quickly and the number of collisions between entities increases. The number of collisions between the entities and the walls of the container also increases, making the container expand and increasing the volume of the gas.

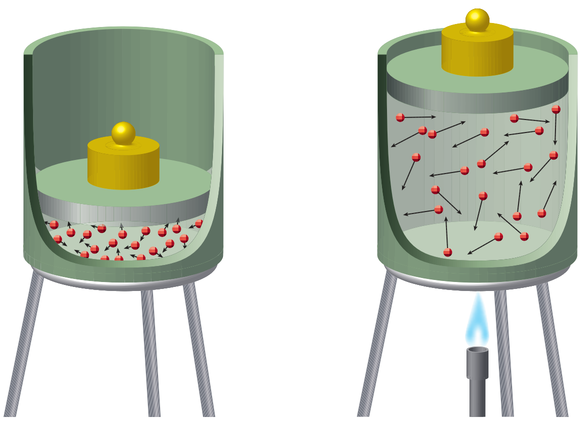


Figure 5 The volume of a gas increases as its temperature is increased. In this case, the container has a movable piston so that volume can change. The pressure of the gas remains constant in this situation.

Figure 6 Jacques Charles (1746–1823) designed and flew the first hydrogen balloon in 1783. His experiences and experiments led to the formulation of Charles' law.

Interestingly, Jacques Charles was a balloon enthusiast. Most balloonists at the time used hot air to lift their balloons. Charles experimented with using hydrogen gas instead (**Figure 6**). He most likely generated the hydrogen by reacting large quantities of acid with iron. In 1783, one of his hydrogen balloons reportedly travelled from Paris to the outskirts of the city in 45 minutes. When it landed, local residents—not knowing what the balloon was—attacked and destroyed it!

Mini Investigation

Soap in the Microwave (Teacher Demonstration)

Skills: Observing, Analyzing, Communicating

SKILLS HANDBOOK A1.2, A2.4

The strange white blob in **Figure 7** is not ice cream or mashed potato, but soap! This activity involves observing the effects of microwaves on a piece of Ivory soap. This soap is well known for floating in the bathtub. During the production of Ivory soap, air pockets are created within the bar itself. This makes it behave in a predictably wild way.

Figure 7 Do not do this at home!

Equipment and Materials: microwave oven; microwave-safe dish; 400 mL beaker; oven mitts; water; half a bar of Ivory soap cut into 2 pieces; a piece of another brand of soap

This activity involves heating substances in a microwave oven. Allow all heated objects to cool completely before touching them.

1. Your teacher will place one piece of Ivory soap in a dish and heat it in the microwave oven at medium to high power for approximately 90 s.
2. Your teacher will circulate a small piece of Ivory soap and a small piece of regular soap. Record any interesting observations.
3. While the microwave is running, your teacher will place a small piece of Ivory soap and a small piece of another brand of soap into a beaker of water. Record your observations.
4. When the microwave has finished running, your teacher will remove the soap using oven mitts. Observe and record any changes.
5. When it has cooled, use the microwaved soap to wash your hands. Record your observations.
6. Your teacher may now heat a piece of the other, regular soap in the microwave as a control.

A. How do the properties of the two soaps differ? T/I

B. Suggest an explanation for the unusual properties of Ivory soap. T/I

C. How does the behaviour of Ivory soap demonstrate Charles' law? K/U

Research This

Evaluating the Use of Nitrogen in Car Tires

Skills: Researching, Analyzing, Evaluating, Identifying Alternatives, Defending a Decision

SKILLS HANDBOOK A5.1

Good tire maintenance is necessary to optimize car performance, reduce gasoline consumption, and minimize environmental impacts. Proper maintenance includes keeping tires inflated to the specifications of the car. Traditionally, nitrogen gas has been used in race car tires. It is now available in some regions for use by everyday consumers in car tires (**Figure 8**).

1. Research the importance of good tire maintenance.
2. Research the benefits and drawbacks of using nitrogen gas in tires.

A. How can underinflated or overinflated tires affect how long tires last? A

B. What is the effect of underinflation and overinflation of tires on gas consumption and on the environment? T/I A

C. Why is nitrogen gas used in race car tires? K/U

D. In your opinion, is using nitrogen in everyday applications beneficial? Explain your answer. A

GO TO NELSON SCIENCE

Figure 8 Should nitrogen be used in car tires?

Tutorial 2 / Using Charles' Law

The mathematical form of Charles' law is very useful for predicting new values of volume and/or temperature when changes have occurred to a sample of gas. Remember that the amount of gas and the pressure do not change in situations where we apply Charles' law.

Sample Problem 1: Temperature and Volume

A sample of gas is drawn into a piston. If the sample occupies 0.255 L at 25 °C, what volume will the gas occupy if it is heated to 80 °C? The pressure and amount of gas are kept constant throughout the process.

Given: initial Celsius temperature, $t_1 = 25\ °C$

final Celsius temperature, $t_2 = 80\ °C$

volume, $V_1 = 0.255\ L$

The amount of the gas and the gas pressure remain constant.

Required: final volume, V_2

Analysis: Apply Charles' law to the situation.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Solution:

Step 1. Convert temperature values to kelvins.

$$\begin{aligned} T &= t + 273 \\ T_1 &= t_1 + 273 \\ &= 25 + 273 \\ T_1 &= 298\ K \\ T_2 &= t_2 + 273 \\ &= 80 + 273 \\ T_2 &= 353\ K \end{aligned}$$

Step 2. Rearrange the Charles' law equation to isolate the unknown variable.

$$V_2 = \frac{V_1 T_2}{T_1}$$

Step 3. Substitute given values (including units) into the equations and solve.

$$\begin{aligned} V_2 &= \frac{0.255\ L \times 353\ \cancel{K}}{298\ \cancel{K}} \\ &= 0.300\ L \end{aligned}$$

Statement: The volume of the gas at 80 °C is 0.300 L.

Check your answer. Does it make sense? Are the units appropriate?

Practice

SKILLS HANDBOOK A6

1. A sample of methane gas occupies an initial volume of 5.25 L at an initial temperature of 200 K. The gas is heated to 300 K while the pressure and the amount of gas remain constant. Determine the new volume. [ans: 7.88 L]
2. A sample of carbon dioxide is placed in a piston. The initial temperature of the gas is 35 °C and it occupies a volume of 2.2 L. Calculate the temperature at which it will occupy 4.4 L. Pressure and the amount of gas remain constant. [ans: 620 K]

11.8 Summary

- Absolute zero on the Kelvin scale (0 K or –273.15 °C) is the theoretical temperature at which entities have no kinetic energy so a gas exerts no pressure and has no volume.
- The absolute temperature scale starts at absolute zero.
- Charles' law states that the volume of a gas is directly proportional to its temperature in kelvins, provided the pressure and amount of gas remain constant.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}, \text{ or } \frac{V}{T} = \text{constant}$$

11.8 Questions

1. Describe what theoretically happens to a sample of matter cooled to absolute zero. K/U
2. Convert the following temperatures to degrees Celsius or kelvins as required: K/U
   (a) A flame from a match can have temperatures in the range of 1700 K.
   (b) A household incandescent light bulb has a tungsten filament that can reach over 2327 °C.
   (c) Helium becomes a liquid below −268 °C.
   (d) Oxygen was first made into a solid at a temperature near 54 K.
3. A balloon has a volume of 0.57 L at 22 °C. If it is plunged into a container of liquid nitrogen at 77 K, what will be the volume of the balloon? T/I
4. A 0.300 L sample of argon is collected at 50.0 °C. What is the volume of this gas at standard ambient temperature? T/I
5. Carbon dioxide gas sublimes directly to the solid state at −110 °C. If 25 cm^3 of carbon dioxide is cooled from −20 °C to −110 °C, what will the volume of the gas be just before it sublimes? T/I
6. Molten lava from a volcano includes pockets of gases. A sample of lava contains 0.45 L of water vapour at 1170 °C. What volume will the water vapour occupy when it cools to 120 °C? T/I
7. A sample of argon gas at constant pressure is heated and its volume is measured (**Table 2**). K/U T/I C
   (a) Plot a volume vs. temperature graph of the data in Table 2. Include a title, label your axes, and show the appropriate units.
   (b) Extrapolate the data to find the experimental value of absolute zero in degrees Celsius.
   (c) Compare your experimental value to the actual value of absolute zero. Suggest why these values might not agree perfectly.

Table 2 Experimental Data on a Sample of Argon

Temperature (°C)	Volume (mL)
11	95.3
25	100.0
47	107.4
73	116.1
159	145.0
233	169.8
258	178.1

8. We have discussed a lower limit to temperature: absolute zero. Is there an upper limit? Research to find out. A
9. Jacques Charles used the reaction between metal and acid to produce hydrogen for his balloons. Other people, including the Montgolfier brothers, were experimenting with balloons at the same time. Research another method that was used to provide lift for balloons. T/I
10. When you dive deep into a swimming pool you experience a large pressure change that might cause you some discomfort. Divers run into this situation all the time and are taught a technique, known as the Valsalva manoeuvre, that works well to equalize the pressures. Research the cause of this discomfort and how to perform the Valsalva manoeuvre. A
11. Canadian Mandy-Rae Cruickshank is one of the world's best free divers. In "constant weight" free-diving competitions, she swims as far down as she can (as far as 90 m!) using weights for ballast. Research the extreme sport of free diving and its effects on the human body. A

GO TO NELSON SCIENCE

11.9 The Gas Laws—Boyle's Law, Gay-Lussac's Law, and the Combined Gas Law

Take a deep breath! Now do it again and think about what is happening with your rib cage, your diaphragm, and your lungs. Finally, take a third deep breath and think about how your body moves during breathing to draw oxygen-rich gas into your lungs and, in turn, exhale carbon dioxide-rich gas.

The process of breathing, which you undertake every few seconds, is a demonstration of another gas law. Your lungs are just like large elastic bags that can expand and contract. Your diaphragm is a muscle that extends across the bottom of your rib cage. When you inhale **(Figure 1(a))**, your diaphragm contracts while your rib cage expands. This allows your lungs to occupy more volume and, as a result, the gas in your lungs acquires a lower pressure. Air moves to the area of lower pressure and thus air enters your lungs. When you exhale **(Figure 1(b))**, your diaphragm rises and your rib cage contracts. This decreases the volume of your lungs and consequently creates a high pressure area. This causes the gas to move out of your lungs.

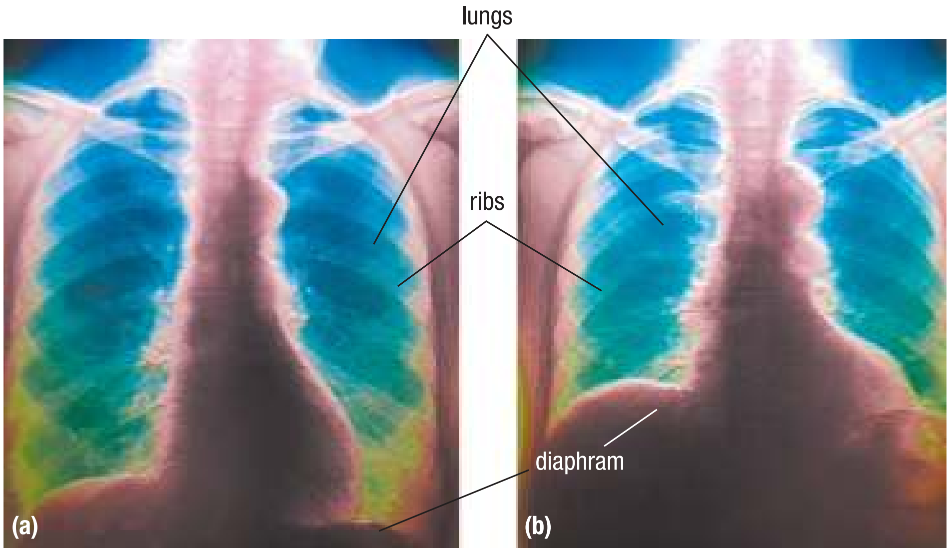


Figure 1 During (a) inhalation the volume of your lungs increases, whereas during (b) exhalation the volume decreases.

Figure 2 Playing a brass or woodwind instrument requires controlled breathing.

Control of the breathing process is of extreme importance in many different types of activities. Singers and wind or brass musicians need to have control of their breathing to achieve the appropriate tone, length of sound and pitch **(Figure 2)**. Competitive swimmers and runners need to monitor their breathing to maximize the delivery of oxygen to their bodies.

Boyle's Law

Boyle's law the statement that as the volume of a gas is decreased, the pressure of the gas increases proportionally, provided that the temperature and amount of gas remain constant; the volume and pressure of a gas are inversely proportional

To recap what happens during the breathing process, as the volume of a gas increases, the pressure of the gas decreases, as long as the temperature remains constant. The breathing process is an example of **Boyle's law**. This law, named for the British scientist Robert Boyle, states that as the volume of a gas is decreased, the pressure of the gas increases proportionally—provided that the temperature and amount of gas remain constant.

Mini **Investigation**

Modelling a Lung

Skills: Performing, Observing, Evaluating, Communicating

SKILLS HANDBOOK A2.1, 2.4

It is quite easy to create a model lung to demonstrate how Boyle's law is connected to the breathing process.

Equipment & Materials: straw; Styrofoam cup or similar container; 2 balloons; clear tape

1. Punch a hole the size of the straw in the bottom of the cup.
2. If the cup is not transparent, cut a small window in the side of the cup and seal the window with clear tape.
3. Place a balloon at the end of the straw.
4. Slide the other end of the straw through the hole in the bottom of the cup so that the balloon is inside the cup.
5. Cut about 1 cm off the top (sealed) end of the other balloon. Tie a knot in the neck of the remaining part. Stretch the balloon across the open end of the cup (**Figure 3**).
6. Pull down on the balloon cover (the diaphragm) and, through the window, observe what happens to the balloon inside the cup. Record your observations.

A. Summarize the effect of expanding the volume of the gas on the balloon inside. Explain, referring to the KMT. K/U

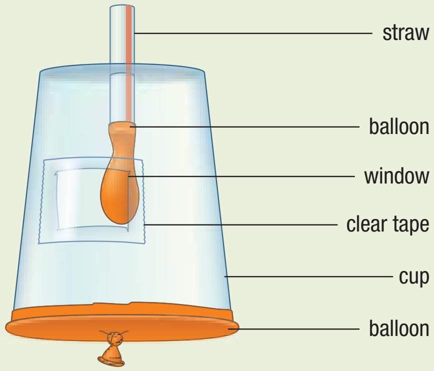


Figure 3

B. Connect and communicate this model to the body parts and processes used in breathing. K/U C

C. How accurate is this model? List some strengths and some weaknesses of this model. T/I

Volume–Pressure Relationships

SKILLS HANDBOOK A6. 4

We will use the data in **Table 1** to investigate the relationship between pressure and the volume of a gas. As usual, a mathematical relationship is easier to understand when plotted as a graph (**Figure 4**). In this case, the volume of the gas is the manipulated (independent) variable and the pressure of the gas is the responding (dependent) variable. Note that the temperature and the amount of gas remain constant throughout this investigation.

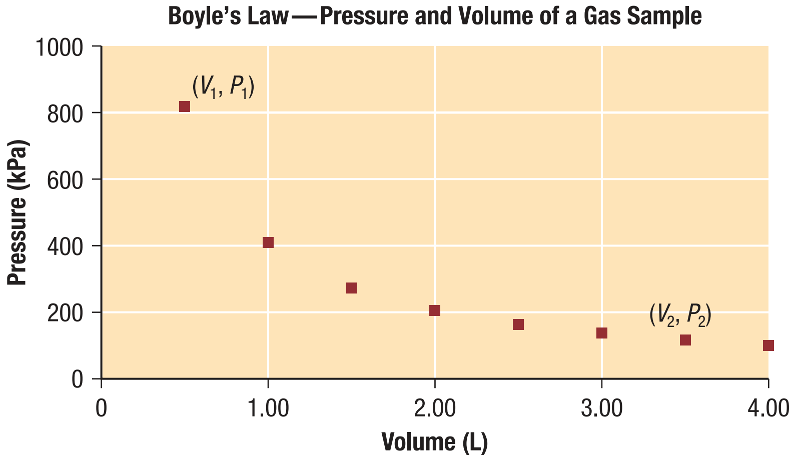


Figure 4 Pressure versus volume data for a closed sample of gas. Temperature and amount of gas have been kept constant.

Table 1 Pressure versus Volume Data for a Closed Sample of Gas

Volume, V (L)	Pressure, P (kPa)	PV (L•kPa)	$1/P$ (kPa^{-1})
0.500	818	409	0.00122
1.00	411	411	0.00243
1.50	273	410	0.00366
2.00	205	410	0.00488
2.50	164	410	0.00610
3.00	137	411	0.00728
3.50	117	410	0.00850
4.00	101	404	0.00990

The data in Figure 4 seem to show an inverse relationship. We can check this by plotting the inverse of pressure ($1/P$) against volume. If the pressure and volume of a gas are inversely related, then the graph of $1/P$ and V will produce a straight line. The graph in **Figure 5** on the next page confirms that there is indeed an inverse relationship between pressure and volume (at constant temperature and amount of gas).

Investigation 11.9.1

The Relationship between Volume and Pressure (p. 564)
In this controlled study you will manipulate one variable (volume) and measure the responding variable (pressure). You will control temperature and amount of gas.

LEARNING TIP

Inverse Relationship
When one variable is changed (increased or decreased) and the other responds in the opposite manner, the two variables are inversely related (or inversely proportional) to each other. As a real-world example, as the number of people buying tickets to a concert increases, the number of available seats decreases. These variables are inversely proportional.

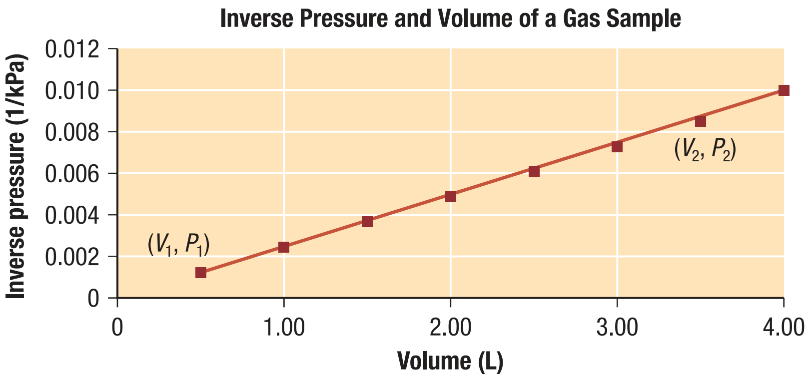


Figure 5 When we plot the inverse of pressure against volume (at constant temperature and amount of gas), the graph indicates a direct relationship.

The data in the *PV* column of Table 1 (page 555) show that the product of pressure and volume gives a constant value. Therefore, the product of any two points in Figure 4 should equal a constant, *k*. In this case, the constant is equal to 409 kPa·L, as follows:

$$P_1V_1 = k \qquad\qquad P_2V_2 = k$$

$$818 \text{ kPa} \times 0.500 \text{ L} = 409 \text{ kPa·L} \qquad 117 \text{ kPa} \times 3.50 \text{ L} = 409 \text{ kPa·L}$$

Since the two equations above both equal a constant, *k*, then, in general

$$P_1V_1 = P_2V_2$$

This relationship is the mathematical expression of Boyle's law.

Boyle's Law

$P_1V_1 = P_2V_2$, or $PV = k$ (k is a constant)

The volume of a gas is inversely proportional to its pressure when the amount of gas and the temperature remain constant.

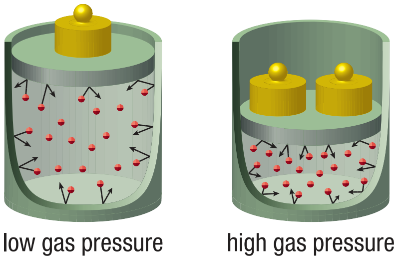


Figure 6 As the volume of the gas decreases, the pressure exerted by the gas increases.

It is important to picture what is happening to the gas molecules as we explore Boyle's law. Consider a fixed amount of a confined gas at a constant temperature. If we were to reduce the size of the container (decreasing the volume), the entities of gas would have less space in which to move. In a reduced space the entities would collide with each other and with the walls of the container more often. The pressure of the gas on the container would increase (**Figure 6**).

Mini Investigation

Cartesian Diver

Skills: Performing, Observing, Analyzing, Communicating — SKILLS HANDBOOK A2.4

A Cartesian diver is named for René Descartes, a famous scientist and mathematician (1592–1650). The "diver" in this investigation is a medicine dropper (**Figure 7**).

Equipment and Materials: empty water or pop bottle with cap; medicine dropper; water

1. Fill the pop bottle with tap water to within a few centimetres of the top.
2. Half-fill a glass medicine dropper with water.
3. Place the medicine dropper in the pop bottle. Adjust the volume of water in the medicine dropper until it floats upright. Cap the bottle securely.
4. Slowly squeeze the pop bottle. Observe the effect.

A. Explain why the diver initially floats. K/U T/I
B. What happened when you applied pressure to the pop bottle? Why? T/I A
C. How might this investigation be related to how submarines dive underwater? T/I A

Figure 7 A Cartesian diver in action

Gay-Lussac's Law

It is the first really cold morning in early winter. You go to start your car. You notice that your tire pressure indicator is on. Yesterday the tire pressure was fine. You check your tire pressure and it is indeed below the recommended level in all four tires. What happened? As the temperature outside decreased, your tire pressure also decreased.

The data in **Table 2** were obtained by measuring the pressure of a constant volume and amount of gas as the temperature of the gas was changed. Temperature, the manipulated variable, is plotted along the *x*-axis (**Figure 8**). Pressure, the responding variable, is plotted along the *y*-axis. The amount of gas and the volume of the gas remain constant throughout this experiment. As the temperature of the gas is increased, the pressure of the gas also increases. There is a direct relationship between temperature and pressure. This relationship is now known as **Gay-Lussac's law**: the pressure of a gas increases proportionally as its temperature increases.

Table 2 Pressure versus Temperature Data for a Fixed Gas Sample

Temperature (K)	Pressure (kPa)
129	111
301	242
403	323
501	429
601	517
704	613

Gay-Lussac's law the statement that as the temperature of a gas increases, the pressure of the gas increases proportionally, provided that the volume and amount of gas remain constant; the temperature and pressure of a gas are directly proportional

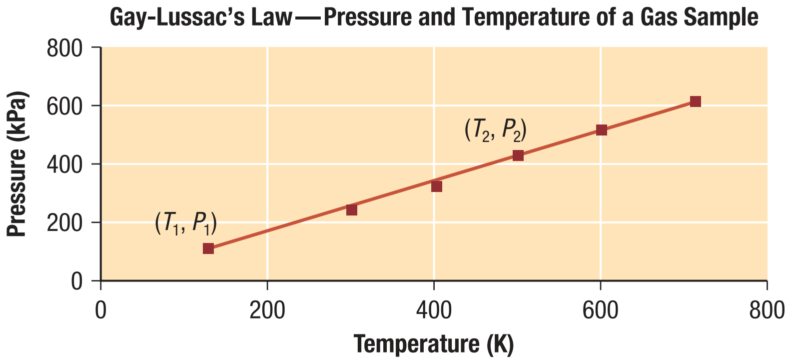


Figure 8 A graph showing the relationship between pressure and temperature for a fixed sample of gas. Volume and amount of gas are held constant.

As in the case of Boyle's law, we may analyze two data points from the plot of pressure versus temperature: (T_1, P_1) and (T_2, P_2). The values of these points are (129, 111) and (501, 425) respectively. If we divide the value of pressure, *P* (*y* value), by temperature, *T* (*x* value), for these two points, we obtain the following results:

For (T_1, P_1)

$$\frac{P_1}{T_2} = \frac{111 \text{ kPa}}{129 \text{ K}} = 0.86 \text{ kPa/K}$$

For (T_2, P_2)

$$\frac{P_2}{T_2} = \frac{429 \text{ kPa}}{501 \text{ K}} = 0.86 \text{ kPa/K}$$

Since the above ratios are equal we can conclude that

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \text{ and } \frac{PT}{T} = k \text{ (constant)}$$

This is the mathematical expression of Gay-Lussac's law.

Gay-Lussac's Law

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}, \text{ or } PT = k \text{ (constant)}$$

The pressure of a gas is directly proportional to its temperature when the amount of gas and volume remain constant.

From our understanding of kinetic molecular theory we know that entities have greater kinetic energy at higher temperatures. With greater kinetic energy, the molecules are more likely to collide with other molecules and with the walls of the container. The result is increased pressure from more frequent collisions with the walls of the container.

Tutorial 1 Using the Gas Laws

The gas laws that we have discussed thus far—Charles' law, Boyle's law, and Gay-Lussac's law—are extremely useful for predicting volumes, temperatures, and/or pressures when various changes have been made to a sample of gas. This tutorial will help you to solve problems involving these gas laws.

Sample Problem 1: Pressure and Volume

A weather balloon is filled with 60.0 L of hydrogen gas at sea level pressure (101.3kPa). It then rises to 900 m above Earth's surface. The atmospheric pressure at this altitude is 90.6 kPa. What is the volume of the balloon at this altitude? Assume there is no change in temperature or amount of gas.

Given: Identify the known variables. Note which variables are being held constant.

initial pressure, $P_1 = 101.3$ kPa

final pressure, $P_2 = 90.6$ kPa

initial volume, $V_1 = 60.0$ L

The amount of gas and the temperature remain constant.

Required: final volume, V_2

Analysis: Use Boyle's law.

$$P_1V_1 = P_2V_2$$

Solution:

Step 1. Rearrange the equation to isolate the unknown variable.

$$V_2 = \frac{P_1V_1}{P_2}$$

Step 2. Substitute given values (including units) into the equations and solve.

$$V_2 = \frac{101.3\ \cancel{\text{kPa}} \times 60.0\ \text{L}}{90.6\ \cancel{\text{kPa}}}$$

$$= 67.1\ \text{L}$$

Statement: The volume of the balloon is 67.1 L at 900 m altitude. Check your answer. Does it make sense? Are the units appropriate?

Sample Problem 2: Pressure and Temperature

A sample of gas is stored in a reinforced steel container at −115 °C, at a pressure of 39.9 kPa. If the pressure reaches 60.8 kPa, what is the final Celsius temperature?

Given: initial pressure, $P_1 = 39.9$ kPa

final pressure, $P_2 = 60.8$ kPa

initial temperature, $t_1 = -115$ °C

The volume and amount of gas remain constant.

Required: final temperature, T_2

Analysis: Use Gay-Lussac's law.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Solution:

Step 1. Convert temperature values to kelvins.

$$T_1 = t_1 + 273$$

$$= -115 + 273$$

$$T_1 = 158\ \text{K}$$

Step 2. Rearrange the equation to isolate the unknown variable.

$$T_2 = \frac{P_2 T_1}{P_1}$$

Step 3. Solve the equation (including units).

$$T_2 = \frac{60.8 \cancel{\text{kPa}} \times 158 \text{ K}}{33.9 \cancel{\text{kPa}}}$$

$$T_2 = 241 \text{ K}$$

Step 4. Convert the final temperature to the required units.

$$t_2 = T_2 - 273$$
$$= 241 - 273$$
$$t_2 = -32 \text{ °C}$$

Statement: The temperature of the gas in the container is −32 °C.

Practice

SKILLS HANDBOOK A6

1. Helium gas has a volume of 8.25 L at 446 kPa. What pressure must be applied to the gas when it occupies 12 L? T/I [ans: 307 kPa]
2. A 1.75 L sample of ammonia gas increases in volume to 6.50 L when the pressure reaches 2.84 kPa. What was the original pressure of this gas? T/I [ans: 10.5 kPa]
3. Soccer balls are typically inflated to between 60 and 110 kPa. A soccer ball is inflated indoors with a pressure of 85 kPa at 25 °C. If it is taken outside, where the temperature on the playing field is −11.4 °C, what is the pressure of the gas inside the soccer ball? T/I [ans: 75 kPa]

The Combined Gas Law

In each of the gas laws that we have studied so far, one variable was manipulated, one was allowed to respond, and all others were kept constant (controlled).

Charles' law: $\frac{V}{T}$ = constant (n and P are controlled)

Boyle's law: PV = constant (n and T are controlled)

Gay-Lussac's law: $\frac{P}{T}$ = constant (n and V are controlled)

Now we can combine Charles' law, Boyle's law, and Gay-Lussac's law into a single law. The **combined gas law** describes the relationship between volume, temperature, and pressure for any fixed amount of gas. The combined equation is

$$\frac{PV}{T} = \text{constant}$$

In other words, the product of the pressure and volume of a gas divided by its absolute temperature is a constant as long as the amount of gas is kept constant. This relationship can be expressed in a convenient form for calculations involving changes in volume, temperature, or pressure for a fixed amount of gas.

combined gas law the statement that the product of the pressure and volume of a gas sample is proportional to its absolute temperature in kelvins

The Combined Gas Law

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}, \text{ or } \frac{PV}{T} = \text{constant}$$

The product of the pressure and volume of a gas divided by its absolute temperature is a constant as long as the amount of gas is kept constant.

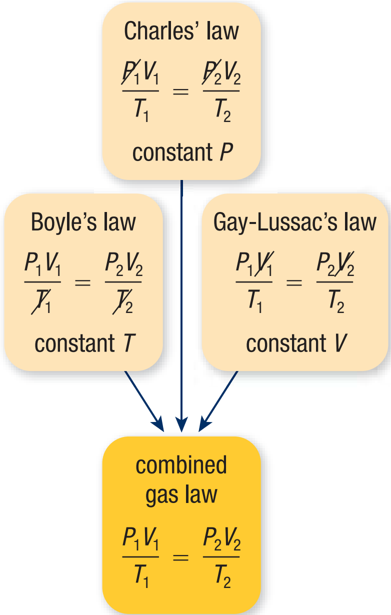


Figure 9 The combined gas law is a combination of three gas laws: Charles' law, Boyle's law, and Gay-Lussac's law. We can use it to find any of the other laws.

LEARNING TIP

Mathematical Mess

As you rearrange the combined gas law, things can get messy quickly. Be sure to take one step at a time and be careful as you perform mathematical operations. Do a quick check to make sure you rearranged the equation correctly. If you include the units in the rearranged equation, you should be left with the correct units after the others have cancelled out.

Tutorial 2 Using the Combined Gas Law

The combined gas law is very useful in predicting the effect of changes in two of the gas variables on the third when the amount of gas remains constant. In addition you can use the combined gas law to derive any of the other three gas laws that we have examined **(Figure 9)**.

Sample Problem 1: Using the Combined Gas Law

A sample of carbon dioxide gas, $CO_2(g)$, occupies a volume of 25.0 L when the pressure is 125 kPa and the temperature is 25 °C. Calculate the volume occupied by this same quantity of carbon dioxide at STP.

Given: Identify the known variables. Decide whether the known variables are the initial or final values. Note which variables are being held constant. Remember that STP stands for standard temperature and pressure and is equal to a temperature of 25 °C and 101.3 kPa of pressure.

initial pressure, $P_1 = 125$ kPa

initial volume, $V_1 = 25.0$ L

initial temperature, $t_1 = 25$ °C

final pressure, $P_2 = 101.3$ kPa

final temperature, $T_2 = 273$ K

Amount of gas remains constant.

Required: final volume, V_2

Analysis: Use the combined gas law.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Solution:

Step 1. Convert temperature values to kelvins.

$$T_1 = 25 + 273$$
$$= 298 \text{ K}$$

Step 2. Rearrange the equation to isolate the unknown variable.

$$V_2 = \frac{P_1 V_1 T_2}{P_2 T_1}$$

Step 3. Substitute given values (including units) into the equations and solve.

$$V_2 = \frac{125 \cancel{\text{kPa}} \times 25.0 \text{ L} \times 273 \cancel{\text{K}}}{101.3 \cancel{\text{kPa}} \times 298 \cancel{\text{K}}}$$

$$V_2 = 28.3 \text{ L}$$

Statement: The volume of the gas at STP is 28.3 L.

Practice

SKILLS HANDBOOK A6

1. A balloon at the top of Mount Logan occupies a volume (V_1) of 775 mL at a temperature of −28 °C (T_1) and a pressure of 92.5 kPa (P_1). What is the pressure (P_2) at the bottom of the mountain if the same balloon has a volume (V_2) of 825 mL at a temperature (T_2) of 15 °C? T/I [ans: 102 kPa]
2. A researcher heated a 2.75 L sample of helium gas at 99.0 kPa from 21.0 °C to 71.0 °C, and recorded that the pressure changed to 105 kPa. Calculate the final volume of the gas. T/I [ans: 3.03 L]
3. A 450 mL sample of propane gas at 253 kPa and 15 °C was compressed to 310 mL at a pressure of 405 kPa. Calculate the final temperature in Celsius. T/I [ans: 45 °C]

Research This

Investigating Gas Law Simulations

Skills: Researching, Evaluating, Communicating, Defining the Issue, Identifying Alternatives

SKILLS HANDBOOK A5.1

Animations and simulations are useful tools for exploring the gas laws. Simulations allow us to "see" gas molecules and how they behave under various situations (**Figure 10**). In this activity you will explore and evaluate several online simulations.

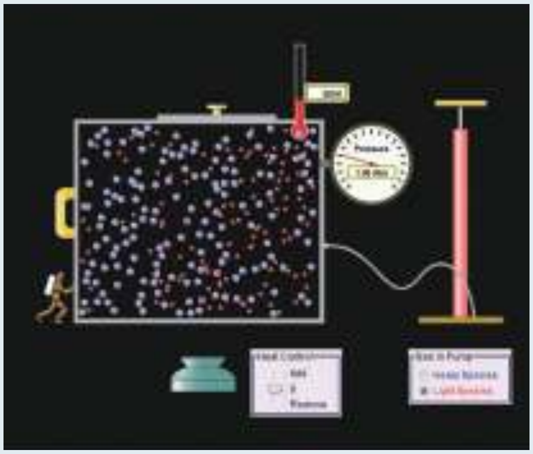


Figure 10

1. Working in a small group, develop criteria to evaluate some online simulations. Create a checklist based on these criteria.
2. Find several simulations for the same gas law.
3. Try each of these simulations, using your checklist to evaluate them.

A. What were the three most important criteria to evaluate the simulations? Briefly describe why these considerations were chosen. T/I A

B. Analyze your collected data. Which online simulation would you recommend? Explain your choice. T/I A

C. What improvements would you like made to the online simulations you evaluated? A

D. In your opinion, how does using these simulations enhance your understanding of gas laws? Explain your answer. K/U

GO TO NELSON SCIENCE

11.9 Summary

- Boyle's law states that as the volume of a sample of gas decreases, the pressure increases as long as temperature and amount of gas remain constant. Pressure and volume are inversely related.

  $P_1V_1 = P_2V_2$, or $PV = \text{constant}$

- Gay-Lussac's law states that the pressure of a gas is directly related to the temperature of the gas when the volume and amount of gas remain constant.

  $\dfrac{P_1}{T_1} = \dfrac{P_2}{T_2}$, or $\dfrac{P}{T} = \text{constant}$

- The combined gas law encompasses Charles' law (Section 11.8), Boyle's law, and Gay-Lussac's law for a controlled amount of gas.

  $\dfrac{P_1V_1}{T_1} = \dfrac{P_2V_2}{T_2}$, or $\dfrac{PV}{T} = \text{constant}$

11.9 Questions

1. Create a concept map of the four gas laws addressed in this chapter. Include descriptions of the laws, their mathematical equations, and the appropriate SI units. K/U C
2. A controlled amount of nitrogen gas initially occupies a volume of 45.2 L. Find the final volume of the nitrogen gas if it is heated from an initial temperature of 280 K to a final temperature of 560 K.
3. The air inside a tire exerts a pressure of 340 kPa at a temperature of 15° C. What will be the new pressure if the air is cooled to a temperature of −15° C? T/I
4. Using the data in **Table 3**, use the appropriate gas law to solve for the unknown quantities. T/I

Table 3 Data for a Constant Amount of Gas

	P_1	V_1	T_1	P_2	V_2	T_2
(a)	101.3 kPa	3.50 L	320 K	?	8.40 L	320 K
(b)	210 kPa	?	415 K	420 kPa	120 mL	415 K
(c)	720 mm Hg	345 mL	420 K	620 mm Hg	?	640 K

5. Neon gas is used in glass tubes to make "neon" lights. The glass must be able to withstand great variations of temperature. One night the temperature is 6.50 °C and the neon gas has a pressure of 130 kPa. The sign is turned on and the neon reaches 670.0 °C. What is the new pressure of the neon gas at this temperature? T/I
6. A compressed-gas tank contains air at 95 kPa at a temperature of 22 °C. The tank can sustain a maximum pressure of 350 kPa. What is the maximum temperature that the tank can withstand? T/I
7. Sometimes biological samples are stored at low temperatures under nitrogen gas. A biological sample in a sealed vessel contains a small volume of nitrogen gas at 101.3 kPa at 20 °C. The sample is stored at –15 °C. What is the final pressure of the nitrogen gas? T/I
8. A glass jar with a volume of 215 cm^3 is designed to withstand a pressure of 253 kPa. The jar is filled with gas at a pressure of 152 kPa at ambient temperature. (See to Section 11.7.) Can the glassware withstand the gas being heated to 200 °C? T/I
9. Your tires are adjusted to 227.5 kPa at 10 °C in the mechanic's garage. You then take your car home and park it outside. The overnight temperature drops to –5 °C. K/U T/I A
   (a) Would you expect the tire pressure to increase or decrease? Explain your answer using the kinetic molecular theory.
   (b) Determine the new tire pressure.
10. A helium balloon escapes from a birthday party. It contained 1250 cm^3 of helium at a pressure of 125 kPa. As the balloon rises it experiences a decrease in pressure to 97 kPa. Assuming no change in temperature, what would be the volume of the balloon at this altitude? T/I
11. Pressurized hydrogen gas is being tested as a fuel for vehicles. Hydrogen used to fuel these cars is stored at great pressure. In one storage tank, 4.50×10^4 L of hydrogen is stored at 2.7574×10^4 kPa. What would be the new pressure if the gas were transferred to a new tank with a volume of 6.00×10^4 L? T/I A
12. We store and use gases under pressure every day. T/I A
   (a) List three familiar examples of gases stored under pressure.
   (b) For each of your examples, list the precautions that you should take when using this product.
   (c) Explain the reason for each of these precautions, referring to the gas laws.
13. A 8.0 mL bubble of gas is released at the bottom of the ocean where the pressure is 685 kPa and the temperature is 12 °C. Calculate the volume of the gas bubble when it reaches the surface of the water, where the pressure is 99.0 kPa and the temperature is 26 °C. T/I
14. A balloon containing helium at 25 °C and a pressure of 112 kPa has a volume of 5.50 L. Calculate the volume of the balloon after it rises 8 km into the upper atmosphere, where the temperature is –34 °C and the outside air pressure is 32.0 kPa. Assume that no helium escapes and that the balloon is free to expand so that the gas pressure within it remains equal to the air pressure outside. T/I
15. A sample of carbon dioxide gas occupies a volume of 893 mL at 44 °C and 116 kPa. At what temperature will this gas occupy a volume of 1.03 L if the pressure is reduced to 102 kPa? T/I

Investigation 11.4.1 CORRELATIONAL STUDY

SKILLS MENU

• Questioning	• Planning	• Observing
• Researching	• Controlling Variables	• Analyzing
• Hypothesizing	• Performing	• Evaluating
• Predicting		• Communicating

Comparing Air Quality in Several Locations

In this investigation you will compare Air Quality Health Index (AQHI) data for several locations. You will also use historical Air Quality Index (AQI) data to investigate air quality trends over extended periods of time.

Purpose

SKILLS HANDBOOK A2.3

To find out whether there is a relationship between human activity and air quality in Canada

Variables

Identify all major variables that you will explore in this study.

Study Design

Part A

In this study you will select four locations in Canada with different levels of human activity. You will develop a criterion for ordering the locations according to their level of human activity. For each location, you will then research air quality data and analyze these data to identify any relationship.

Part B

Research the implementation of a government initiative to improve air quality at one of your locations. Research historical air quality data for this location and follow these data over time. Analyze these data to identify any relationship between the implementation of the initiative and changes in air quality.

Procedure

Part A

1. Choose four Canadian locations to research: your own (if possible), one major city in Ontario, and two others. The locations should have different levels of human activity (such as industry or transportation).
2. Decide on a criterion for ranking the locations and apply your criterion.
3. Go to the AQHI page on Environment Canada's website. Select "Your Local AQHI Conditions" for each of your four locations. Record today's AQHI conditions, for each location.

Part B

SKILLS HANDBOOK A5.1

4. Research a government initiative designed to improve air quality in Ontario. Find out when this initiative was implemented and whether it has changed in any way.
5. Go to the Air Quality Ontario website. Research historical air quality data using the Air Quality Index (AQI). Select at least six different years for the major Ontario city that you are investigating. At least two of the years should be *before* the air quality improvement initiative was implemented, and at least two should be *after* this date. Determine how many days, each month, had "poor" or "very poor" air quality at your chosen location. Record your observations.

Analyze and Evaluate

(a) What variables were recorded in this investigation? T/I

(b) Analyze your data in Part A. T/I A
  (i) Identify any relationship between the variables.
  (ii) Compare your local air quality conditions to conditions at the three other locations.

(c) Analyze your data in Part B. T/I C A
  (i) Plot a graph of the number of days with poor or very poor air quality each month. Show the data for each year using a separate line. Indicate on your graph the year in which the air quality improvement initiative (researched in Step 4) was implemented.
  (ii) Interpret your graph.

(d) Write a paragraph summarizing your findings about the effects of human activity on air quality in Ontario and the effectiveness and value of initiatives to improve air quality. Refer to specific data to support your conclusions. T/I C A

(e) What statements, if any, can you make about the cause of poor air quality? T/I

(f) In this investigation you used both AQHI and AQI data. Comment on the difference between them, and the advantages and disadvantages of each. T/I A

Apply and Extend

(g) Select another major city in Ontario. Search the AQI data and find four days that had poor air quality readings for this location within the last year. Also go to Environment Canada's National Climate Data and Information Archive to find the weather report for those days. Record answers to the following questions: T/I C A
  (i) At which time of the day were air quality readings the worst? (Select more readings to view data throughout the day.)
  (ii) Which categories of air quality readings are listed as "poor"?
  (iii) What was the weather like on those days?
  (iv) What trends, if any, do you notice? If possible, write a hypothesis linking the variables.

(h) Choose a location 20 to 40 km away from the major city that you chose in (g). Compare the air quality for the same days. What trends, if any, do you notice? Explain. T/I C A

(i) Find three locations in Ontario that did not show poor air quality results for last year. Why do you think these locations did not show any cases of poor air quality? T/I C A

GO TO NELSON SCIENCE

Investigation 11.9.1 CONTROLLED EXPERIMENT

The Relationship between Volume and Pressure

SKILLS MENU

• Questioning	• Planning	• Observing
• Researching	• Controlling Variables	• Analyzing
• Hypothesizing	• Performing	• Evaluating
• Predicting		• Communicating

In this investigation we will observe the relationship between the volume and the pressure of a gas. You will be provided with the necessary materials to carry out this investigation. You will design your own procedure based on the materials and information provided.

Testable Question

What is the effect of increasing pressure on the volume of a confined gas?

Hypothesis

SKILLS HANDBOOK A2.2

Consider what you know about the behaviour of gases from the kinetic molecular theory and real-world examples. Develop a hypothesis for this relationship, including explanations.

Variables

List the variables that will be controlled in this experiment. Identify which variable will be manipulated and which will be responding.

Experimental Design

You will vary the pressure of a gas trapped inside a large syringe. You will change the pressure by changing the force applied to the syringe plunger. You will record the corresponding volume for each pressure reading.

Equipment and Materials

SKILLS HANDBOOK A1.2

- Boyle's law apparatus
- 35 mL plastic syringe
- large rubber stopper or rubber block with a hole bored to fit the syringe
- 5 identical textbooks or 5 standard masses (1 kg)
- retort stand with clamp
- balance

Alternatively:

- digital gas pressure sensor and syringe

Take care not to drop the masses on your feet.

Procedure

1. Using the list of equipment and materials and **Figure 1**, develop and write a procedure to test the effects of changing pressure on the volume of air contained in the syringe. When you have your teacher's approval, proceed with your investigation.

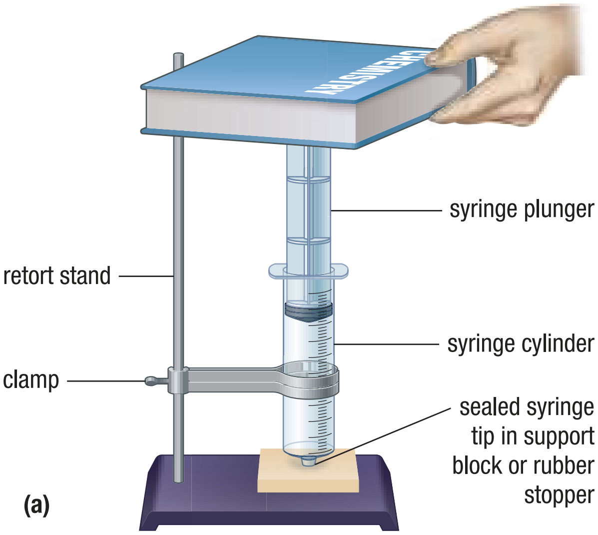


Figure 1 (a) traditional experimental setup (b) electronic equipment

Observations

Record your data in a table. C

Analyze and Evaluate

(a) What variables were kept constant in this investigation? Why? T/I

(b) Did your procedure allow you to collect suitable data so that you could test your hypothesis? Explain. T/I

(c) Plot a graph of volume (in mL or cm^3) versus pressure (in number of books). Draw a line or curve of best fit to show the relationship between these variables. T/I C

(d) Answer the testable question. T/I C

(e) Compare your answer to the testable question in (d) to your hypothesis. Account for any differences. T/I

(f) Select two points on or near the line of best fit. For each point determine the product of $P \times V$. Do these values correspond to what is predicted by Boyle's law? Explain. K/U T/I

(g) What are some of the sources of error in this experiment? Explain how each error would affect the outcome of the experiment. T/I

(h) How could the design be improved? Be specific in your answer. T/I

Apply and Extend

(i) How would your results have changed if, during the experiment, the temperature in the room had increased substantially? T/I A

(j) Pressure exerted by water at the ocean floor can be very high. The Mariana Trench has a depth of 11 033 m. At this depth the pressure exerted on an object is over 1.03×10^5 kPa, about 1000 times greater than atmospheric pressure. T/I A

(i) What would happen if you tried to scuba dive at this depth?

(ii) Outline what design considerations you would incorporate into an exploration vessel for this region.

(iii) Research whether there is any life at these depths.

GO TO NELSON SCIENCE

CHAPTER 11 SUMMARY

Summary Questions

1. Look back at the points listed in the margin on page 514. For each point, write three or four questions. Answer these questions using the information in the chapter. Include examples and diagrams in the answers. Share your questions and answers with classmates as a study aid.
2. Look back at the Starting Points questions on page 514. Answer these questions using what you have learned in this chapter. Compare your latest answers with those that you wrote at the beginning of the chapter. Note how your answers have changed.

Vocabulary

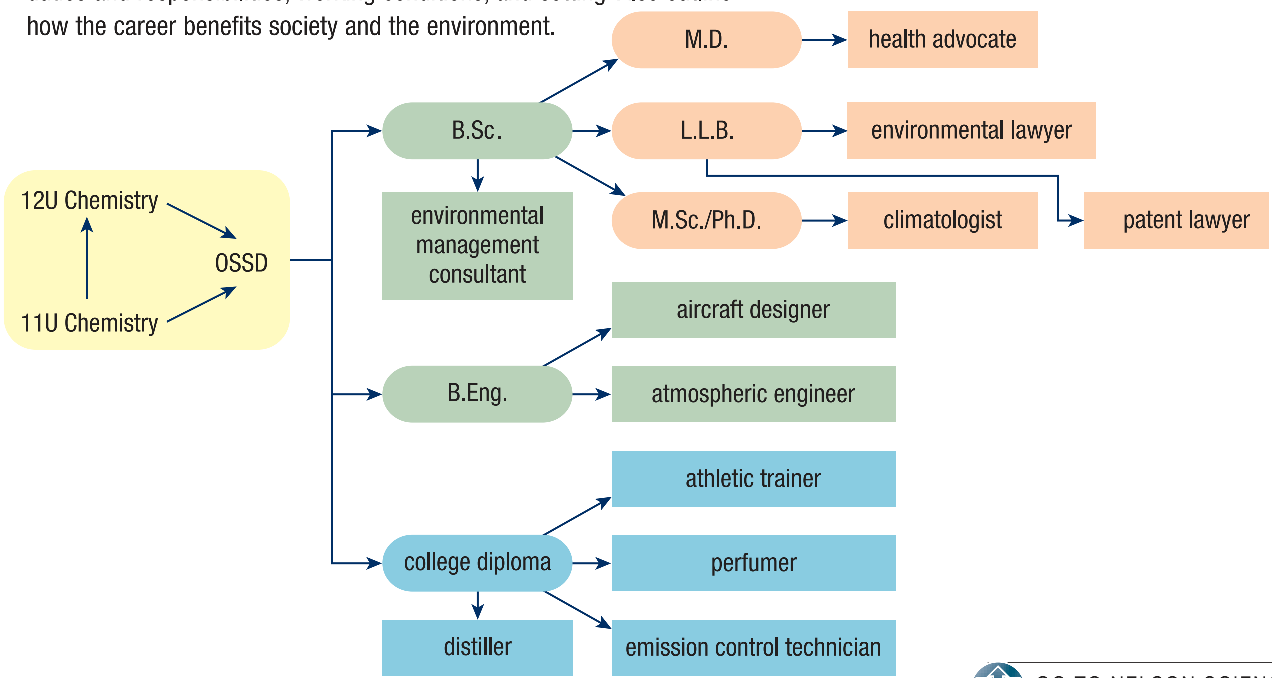
Brownian motion (p. 517)
kinetic molecular theory (p. 517)
kinetic energy (p. 517)
temperature (p. 518)
greenhouse gas (GHG) (p. 521)
greenhouse effect (p. 522)
carbon sequestration (p. 523)
photochemical smog (p. 528)
particulate matter (p. 528)
Air Quality Health Index (AQHI) (p. 532)
off-gassing (p. 534)
pressure (*P*) (p. 541)
atmospheric pressure (p. 542)
standard pressure (p. 542)
standard temperature and pressure (STP) (p. 542)
standard ambient temperature and pressure (SATP) (p. 542)
absolute zero (p. 548)
Kelvin temperature scale (p. 548)
absolute temperature (p. 548)
Charles' law (p. 549)
Boyle's law (p. 554)
Gay-Lussac's law (p. 557)
combined gas law (p. 559)

CAREER PATHWAYS

SKILLS HANDBOOK A7

Grade 11 Chemistry can lead to a wide range of careers. Some require a college diploma or a B.Sc. degree. Others require specialized or post-graduate degrees. This graphic organizer shows a few pathways to careers mentioned in this chapter.

1. Select two careers, related to the gas state and gas laws, that you find interesting. Research the educational pathways that you would need to follow to pursue these careers. What is involved in the required educational programs? Prepare a brief report of your findings.
2. For one of the two careers that you chose above, describe the career, main duties and responsibilities, working conditions, and setting. Also outline how the career benefits society and the environment.



GO TO NELSON SCIENCE

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

For each question, select the best answer from the four alternatives.

1. Which of the following shows the layers of the atmosphere in order of lowest (ground level) to highest? (11.2) K/U
   (a) troposphere, stratosphere, mesosphere, thermosphere
   (b) troposphere, mesosphere, stratosphere, thermosphere
   (c) stratosphere, mesosphere, thermosphere, troposphere
   (d) thermosphere, troposphere, mesosphere, stratosphere
2. The gas in the stratosphere that protects life on Earth from harmful solar ultraviolet radiation is
   (a) methane
   (b) ozone
   (c) nitrogen monoxide
   (d) sulfur trioxide (11.2) K/U
3. Suppose that the oxygen, water vapour, and carbon dioxide are removed from a sample of air. The two gases that would remain in significant quantities are
   (a) hydrogen and methane
   (b) ozone and nitrogen monoxide
   (c) neon and helium
   (d) nitrogen and argon (11.3) K/U
4. Which of the following substances gives smog its reddish-brown colour? (11.4) K/U
   (a) carbon monoxide
   (b) sulfur dioxide
   (c) nitrogen dioxide
   (d) ammonia
5. Which of the following indoor air pollutants poses a significant health risk because of its radioactivity? (11.5) K/U
   (a) radon
   (b) methanal (formaldehyde)
   (c) carbon monoxide
   (d) particulate matter from combustion
6. A pressure of 1.20 atm is equivalent to
   (a) 633 Torr
   (b) 122 kPa
   (c) 84.4 N/m^2
   (d) 63.3 mm Hg (11.7) K/U
7. As altitude increases, the number of entities in 1 m^3 of air
   (a) decreases
   (b) increases
   (c) remains constant
   (d) decreases at first, then increases (11.7) K/U
8. If the absolute temperature of a given amount of gas is doubled at constant pressure, the new volume will be
   (a) unchanged
   (b) half as great
   (c) twice as great
   (d) three times as great (11.8) K/U
9. If the pressure of a given amount of gas is doubled at constant temperature, the new volume will be
   (a) unchanged
   (b) half as great
   (c) twice as great
   (d) three times as great (11.9) K/U

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

10. The least compressible states of matter are the gaseous state and the solid state. (11.1) K/U
11. When a substance is in the liquid state, its entities can only undergo vibrational motion. (11.1) K/U
12. Oxygen is the most abundant gas in Earth's atmosphere. (11.2) K/U
13. A low AQHI value indicates that outdoor activities should be avoided by people who have respiratory problems. (11.4) K/U A
14. A car engine is more polluting when idling than when operating at highway speeds. (11.6) K/U
15. Children have a greater risk of health effects from air pollution because they breathe more frequently and inhale more pollutants per kilogram of body mass. (11.6) K/U
16. Pressure is defined as area per unit force. (11.7) K/U
17. A temperature in degrees Celsius is converted to kelvins by adding 273. (11.8) K/U
18. In the equation below, the temperatures must be expressed in degrees Celsius. (11.9) K/U

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

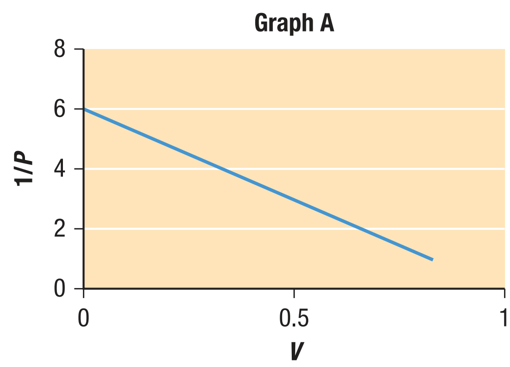
To do an online self-quiz,
GO TO NELSON SCIENCE

Knowledge

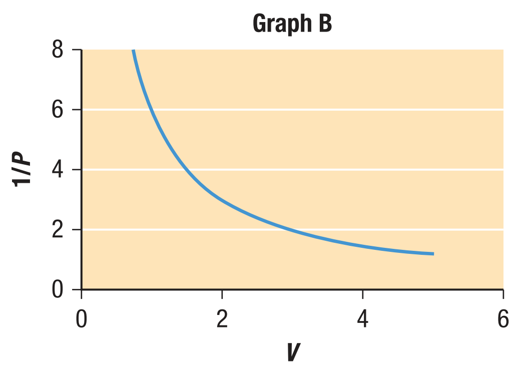
For each question, select the best answer from the four alternatives.

1. Which of the following is the layer in which 99 % of Earth's atmospheric water vapour is found? (11.2) K/U
   (a) thermosphere
   (b) mesosphere
   (c) stratosphere
   (d) troposphere
2. Which shows the proper ordering of atmospheric gases from most abundant to least abundant by volume? (11.2) K/U
   (a) N_2, Ar, CO_2, O_2
   (b) N_2, O_2, Ar, CO_2
   (c) O_2, N_2, Ar, CO_2
   (d) Ar, CO_2, O_2, N_2
3. In 1952, terrible smog killed over 4000 people in
   (a) Los Angeles
   (b) Beijing
   (c) Budapest
   (d) London (11.4) K/U
4. Approximately how many premature deaths in Ontario in 2008 were attributed in part to poor air quality? (11.4) K/U
   (a) 500
   (b) 1000
   (c) 4000
   (d) 9000
5. Which of the following indoor air pollutants is a biological pollutant? (11.5) K/U
   (a) mould
   (b) radon
   (c) carbon monoxide
   (d) VOCs
6. Which of the following chemical equations shows how smelting operations contribute to acid precipitation? (11.4) K/U T/I
   (a) $Cu_2S(s) + O_2(g) \rightarrow 2\ Cu(l) + SO_2(g)$
   (b) $C_8H_{18}(l) + 11\ O_2(g) \rightarrow C(s) + CO(g) + 6\ CO_2(g) + 9\ H_2O(g)$
   (c) $CO(g) + 2\ O_2(g) \rightarrow CO_2(g) + O_3(g)$
   (d) $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$
7. According to Health Canada, which of the following indoor air pollutants is the leading cause of lung cancer among non-smokers? (11.5) K/U
   (a) methanal (c) carbon monoxide
   (b) radon (d) VOCs
8. Which of the following graphs could represent how inverse pressure varies with volume for a fixed amount of gas at constant temperature? (11.9) K/U

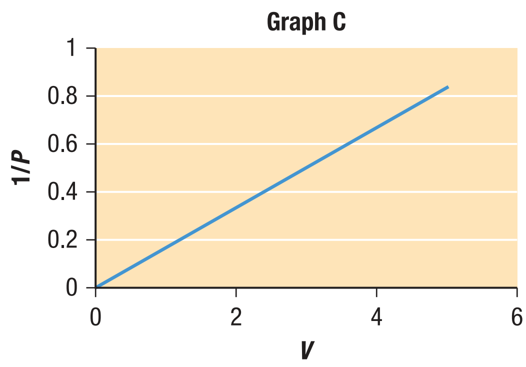
(a)



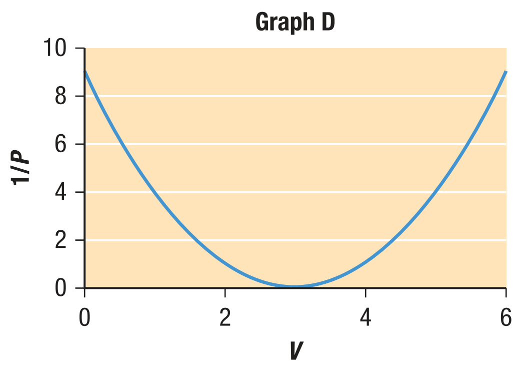
(b)



(c)



(d)



9. The shore of the Dead Sea is the lowest land location in the world, with an elevation of 413 m below sea level. Which of the following is most likely the average atmospheric pressure at the surface of the Dead Sea? (11.7) K/U
   (a) 80 kPa
   (b) 99 kPa
   (c) 101 kPa
   (d) 106 kPa
10. Which of the following quantities is equivalent to 1 N/m^2? (11.7) K/U
   (a) 1 Pa
   (b) 1 Torr
   (c) 1 atm
   (d) 1 mm Hg
11. Which of the following temperature changes would cause a fixed amount of gas at constant pressure to double in volume? (11.8) K/U
   (a) increase from 20 °C to 40 °C
   (b) decrease from 60 °C to 30 °C
   (c) increase from 150 K to 300 K
   (d) decrease from 400 K to 200 K

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

12. When a substance changes from a solid to a liquid and then from a liquid to a gas, the degree of order in its molecules increases. (11.1) K/U
13. With an increase in thermal energy of its entities, the liquid in a thermometer expands. (11.1) K/U
14. The layers of the atmosphere in which temperature rises with increasing altitude are the troposphere and the stratosphere. (11.2) K/U
15. Carbon dioxide gas in the atmosphere is converted by bacteria into a form that plants can readily absorb through their roots. (11.2) K/U
16. When a fire occurs in a residential building, the main cause of death will typically be burns. (11.4) K/U A
17. About half of the fine particulate matter in Ontario's air comes from the United States. (11.4) K/U
18. The greatest proportion of carbon monoxide air pollution in Ontario comes from the production of metals. (11.4) K/U
19. For a sample of gas in a sealed chamber of constant volume, the pressure of the gas will increase as thermal energy is supplied to raise its temperature. (11.9) K/U

Match each term on the left with the most appropriate description on the right.

20. (a) troposphere
   (b) stratosphere
   (c) mesosphere
   (d) thermosphere

   (i) layer where the ozone absorbs solar UV rays, resulting in warming
   (ii) layer where all weather occurs
   (iii) layer where low-pressure gases receive and absorb the most solar energy
   (iv) layer starting at 70 km that cools with increasing altitude (11.2, 11.7) K/U

Write a short answer to each question.

21. For each property, write "solid," "liquid," or "gas." Note that in some cases, you might write more than one of these terms. (11.1) K/U
   (a) virtually incompressible
   (b) takes the shape of the container, but has a definite volume
   (c) definite shape and volume
   (d) easily compressed
   (e) takes the shape and volume of the entire container
   (f) flows readily
22. Explain how supplying thermal energy to a solid can cause it to melt. (11.1) K/U
23. What change to the periodic table ultimately resulted from Sir William Ramsay's discovery of argon? (11.3) K/U
24. In what layer of the atmosphere is ozone a bad thing? Explain your answer. (11.2, 11.4) K/U
25. What organism lives in linens, mattresses, pillows, and carpets and feeds on dander? (11.5) K/U
26. Copy and complete **Table 1** in your notebook. Assume that the amount of gas and temperature are kept constant. (11.9) K/U

Table 1 Observations of a Sample of Gas

Pressure (kPa)	Volume (L)
1.00	6.00
1.50	
2.00	
	2.00
4.00	
	1.00

Understanding

27. At room temperature, iodine, I_2, is a solid, chlorine, Cl_2, is a gas, and bromine, Br_2, is a liquid. List these substances in order of decreasing strength of intermolecular attraction. State reasons for your listing. (11.1) K/U
28. Consider samples of methane, $CH_4(g)$, water, $H_2O(l)$, and sucrose, $C_{12}H_{22}O_{11}(s)$, at room temperature. Compare and contrast the motions of the molecules in these samples. (11.1) K/U
29. How could mowing the lawn with a push mower help in the fight against climate change? (11.2) K/U A
30. What evidence do we have that ozone high in the atmosphere is a very efficient absorber of ultraviolet radiation from the Sun? (11.2) K/U
31. What volume of argon could be obtained from 2.00 L of air? (11.2) K/U
32. If you hold a piece of ceramic tile above a yellow candle flame, the tile soon becomes coated with a black deposit. What does this observation demonstrate about incomplete combustion and air pollution? (11.4) K/U
33. Explain how setting a home's air conditioning temperature a few degrees higher could help urban air quality in the summer months. (11.4) K/U A
34. Ozone is a component of photochemical smog. The chemical equations below represent the formation of ozone. Refer to these reactions as you answer the following questions: (11.4) K/U A

    I. $N_2(g) + O_2(g) \rightarrow 2\,NO(g)$

    II. $2\,NO(g) + O_2(g) \rightarrow 2\,NO_2(g)$

    III. $NO_2(g) \xrightarrow{\text{UV radiation}} NO(g) + O(g)$

    IV. $O(g) + O_2(g) \rightarrow O_3(g)$

    (a) Why does ozone formation stop at night?
    (b) Why does less ozone form on days when the traffic is very light?
35. Explain how indoor air could be more polluted than outdoor air. (11.5) K/U A
36. Why would cracked exterior basement walls be of particular concern in a region that has relatively high levels of uranium in the soil? (11.5) K/U A
37. Explain why idling cars lined up to pick up children at an elementary school are a health concern. (11.6) K/U A
38. Compare the number of entities in 1 m^3 of air at sea level to the number in 1 m^3 of air at an elevation of 3000 m. Explain your answer. (11.7) K/U
39. When a sample of water freezes, do the water molecules become motionless? Explain your answer. (11.1, 11.8) K/U T/I
40. Why should you store unpopped popcorn kernels in sealed containers, especially if the air is very dry where you live? (11.8) K/U A
41. 4.000×10^3 L of propane gas is held in a tank at 25 °C. The tank has a moveable diaphragm to keep the pressure constant at 200 kPa. If the temperature falls to −5 °C on a cold winter day, what volume will the gas occupy? (11.8) T/I
42. 12.0 L of nitrogen gas is at a pressure of 100.0 kPa and temperature of 27 °C. To what temperature must the nitrogen be cooled for its volume to become 10.6 L, still at 100.0 kPa? Give your answer in degrees Celsius. (11.8) T/I
43. You and some friends go out on a winter day to play soccer. Indoors the ball seemed fine, but outdoors it becomes soft and somewhat deflated. Use the combined gas law and the kinetic molecular theory to explain what has happened to the air in the ball. (11.1, 11.9) K/U A
44. In a laboratory experiment, a team of students collects 100 mL of hydrogen gas at 22 °C and 90.659 kPa. Determine the volume the gas sample would have at standard temperature and pressure (273 K and 101.325 kPa). Show calculations to support your answer. (11.9) T/I C

Analysis and Application

45. Some spring water smells strongly of rotten eggs due to dissolved hydrogen sulfide, H_2S. If you open a bottle of this spring water, it takes quite a few seconds to smell the hydrogen sulfide from across the room. Yet at room temperature, the average speed of gaseous hydrogen sulfide molecules is 1700 km/h. Explain why the smell of the hydrogen sulfide takes so long to reach across the room, especially when the air is still. (11.1) K/U T/I
46. As a demonstration, a chemistry teacher places a spoonful of baking soda, $NaHCO_3(s)$, into 20 mL of vinegar in a tall glass. A chemical reaction occurs between the baking soda and the ethanoic acid, $C_2H_4O_2(l)$, in the vinegar, producing a great deal of bubbling. When the bubbling has died down, the teacher tips the glass, as if pouring water, over a beaker in which a short candle is burning. Even though the teacher pours no liquid into the beaker, the candle is extinguished! (11.1) K/U T/I C

    (a) Write a balanced chemical equation for the reaction between the baking soda and the ethanoic acid.
    (b) Use the concepts of density and the properties of gases to explain why the candle was extinguished.

47. Diffusion is the spread of one substance through another due to the random motions of the entities that comprise each substance. Predict the relative rates of diffusion between two solids in contact, two liquids in contact, and two gases in contact. Refer to the kinetic molecular theory to explain your prediction. (11.1) K/U T/I

48. The average concentration of ozone in the stratosphere is roughly three ozone molecules for every one million molecules. The average concentration of all gas entities in the stratosphere is approximately 1×10^{-3} mol/L. If the total volume of the stratosphere is approximately 2×10^{22} L, what is the approximate mass of ozone in the stratosphere, in tonnes? Explain your calculations. (11.2) K/U T/I

49. A class of students attempts to verify that magnesium oxide, $MgO(s)$, is 60.3 % magnesium by mass, the theoretical value (Section 6.7). Each group of students burns a known mass of magnesium in oxygen:

    $2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$

    The students then determine the mass of the magnesium oxide and divide the mass of magnesium by the mass of magnesium oxide to get the experimental mass percentage of magnesium. The students find that their class average is bigger than 60.3 % by a significant margin. Sir William Ramsay used a reaction involving magnesium to isolate argon from atmospheric nitrogen. Using Ramsay's reaction as a clue, offer an explanation that could account for the class result being too high. Include a balanced chemical equation and relevant calculations in your answer. (11.3) K/U T/I C

50. In hot weather, city authorities might recommend that citizens fill their cars with gasoline only in the early morning and late evening hours to help air quality. Explain the reasoning behind this suggestion. (11.4) K/U T/I A

51. You are watching the news early in the morning on a summer day and the weather forecaster says that the local AQHI is 3, but is expected to climb to 7 by late afternoon. Given this information, how would you encourage your family members to plan their daily activities? Explain your thinking. (11.4) K/U T/I A

52. A neighbour is at home on a very cold winter day when the power fails, including power to the furnace controls. To heat the house, the neighbour brings a propane barbecue grill indoors and lights it. Is this a wise plan? Why or why not? (11.5) K/U A

53. Suppose your family is considering buying a home, preferably a previously occupied residence. To make sure that indoor air pollution will not be a problem in a prospective home, prepare a list of at least six questions to consider as you look around homes for sale. Include a detailed rationale for each item on the list. (11.5) K/U T/I A

54. The chemical equations below represent two of many possible reactions that occur when gasoline is burned in an idling car engine. (11.4, 11.6) K/U A

    $$2\ C_5H_{12}(l) + 7\ O_2(g) \rightarrow 2\ CO(g) + 2\ CO_2(g) + 8\ H_2O(g) + 2\ C(s) + 2\ C_2H_4(g)$$

    $$N_2(g) + O_2(g) \rightarrow 2\ NO(g)$$

    (a) Identify any air pollutants that are produced by these reactions, and briefly describe the problems that each pollutant causes.

    (b) Why might an idling car be more polluting than a car driving down the road?

55. A patient comes into an emergency room complaining of a bad headache. Her lips and cheeks are exceptionally red. Why would the doctor ask if she has been sitting in an idling car for an extended time or if she lives in a home with an old or poorly adjusted furnace? (11.5, 11.6) K/U T/I A

56. At sea level, the pressure of the air is equivalent to about 1 kg resting on every square centimetre of your body. Yet people at sea level are not crushed by all of this weight. Explain this apparent contradiction. (11.1, 11.7) K/U T/I

57. In an action movie, a hole is blown in the wall of an airplane's passenger cabin at high altitude. Passengers are sucked out through the hole. Are they really "sucked" out, though? Explain what is actually going on. (11.7) T/I

58. Encana Corporation, headquartered in Calgary, Alberta, is one of North America's largest producers of natural gas. The gas in a typical long-distance pipeline has a pressure of 60.0 atm and a cross-sectional area of 6600 cm^2. For each part below, show your work fully. (11.7) K/U T/I C

    (a) Convert the pressure in the pipeline to kilopascals.

    (b) How does this value compare to standard atmospheric pressure?

    (c) Suggest at least two hazards that might be associated with transporting gas in a pipeline such as this.

59. The home stadium of the Colorado Rockies baseball team, in Denver, Colorado, is at a high elevation compared with most baseball cities. When the stadium was built, some people predicted that batters would score an unusual number of home runs. Explain this prediction. (11.7) K/U T/I A

60. Why would it be extremely dangerous to put dry ice, $CO_2(s)$, into a sealed container like a plastic pop bottle? (11.1, 11.7) K/U T/I

61. If the entities of a gas had no attractive forces between them and if they were infinitely small in diameter, the gas sample could contract to a zero volume if its temperature were reduced to absolute zero. However, the entities of a gas do have some attraction for each other, and they do individually take up space. Assuming constant pressure, sketch a graph of volume versus temperature for a sample of gas that does have attractive forces between its entities and whose entities do individually take up space. Explain your prediction using the concepts of the kinetic molecular theory. (11.1, 11.8) K/U T/I C

62. A latex rubber balloon inflated with helium will soar up into the atmosphere if it is released. (11.7, 11.9) K/U
    (a) How will the balloon's volume change as it rises? Why?
    (b) What might be the ultimate fate of the balloon?

63. The pressure inside a welding bottle of ethyne (acetylene) gas is 13 800 kPa at 26 °C. The valve on the tank can withstand 68 900 kPa before the bottle "blows its top" due to valve failure. If the bottle is heated to 1400 °C in a warehouse fire, is the valve in danger of failing, sending the bottle off like a violent, fiery torpedo? Show calculations to support your answer. (11.9) T/I C

64. A gas sample is sealed in a cylinder by a freely moveable piston (**Figure 1**). Predict what will happen if one of the masses is removed from the piston while the temperature is held constant. Use the kinetic molecular theory to explain your prediction. (11.1, 11.9) K/U T/I

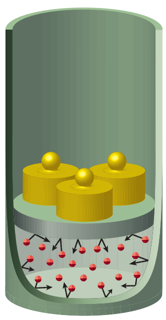


Figure 1

65. An engineer is designing a new tank to hold hydrogen gas for use in an experimental vehicle's fuel system. The tank is to be made out of a special metal alloy that is much less dense than steel. The engineer needs to know whether or not the alloy will be strong enough to hold a sufficient quantity of hydrogen under pressure. The specifications call for 1.20×10^3 L of hydrogen at normal atmospheric pressure, 101 kPa to be compressed into a 40.0 L tank made of the light alloy. The alloy can readily withstand pressures up to 5050 kPa. Assuming constant temperature, calculate the pressure that will exist in a filled tank. Predict whether the alloy will be able to hold the hydrogen safely. Support your prediction with calculations. (11.9) T/I C

Evaluation

66. Suppose that you live in a city and are interested in promoting carbon sequestration through the planting of vegetation. You will need a convincing, comprehensive argument to get citizens to participate in planting efforts. What would you have to find through research to construct your argument? Explain your thinking. (11.2) T/I A

67. Paints usually contain a pigment and a binder, both of which are suspended or dissolved in a mixture of solvents. When the paint dries, the solvents evaporate and the binder forms a pigmented film that covers a surface. Some paints, such as latex formulations, are mostly water-based with smaller amounts of organic solvents. Others are oil-based and are formulated with several organic solvents. Taking into account the environmental impact from the application of a paint product as well as the effectiveness of the product, what would you have to find out to evaluate which paint is best for a particular project? (11.4) K/U T/I A

68. A relative has read that sealing all cracks between window frames and walls, putting plastic covers over unused electrical outlets, and installing weatherstripping and draft guards on all exterior doors can prevent cold air from entering a house in winter. He thinks that this is a good idea, but he asks for your opinion. (11.5) K/U T/I A
    (a) What pros and cons would you describe to your relative about making his home "airtight"?
    (b) On balance, what course of action would you recommend to your relative? Explain your answer.

69. There are several companies that give hot-air balloon rides in Ontario. An engineer who works for a ballooning company is designing a new hot-air balloon. The balloon must lift four people, as well as the envelope, the burner, and the basket, all of which will have a maximum total mass of 760 kg. The mass that a balloon can lift is the mass of the air it displaces minus the mass of the gas inside its envelope. (11.8) K/U T/I C A

(a) The engineer plans to use an envelope that inflates to a volume of 2650 m^3, meaning the inflated balloon will displace 2650 m^3 of air. She assumes environmental conditions (outside of the balloon envelope) of 20 °C and 100.0 kPa for her calculations. If the density of air at 20 °C and 100.0 kPa is 1.23 kg/m^3, what mass of air will the fully inflated balloon displace?

(b) The engineer knows the best average temperature for the hot air in the balloon is 90 °C. Suppose 2650 m^3, the volume that would fill the balloon, of 90 °C air is cooled to 20 °C at the constant environmental pressure of 100.0 kPa. What will the new volume of the air be?

(c) Find the mass of the volume of air you found in (b) using a density of 1.23 kg/m^3.

(d) Predict whether the balloon as designed will work properly. Explain your reasons. If the balloon will not work properly, suggest and explain, quantitatively, a change that the engineer might make in the design.

70. A student calculates the new pressure of a fixed amount of gas in a rigid container after a temperature change. The original pressure of the gas was 0.800 kPa and the temperature was changed from 10 °C to 50 °C. The student's problem-solving strategy is shown below. Evaluate the student's strategy. Summarize the positive aspects and negative aspects and show how the negative aspects could be improved. (11.9) K/U T/I C

Given: $P_1 = 0.800$ kPa

$T_1 = 10$ °C

$T_2 = 50$ °C

The volume and amount of gas remain constant.

Required:

final pressure, P_2

Analysis:

Use Gay-Lussac's law, $\dfrac{P_1}{T_1} = \dfrac{P_2}{T_2}$

Solution:

$$P_2 = \frac{P_1}{T_1} \times T_2$$

$$P_2 = \frac{0.800}{10} \times 50$$

$$= 4.0 \text{ kPa}$$

Statement:

When the temperature of a fixed amount of gas at 0.800 kPa and constant volume is raised from 10 °C to 50 °C, the new pressure is 4.0 kPa.

Reflect on Your Learning

71. State something you learned in this chapter that surprised you or that you found very interesting. Explain your surprise or interest in this subject. K/U

72. You learned a great deal about air pollution in this chapter. State three strategies for improving outdoor air quality that you will implement in your own daily life or that you will encourage others to implement in their lives. Explain the importance of each strategy. K/U A

73. After studying this chapter, how concerned are you about indoor air pollution in your home? Explain your answer. K/U A

74. You learned about atmospheric pressure in this chapter. State and explain an example of how this knowledge applies to your daily life. K/U C A

Research

GO TO NELSON SCIENCE

75. Research and report on the industrial production and uses of argon gas. Include and explain the terms "distillation" and "cryogenic" in your report. Also include a diagram of the machinery and processes that are used to isolate argon from air. K/U C A

76. In August of 1986, Lake Nyos in Cameroon suddenly became a killer, not by a flood of water, but by a flood of carbon dioxide. Some 1700 people in the area were killed. Research this disaster and what has happened in its aftermath. Report your findings and include a discussion of how the kinetic molecular theory and the properties of gases explain how the carbon dioxide spread and then dissipated. K/U A

77. Research the work of Luis W. Alvarez and Walter Alvarez to learn how their research on atmospheric particulate matter resulted in a theory about the extinction of the dinosaurs 65 million years ago. Report your findings in a format of your choice. T/I C A

78. Many drivers believe that it is bad to start a cold car engine and immediately drive away. They believe it is better to start the engine and let it idle for several minutes, especially on cold days, so that the engine can properly "warm up." Is this a myth or a fact? Find out through research. Report your findings as either a short oral presentation or in the form of a web page. T/I C A

79. Prior to the twentieth century, scientists believed that the motion of all entities completely ceased at absolute zero. The theory of quantum mechanics, which was developed in the 1920s and beyond, changed this view. Research what quantum mechanics says about the motion of entities at absolute zero. Also research whether or not absolute zero can ever be attained according to the laws of thermodynamics. Summarize your findings in a brief report. T/I C

CHAPTER

12 Gas Laws, Gas Mixtures, and Gas Reactions

KEY CONCEPTS

After completing this chapter you will be able to

- analyze the cumulative effects of human activities and technologies on air quality
- explain Avogadro's law and how his contribution to the gas laws has increased our understanding of the chemical reactions of gases
- describe the quantitative relationships that exist between the pressure, volume, temperature, and amount of substance for an ideal gas
- explain Dalton's law of partial pressures
- use stoichiometry, the ideal gas law, and Dalton's law of partial pressures to solve quantitative problems involving gases

How Can an Understanding of the Gas Laws, Gas Mixtures, and Gas Stoichiometry Help Us in Our Daily Lives?

We are surrounded by gases. We breathe them, ride on them, use them to propel our cars and planes, and, sometimes, fear their destructive power. They affect every aspect of our lives. Clearly, it is important to understand the behaviour of gases. As you learned in Chapter 11, the relationships between the volume, temperature, and pressure of gases have been so thoroughly tested that they are now accepted as laws. Applications of these laws can sometimes seem mundane and ordinary, such as the process of breathing that was explored in Section 11.9. At other times, the applications are large, dramatic, and extraordinary.

Consider the eruptions of Mt. Eyjafjallajökull in Iceland in April, 2010, shown in the photograph opposite. These eruptions sent volcanic ash high into the atmosphere and interrupted air travel worldwide for days. Magma (molten rock) deep beneath Earth's surface contains a variety of dissolved gases, including carbon dioxide and sulfur compounds. As the magma rises toward the surface, the pressure on the gases decreases and the gases come out of solution—just like bubbles in a bottle of pop when you unscrew the lid. As the magma moves closer to the surface, the gas bubbles continue to expand, leading to explosive eruptions.

In this chapter, we will continue to investigate the laws that govern gases. In particular, we will explore the behaviour of gas mixtures and the stoichiometry of chemical reactions involving gases. Finally, we will explore the applications of these laws and concepts to human activities that involve gases.

STARTING **POINTS**

Answer the following questions using your current knowledge. You will have a chance to revisit these questions later, applying concepts and skills from the chapter.

1. Explain how the gases in magma behave as they rise through Earth's crust.
2. How does the behaviour of gases play a role in natural phenomena such as geysers and underwater vents?
3. Is the analogy between a volcanic eruption and opening a bottle of pop a good one? Suggest a strength and a weakness of this analogy.
4. What would happen if you heated an "empty" test tube of air that has been lightly stoppered? Is the test tube really empty?
5. Gases: "we breathe them, ride on them, use them to propel our cars and planes, and, sometimes, fear their destructive power." Give an example of each of these situations.
6. Why are reactions that produce a lot of thermal energy dangerous if they involve gases?
7. Brainstorm a list of important gas mixtures and important chemical reactions that involve gases.

Mini Investigation

Avoiding a Messy Opening

Skills: Predicting, Planning, Performing, Observing, Analyzing, Evaluating

SKILLS HANDBOOK A1.2, A2.4

How can you safely (and tidily) open a can of pop that has been shaken and is ready to spray pop everywhere (**Figure 1**)? In this investigation, you will examine what happens inside a shaken can of pop and test methods of effectively "disarming" it.

Figure 1 What can you do to avoid this situation?

Equipment and Materials: chemical safety goggles; lab apron; 2 cans of pop; transparent plastic bottle; balloon; tape; paper towels

1. Put on your safety goggles and apron.
2. Shake one of the cans of pop. Think about what is happening inside the can. Set it aside.
3. Pour another can of pop into the bottle. Stretch a balloon over the mouth of the bottle. Use tape to ensure a tight seal.
4. Shake the bottle. Observe what happens.
5. Perform an action that might return the gases to the bottle.
6. Perform the same action on the shaken pop can.
7. Open the shaken pop can carefully.

A. What do you think happened to the dissolved carbon dioxide when you shook up the pop can? Why should you be nervous about opening it? K/U A

B. After shaking the pop in the bottle, what did you observe? Where is the gas located? Why might the location of the gas bubbles result in a messy situation? T/I

C. What action did you propose to move the gas bubbles to the top of the bottle? T/I

D. What happened in Step 7? Was your action successful? T/I

12.1 Avogadro's Law and Molar Volume

What happens when we change the quantity (number of molecules, or mass) of a gas? So far, we have considered the relationships between the pressure, volume, and temperature of a constant amount of gas. How does the amount of a gas change this relationship? The early study of gases was primarily empirical: it was based on experimental evidence and observations. Indeed, Charles' law and Boyle's law were published before Dalton's atomic theory in 1803. The kinetic molecular theory followed in 1860, neatly offering an explanation for the laws. Before any theory is accepted by the scientific community, supporting evidence must be provided. The more situations a theory can explain, the better the theory. In Chapter 11, we used experimental data to develop the individual gas laws. In this section, we will use the predictive power of the kinetic molecular theory to extend our understanding of gases.

Figure 1 Electrical energy can be used to decompose water and separately collect hydrogen and oxygen gas. The battery is connected to electrodes inside inverted test tubes.

law of combining volumes the statement that, when measured at the same temperature and pressure, volumes of gaseous reactants and products of chemical reactions are always in simple ratios of whole numbers

The Law of Combining Volumes

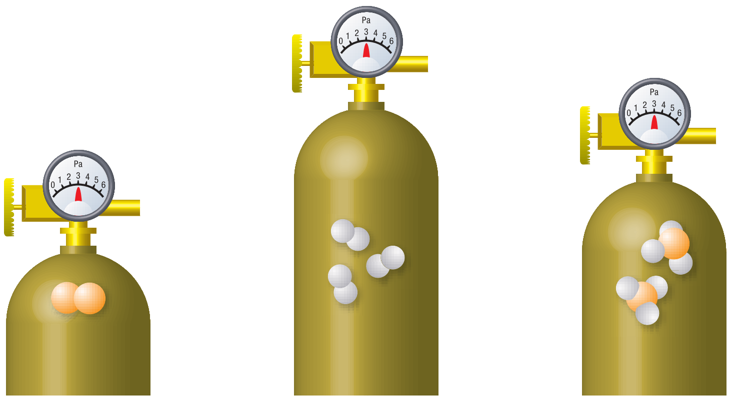
Electrical energy can be used to decompose water into its elements (**Figure 1**). Reactions occur at the two electrodes due to the electric current supplied by the battery. Oxygen and hydrogen gas bubbles are produced during the reaction:

$$2\ H_2O(l) + \text{energy} \rightarrow 2\ H_2(g) + O_2(g)$$

Each water molecule consists of 2 hydrogen atoms and 1 oxygen atom. When water decomposes, therefore, twice the volume of hydrogen as oxygen is trapped in the collection tubes.

In 1808, French scientist Joseph Gay-Lussac measured the relative volumes of gases involved in various chemical reactions. His observations showed that the volumes of gaseous reactants and products of chemical reactions are always in simple, whole-number ratios. For example, when he combined hydrogen, $H_2(g)$, and chlorine, $Cl_2(g)$, to produce hydrogen chloride, $HCl(g)$, he found that 1.0 unit volume of hydrogen added to 1.0 unit volume of chlorine always produced 2.0 units of hydrogen chloride. Gay-Lussac formulated the **law of combining volumes**, which states that gases always react to produce products in whole-number ratios.

Consider another example. Ammonia gas can be produced by combining nitrogen, $N_2(g)$, and hydrogen, $H_2(g)$ (**Figure 2**).



	$N_2(g)$	+	$3\ H_2(g)$	$\longrightarrow$	$2\ NH_3(g)$
coefficients	1		3		2
mole ratio	1	:	3	:	2
volume ratio	1	:	3	:	2

Figure 2 One volume of nitrogen reacts with 3 volumes of hydrogen, producing 2 volumes of ammonia, when all are measured at the same temperature and pressure.

Here, 1 unit volume of nitrogen reacts with 3 unit volumes of hydrogen to produce 2 unit volumes of ammonia. While Gay-Lussac proposed the law of combining gases, he did not have an explanation for it. Remember that at the time, the kinetic molecular theory had not yet been developed. It would take another scientist, Amedeo Avogadro, to propose an explanation for the law of combining volumes.

Avogadro's Law

Consider this thought experiment. How can two samples of gas that contain particles of different mass have equal pressures at the same temperature? Imagine samples of two different gases, perhaps hydrogen and oxygen, in two identical containers. The volume, temperature, and pressure are the same for both gas samples. Temperature is a measure of the average kinetic energy of entities in a sample (Section 11.1). Using the kinetic molecular theory, we know that, at a given temperature, all gas entities have the same average kinetic energy regardless of their size or mass. Since the temperature of the two samples is the same, the average kinetic energy of the entities in the two samples is the same. Any difference in pressure exerted on the walls of the containers can only be due to the number of entities present. We have already declared that the pressures are equal, so we therefore conclude that there must be an equal number of entities in the two containers (**Figure 3**). The Italian scientist Amedeo Avogadro made similar observations.

LEARNING **TIP**

Thought Experiment
A "thought experiment" is a way to solve a problem using your imagination. In a thought experiment, instead of carrying out an actual experiment, you imagine the practical outcome.

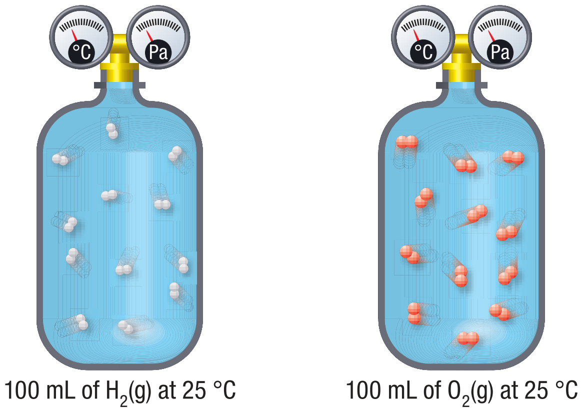


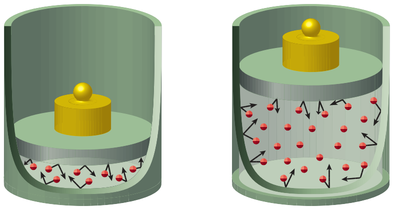
Figure 3 Under identical conditions, equal volumes of gases contain the same number of entities

Shortly after the law of combining volumes was discovered, Avogadro proposed an explanation. He was testing the laws developed by other scientists. His observations led him to combine Gay-Lussac's, Charles', and Boyle's laws to form a hypothesis that stated: Equal volumes of gases, at the same temperature and pressure, contain the same number of molecules. Avogadro's hypothesis provided a way to predict the volumes of gases, either reactants or products, involved in a chemical reaction. In this way, it explained the law of combining volumes. As Avogadro's hypothesis explains, the mole ratios provided by the balanced equation are also the ratios of volumes (Figure 2).

At the time, Avogadro's hypothesis was met with great skepticism and dismissed for over 50 years. Since then, this hypothesis has been thoroughly tested and is now accepted as a law. It is extraordinary that Avogadro was able to develop such a concept before the kinetic molecular theory was developed.

Consider an extension of our thought experiment. Take the two identical containers of gases mentioned earlier, and now connect them with a tube with a closed valve. The volumes, pressures, temperatures, and amounts are equal in the two containers. Now open the valve. What happens to the pressure, temperature, volume, and number of gas molecules? Pressure and temperature are still identical, but we can see how the volume doubles as the amount of gas doubles.

Avogadro's law the statement that the volume of a gas is directly related to the amount of gas, when the temperature and pressure of the gas remain constant; equal volumes of gases, under identical conditions, contain the same number of entities



$n = 1$ mol
$V = 1$ L

$n = 3$ mol
$V = 3$ L

Figure 4 The volume of a gas and the number of moles of gas present are directly proportional when temperature and pressure remain constant.

Avogadro's hypothesis is now known as **Avogadro's law**: the volume (V) of a gas is directly related to the amount (n) of the gas when temperature and pressure remain constant (**Figure 4**):

$$\frac{V_1}{n_1} = \frac{V_2}{n_2} \quad \text{or} \quad \frac{V}{n} = \text{constant}$$

This expression is sometimes referred to as Avogadro's hypothesis, Avogadro's theory, or even Avogadro's principle. All these names refer to the same concept.

Tutorial 1 Using Avogadro's Law

We can use Avogadro's law to calculate the volume of a gas when we know the amount of gas.

Sample Problem 1: Using Avogadro's Law to Determine Volume

A balloon with a volume of 34.5 L is filled with 3.2 mol of helium gas. To what volume will the balloon expand if another 8.0 g of helium is added? (Assume that pressure and temperature do not change.)

Given: Identify the known variables. Note what variables are being held constant.

$$\text{initial volume, } V_1 = 34.5 \text{ L}$$
$$\text{initial amount of gas, } n_1 = 3.2 \text{ mol}$$
$$\text{final amount of gas, } n_2 = n_1 + n_{\text{He added}}$$
$$\text{mass of helium added, } m_{\text{He added}} = 8.0 \text{ g}$$
$$\text{molar mass of helium added, } M_{\text{He}} = 4.00 \text{ g/mol}$$

The pressure and temperature remain constant.

Required: Identify the unknown variable.

$V_2 = ?$

Analysis: Use Avogadro's law to find the final volume of helium.

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Solution:

Step 1. Determine the amount of helium added, n_{He}, using the appropriate conversion factor derived from the molar mass of helium.

$$n_{\text{He added}} = m_{\text{He added}} \times \frac{1 \text{ mol}}{4.00 \text{ g}}$$
$$= 8.0 \cancel{\text{g}} \times \frac{1 \text{ mol}}{4.00 \cancel{\text{g}}}$$
$$n_{\text{He added}} = 2.0 \text{ mol}$$

Step 2. Determine the final total amount of helium, n_2

$$n_2 = n_1 + n_{\text{He added}}$$
$$= 3.2 \text{ mol} + 2.0 \text{ mol}$$
$$n_2 = 5.2 \text{ mol}$$

Step 3. Rearrange the Avogadro's law equation to isolate the unknown variable, substitute in the known values, and solve the equation.

$$V_2 = \frac{V_1 n_2}{n_1}$$
$$V_2 = \frac{34.5 \text{ L} \times 5.2 \cancel{\text{mol}}}{3.2 \cancel{\text{mol}}}$$
$$V_2 = 56 \text{ L}$$

Statement: The balloon will expand to 56 L when the additional helium is added. Check your answer. Does it make sense? Are the units appropriate?

Practice

SKILLS HANDBOOK A6

1. If 2.4 mol of a gas (n_1) occupies a volume of 43 L (V_1), what volume (V_2) will 4.8 mol (n_2) of the same gas occupy? Assume pressure and temperature are kept constant. T/I [ans: 86 L]
2. A sample containing 1.80 mol of argon gas has a volume of 10.00 L. What is the new volume of the gas, in litres, when each of the following changes occurs in the quantity of the gas? Assume that pressure and temperature remain constant. The changes are not cumulative. T/I
   (a) An additional 1.80 mol of argon gas is added to the container. [ans: 20.0 L]
   (b) A sample of 25.0 g of argon gas is added to the container. [ans: 13.5 L]
   (c) A hole in the container allows half of the gas to escape. [ans: 5.00 L]
3. A balloon that contains 4.80 g of carbon dioxide gas has a volume of 12.0 L. Assume that the pressure and temperature of the balloon remain constant. What is the new volume of the balloon if an additional 0.50 mol of CO_2 is added? T/I [ans: 67 L]

Molar Volume

Avogadro's law tells us that equal volumes of any gas at the same temperature and pressure contain an equal number of entities. Incorporating the mole concept, we can understand that one mole of chlorine gas, $Cl_2(g)$, and one mole of methane gas, $CH_4(g)$, under identical conditions of pressure and temperature, should occupy the same volume. The **molar volume** is the volume occupied by one mole of a gas. It is the same for all gases.

- One mole of any gas at standard temperature and pressure (STP, 0 °C and 101.325 kPa) occupies 22.4 L (**Figure 5**).
- One mole of any gas at standard ambient temperature and pressure (SATP, 25 °C and 100 kPa) occupies 24.8 L.

molar volume the volume that one mole of gas occupies at a specified temperature and pressure

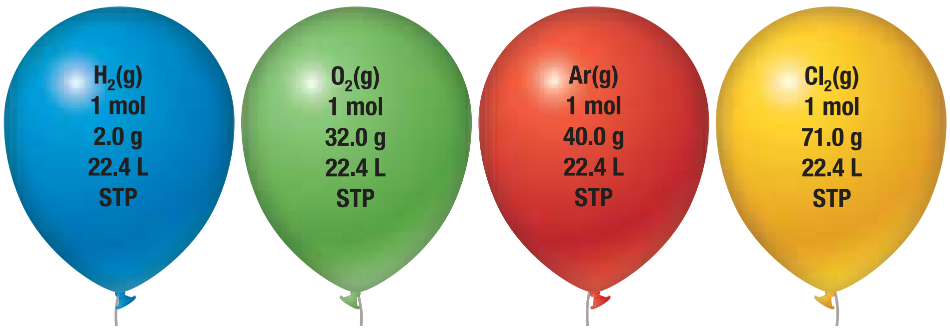
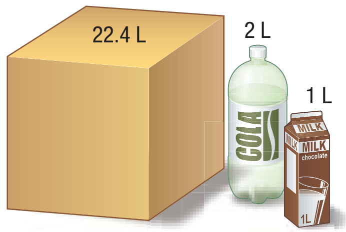


Figure 5 At STP, one mole of any gas occupies 22.4 L of volume.

LEARNING TIP

How Big Is 22.4 L?

To help picture how much space 22.4 L occupies, imagine about twenty-two 1 L milk cartons or eleven 2 L pop bottles, or a box 50 cm × 20 cm × 22 cm.



Tutorial 2 Converting among Amount, Mass, and Volume

When making conversions among the amount, mass, and volume of a gas, remember to note whether the conditions are STP or SATP. Use **Figure 6** as a reminder when converting between amount and volume using molar volume at STP.

Sample Problem 1: Converting from Amount to Volume

A party balloon has 2.50 mol of helium gas in it at STP. What is the volume of the balloon?

Given: amount of helium, n = 2.50 mol; STP conditions

Required: volume, V

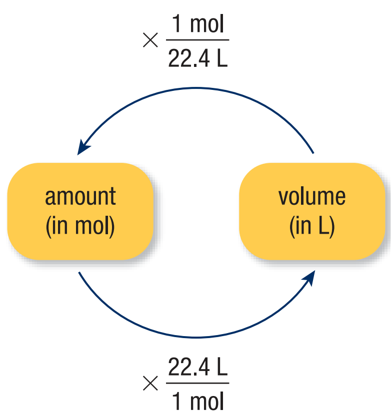


Figure 6 Converting from amount to volume of a gas using molar volume at STP

Analysis: To solve this problem, multiply the amount of helium gas in the balloon by the appropriate conversion factor derived from the molar mass ratio at STP. Since 1 mol of any gas at STP occupies 22.4 L, the following conversion factors may be considered: 22.4 L/1 mol or 1 mol/22.4 L. In this case, we use the factor, 22.4 L/1 mol as follows:

$$V = n \times \frac{22.4\ \text{L}}{1\ \text{mol}}$$

Solution: Substitute in the known values and solve the mathematical equation.

$$V = 2.50\ \cancel{\text{mol}} \times \frac{22.4\ \text{L}}{1\ \cancel{\text{mol}}}$$

$$= 56.0\ \text{L}$$

Statement: The balloon containing 2.50 mol of helium gas has a volume of 56.0 L at STP.

Sample Problem 2: Converting from Mass to Volume

A sample of helium at SATP has a mass of 32.0 g. What volume does this mass of gas occupy?

Given: $m_{He} = 32.0$ g; $M_{He} = 4.00$ g/mol

Required: volume of helium, V_{He}

Analysis: $V = n \times \frac{24.8\ \text{L}}{1\ \text{mol}}$

Solution:

Step 1. Determine the amount of helium gas in the sample by multiplying the mass of helium by an appropriate conversion factor derived from the molar mass of helium.

$$n_{He} = m_{He} \times \frac{1\ \text{mol}}{4.00\ \text{g}}$$

$$= 32.0\ \cancel{\text{g}} \times \frac{1\ \text{mol}}{4.00\ \cancel{\text{g}}}$$

$$n_{He} = 8.00\ \text{mol}$$

Step 2. Determine the volume of helium gas in the sample by multiplying the amount of helium by an appropriate conversion factor derived from the molar volume of a gas at SATP. In this case, the necessary conversion factor is 24.8 L/mol.

$$V_{He} = n \times \frac{24.8\ \text{L}}{1\ \text{mol}}$$

$$V_{He} = 8.00\ \text{mol} \times 24.8\ \text{L/mol}$$

$$V_{He} = 198\ \text{L}$$

Statement: The volume of 32.0 g of helium gas is 198 L at STP.

Practice

SKILLS HANDBOOK A6

1. Determine the volume occupied by the following amounts of nitrous oxide, $N_2O(g)$, at STP: T/I
   (a) 1.0 mol [ans: 22 L]
   (b) 2.0 mol [ans: 45 L]
   (c) 4.5 mol [ans: 1.0×10^2 L]
2. A container of oxygen gas has a volume of 145.6 L. If the pressure of the gas is 101.3 kPa and the temperature is 0 °C, determine the amount of oxygen gas in the container. T/I [ans: 6.500 mol]
3. Determine the mass of hydrogen gas collected in a container if the gas occupies 44.8 L at STP. T/I [ans: 4.04 g]

12.1 Summary

- Gay-Lussac's law of combining volumes states that volumes of gaseous reactants and products of chemical reactions, when measured at the same temperature and pressure, are always in simple ratios of whole numbers.
- Avogadro's law states that the volume of a gas is directly proportional to the amount of the gas when the temperature and pressure of the gas remain constant:

$$\frac{V_1}{n_1} = \frac{V_2}{n_2} \quad \text{or} \quad \frac{V}{n} = \text{constant}$$

- The molar volume of a gas is the volume that one mole of a gas occupies at a specific temperature and pressure. The molar volume of any gas at STP is 22.4 L/mol; the molar volume of any gas at SATP is 24.8 L/mol.

12.1 Questions

1. If you triple the amount of a gas in a balloon, what happens to the volume of the balloon? (Assume that temperature and pressure remain constant.) T/I
2. Determine the volume of 1.5 mol of butane gas if a 2.5 mol sample of butane has a volume of 38.5 L. Assume temperature and pressure are kept constant. T/I
3. Use the molar volume of a gas at SATP to determine the following values at SATP: T/I
   (a) the amount of nitrogen in 44.8 L of pure gas
   (b) the volume (in litres) of 4.8 mol of propane gas, $C_3H_8(g)$
   (c) the mass of carbon dioxide in 34.6 L of carbon dioxide gas, $CO_2(g)$
   (d) the volume (in mL) of 1250 g of methane, $CH_4(g)$
   (e) the amount of oxygen in 36.5 L of O_2 gas
4. Hydrogen gas is produced when we react magnesium metal with hydrochloric acid:

   $Mg(s) + 2\,HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$ T/I

   (a) If 4.50 g of hydrogen gas is collected at STP, what volume of hydrogen does this represent?
   (b) If 0.52 mol of magnesium completely reacted, what amount of hydrogen gas (in moles) would be produced at STP? How many grams of hydrogen is this?
5. Suppose you are pumping air into a deflated basketball, so that it is at a suitable pressure to play a game. Is this a good example of Avogadro's law? Explain your answer. A
6. When gases are sold they are usually compressed to high pressures. For example, when you buy propane for your barbecue, your tank is placed on a scale while it is filled under high pressure (**Figure 7**). This lets the service person known when the tank is "full." Discuss why gases, such as barbecue propane gas, are sold by mass and not by volume. T/I A

Figure 7 Propane is sold by mass.

7. John McLennan (1867–1935), a University of Toronto research scientist and inventor, devised a unique extraction process to obtain helium from Alberta natural gas. The process brought down the price of helium from \$20 000/m^3 to \$3/m^3. Research McLennan's work. T/I A

GO TO NELSON SCIENCE

12.2 Ideal Gases and the Ideal Gas Law

Ideally, we always arrive on time for our appointments, with everything we need in good order. Ideally, we always prepare for upcoming challenges. Ideally, we always treat one another with respect and our classroom or workplace is a harmonious, exciting, and wonderful place to be. How does the ideal compare with reality?

An Ideal Gas

Our ideas about how gases behave have been based, so far, on the gas laws. We have used the main points of the kinetic molecular theory to explain these ideas. All laws and calculations have been based on the assumption that the gases behave "ideally." Do you think this is actually the case? Just what is an "ideal gas"?

An ideal gas has the following properties:

ideal gas a hypothetical gas composed of particles that have no size, travel in straight lines, and have no attraction to each other (no intermolecular forces)

- All entities of an **ideal gas** have high translational energy, moving randomly in all directions in straight lines.
- When ideal gas entities collide with each other or with the container walls, the collisions are perfectly elastic. (There is no loss of kinetic energy.)
- The volume of an ideal gas entity is insignificant (zero) compared to the volume of the container.
- There are no attractive or repulsive forces between ideal gas entities.
- Ideal gases do not condense into liquids when cooled.

There is actually no such thing as an ideal gas. The ideal is an imaginary standard to which the behaviour of a known gas is compared. This imaginary gas has molecules with no volume and no mutual attraction to one another. It is simply a convenient approximation—or model—that works very well as we try to predict the behaviour of gases. At ordinary conditions, most gases obey the gas laws fairly well and their behaviour resembles that of an ideal gas.

We will continue—for now—with the assumption that gases behave ideally. Assuming this, we can put all our observations and gas laws together into a single equation known as the ideal gas law.

Putting It All Together: The Ideal Gas Law

In order to develop the ideal gas law, we will need to consider Charles' law, Avogadro's law, and Boyle's law as mathematical relationships:

Charles' law: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ (at constant n and P)

$$\frac{V}{T} = k \text{ or } V = kT$$

Boyle's law: $P_1V_1 = P_2V_2$ (at constant n and T)

$$PV = a \text{ or } V = \frac{a}{P}$$

Avogadro's law: $\frac{V_1}{n_1} = \frac{V_2}{n_2}$ (at constant P and T)

$$\frac{V}{n} = b \text{ or } V = bn$$

(Note that we do not need to consider Gay-Lussac's law, as it does not introduce any new variables.)

These three relationships show that the volume of a gas depends on the temperature, pressure, and amount of gas present. The letters k, a, and b are constants. We can now combine these relationships into one equation, as follows:

$V = R \times \frac{Tn}{P}$, where R is a single constant incorporating k, a, and b.

UNIT TASK **BOOKMARK**

As you work on your Unit Task on page 616, think about which gas laws researchers must consider when developing carbon capture and storage technologies.

The constant that turns all of the individual gas laws into a single equation is known as the **universal gas constant**. It is represented in the equation as *R*.

When pressure is measured in kilopascals (kPa), volume in litres (L), and the amount of a gas mol in moles (mol), the value of *R* is $8.314\ \text{kPa} \cdot \text{L} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$.

If we rearrange the above equation, we arrive at the most common form of the **ideal gas law** (**Figure 1**).

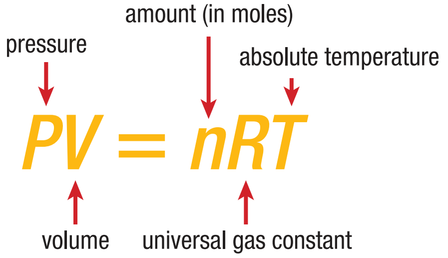


Figure 1 The ideal gas law

The ideal gas law states that the product of pressure and volume of a gas is equal to the product of the amount of gas, the universal gas constant, and the absolute temperature.

Ideal Gas Law

$$PV = nRT$$

The product of pressure and volume is equal to the product of the amount, the universal gas constant, and the absolute temperature.

Using the Ideal Gas Law

We can use the ideal gas law to find an unknown variable if the three other variables are known. While the other gas laws all compare two sets of variables, the ideal gas law works for one set of data.

Tutorial 1 Using the Ideal Gas Law

As you work through this tutorial, you will see that we can extend the use of the ideal gas law equation to include and/or determine other variables. For example, we can adapt the ideal gas law to consider the molar mass of a gas or its density. You already know that the amount of a substance (*n*) is calculated from its mass (*m*) and its molar mass (*M*):

$$n = \frac{m}{M}$$

You also know that density (*d*) is calculated by dividing its mass (*M*) by its volume (*V*):

$$d = \frac{m}{V}$$

When solving problems using the ideal gas law, make sure that your variables are in the correct units. For example, if $R = 8.314\ \text{kPa} \cdot \text{L} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$, then *P* must be in kPa, *V* must be in L, *n* must be in mol, and *T* must be in K. When you have solved the problem, check your answer. Does it make sense? Are the units appropriate?

Sample Problem 1: Determining Amount Using the Ideal Gas Law

The Goodyear blimp has a volume of 2.5×10^7 L and usually operates with the gas at a temperature of 12 °C and a pressure of 112 kPa (**Figure 2**). What amount of helium does the blimp contain? What is the mass of this amount of helium?

universal gas constant (*R*) the constant in the ideal gas law equation that relates the pressure, volume, amount, and temperature of an ideal gas

ideal gas law the statement that the product of the pressure and volume of a gas is directly proportional to the amount and the absolute temperature of the gas; $pV = nRT$

LEARNING TIP

***R* and the Ideal Gas Law**
Look closely at the units of *R* and you will see that they represent the variables in the ideal gas law: "kPa" for pressure, "L" for volume, and so on.

Figure 2 Blimps generate lift by using gases that are lighter than air. Although hydrogen is lighter than helium, it is no longer used because it is flammable.

Given: $V = 2.5 \times 10^7$ L

$t = 12$ °C

$P = 112$ kPa

$R = 8.314$ kPa · L · mol^{-1} · K^{-1}

$M_{He} = 4.0$ g/mol

Required: amount of helium, n_{He}; mass of helium, m_{He}

Analysis: Use the ideal gas law equation to calculate the amount of gas in the blimp.

$PV = nRT$

Following this, convert the amount to mass.

Solution:

Step 1. Convert the temperature value(s) to kelvins.

$$\begin{aligned} T &= 12 + 273 \\ &= 285 \text{ K} \end{aligned}$$

Step 2. Rearrange the ideal gas law equation to find the amount of gas.

$$n = \frac{PV}{RT}$$

Step 3. Substitute known values into the equation and solve.

$$\begin{aligned} n &= \frac{112 \cancel{\text{kPa}} \times 2.5 \times 10^7 \cancel{\text{L}}}{8.314 \cancel{\text{kPa}} \cdot \cancel{\text{L}} \cdot \text{mol}^{-1} \cdot \cancel{\text{K}^{-1}} \times 285 \cancel{\text{K}}} \\ &= 1.18 \times 10^6 \text{ mol} \end{aligned}$$

Step 4. Convert amount into mass using the appropriate conversion factor involving molar mass.

$$\begin{aligned} m_{He} &= n_{He} \times \frac{4.0 \text{ g}}{1 \text{ mol}} \\ &= 1.18 \times 10^6 \cancel{\text{mol}} \times \frac{4.0 \text{ g}}{1 \cancel{\text{mol}}} \\ m_{He} &= 4.7 \times 10^6 \text{ g} \end{aligned}$$

Statement: The blimp contains 1.2×10^6 mol of helium, with a mass of 4.7×10^6 g.

Sample Problem 2: Determining Volume Using the Ideal Gas Law

Calculate the volume of 32.4 g of nitrogen gas, $N_2(g)$, in a container at 25 °C and 96.4 kPa.

Given: $t = 25$ °C

$P = 96.4$ kPa

$R = 8.314$ kPa · L · mol^{-1} · K^{-1}

$m_{N_2} = 32.4$ g

$M_{N_2} = 28.02$ g/mol

Required: volume of nitrogen gas, V

Analysis: Use the ideal gas law.

$PV = nRT$

Solution:

Step 1. Convert the temperature value(s) to kelvins.

$$\begin{aligned} T &= 25 + 273 \\ &= 298 \text{ K} \end{aligned}$$

Step 2. Convert mass into amount using the appropriate conversion factor involving molar mass. Remember that we are considering nitrogen gas, which consists of diatomic molecules.

$$n_{N_2} = m_{N_2} \times \frac{1 \text{ mol}}{28.02 \text{ g}}$$

$$= 32.4 \cancel{\text{g}} \times \frac{1 \text{ mol}}{28.02 \cancel{\text{g}}}$$

$$n_{N_2} = 1.156 \text{ mol}$$

Step 3. Rearrange the ideal gas law equation to isolate the required variable.

$$V = \frac{nRT}{P}$$

Step 4. Solve the equation (including units).

$$V = \frac{1.156 \cancel{\text{mol}} \times 8.314 \cancel{\text{kPa}} \cdot \text{L} \cdot \cancel{\text{mol}^{-1}} \cancel{\text{K}^{-1}} \times 298 \cancel{\text{K}}}{96.4 \cancel{\text{kPa}}}$$

$$= 29.7 \text{ L}$$

Statement: The container has a volume of 29.7 L.

Sample Problem 3: Determining Density Using the Ideal Gas Law

Determine the density of 1.00 mol of pure carbon dioxide gas at STP.

Given: $T = 273$ K

$P = 101.3$ kPa

$R = 8.314 \text{ kPa} \cdot \text{L} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$

$M_{CO_2} = 44.01$ g/mol

$n_{CO_2} = 1.00$ mol

Required: density of carbon dioxide gas, d_{CO_2}

To find the density, you first have to determine the mass, m, and the volume, V.

Analysis: First use the ideal gas law to determine the volume.

$$PV = nRT$$

Next, convert the given amount of gas to mass.

Finally, use the expression $d = m/V$ to calculate the density of the gas.

Solution:

Step 1. Rearrange the ideal gas law equation to isolate the unknown variable, V.

$$V = \frac{nRT}{P}$$

Step 2. Substitute known values into the equation and solve.

$$V = \frac{1.00 \cancel{\text{mol}} \times 8.314 \cancel{\text{kPa}} \cdot \text{L} \cdot \cancel{\text{mol}^{-1}} \cdot \cancel{\text{K}^{-1}} \times 273 \cancel{\text{K}}}{101.3 \cancel{\text{kPa}}}$$

$$= 22.4 \text{ L}$$

You may recall seeing this number before!

Step 3. Convert amount of CO_2 into mass of CO_2 using the appropriate conversion factor involving molar mass.

$$n_{CO_2} = \frac{m_{CO_2}}{M_{CO_2}}$$

$$m_{CO_2} = n_{CO_2} \times \frac{44.01\text{ g}}{\text{mol}}$$

$$= 1.00\ \cancel{\text{mol}} \times \frac{44.01\text{ g}}{1\ \cancel{\text{mol}}}$$

$$m_{CO_2} = 44.01\text{ g}$$

Step 4. Use the relationship $d_{CO_2} = m/V$ to determine the density of CO_2.

$$d_{CO_2} = \frac{m}{V}$$

$$= \frac{44.01\text{ g}}{22.4\text{ L}}$$

$$d_{CO_2} = 1.96\text{ g/L}$$

Statement: The density of pure carbon dioxide gas at STP is 1.96 g/L.

Sample Problem 4: Determining Molar Mass Using the Ideal Gas Law

10.24 g of a pure gas occupies 2.10 L at 123 °C and 99.7 kPa. Calculate the molar mass of this gas.

Given: $t = 123\ °C$

$V = 2.10\text{ L}$

$P = 99.7\text{ kPa}$

$R = 8.314\ \text{kPa} \cdot \text{L} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$

$m = 10.24\text{ g}$

Required: amount of gas, n; molar mass of gas, M

Analysis: First, use the ideal gas law to find the amount of gas.

$$PV = nRT$$

Next, use the amount of gas to determine the molar mass.

Solution:

Step 1. Convert the temperature value(s) to kelvins.

$$T = 123 + 273$$

$$= 396\text{ K}$$

Step 2. Rearrange the ideal gas law equation to isolate the unknown variable, n.

$$n = \frac{PV}{RT}$$

Step 3. Substitute known values into the equation and solve.

$$n = \frac{99.7\ \cancel{\text{kPa}} \times 2.10\ \cancel{\text{L}}}{8.314\ \cancel{\text{kPa}} \cdot \cancel{\text{L}} \cdot \text{mol}^{-1} \cdot \cancel{\text{K}^{-1}} \times 396\ \cancel{\text{K}}}$$

$$= 0.06359\text{ mol}$$

Step 4. Find the molar mass of the gas by rearranging the values for mass and amount to give a value with the units g/mol.

$$M = \frac{10.24\text{ g}}{0.06359\text{ mol}}$$

$$= 161\text{ g/mol}$$

Statement: The molar mass of the gas is 161 g/mol.

Practice

SKILLS HANDBOOK A6

1. A neon gas in a sign has a volume of 42.5 L at 25 °C and 3.5 kPa. Calculate the mass of neon. T/I [ans: 1.2 g]
2. What amount of ethane gas is present in a sample that has a volume of 320 mL at 25.0 °C and 420 kPa? T/I [ans: 0.054 mol]
3. A tank contains 40 kg of propane gas, $C_3H_8(g)$, at a pressure of 170 kPa and a temperature of 25 °C. What is the volume of the tank? T/I [ans: 1.3×10^4 L]
4. A certain diatomic gas is found to have a density of 3.17 g/L at STP. Calculate the molar mass of this gas. T/I [ans: 71.0 g/mol]

Molar Volume for Ideal Gases and Real Gases

We will now revisit the concept of molar volume. We can use the ideal gas law to determine the molar volume of a gas at any temperature and pressure. To find the volume, V, at STP, we simply rearrange the equation and substitute in the known variables. Remember that the molar volume is the volume occupied by one mole of an ideal gas at STP (or SATP).

$$V = ?$$
$$t = 0\ °\text{C}$$
$$T = 0 + 273\ \text{K}$$
$$= 273\ \text{K}$$
$$P = 101.3\ \text{kPa}$$
$$n = 1.00\ \text{mol}$$
$$R = 8.314\ \text{kPa} \cdot \text{L} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$$
$$PV = nRT$$
$$V = \frac{nRT}{P}$$
$$= \frac{1.00\ \cancel{\text{mol}} \times 8.314\ \cancel{\text{kPa}} \cdot \text{L} \cdot \cancel{\text{mol}^{-1}} \cdot \cancel{\text{K}^{-1}} \times 273\ \cancel{\text{K}}}{101.3\ \cancel{\text{kPa}}}$$
$$= 22.4\ \text{L}$$

Thus, the molar volume of an ideal gas at STP is 22.4 L. This value appeared in Sample Problem 3 in the previous tutorial. This is also the value that was described in Section 12.1. This calculation shows how the molar volume can be determined mathematically.

Deviations from the Ideal Gas Law—Real Gases

If we analyze experimental molar volumes for different gases at STP, we see that they vary slightly from the ideal value of 22.4 L (**Table 1**). Such data remind us that, when we use the ideal gas equation, we are *imagining* ideal gases. When we experimentally measure the actual molar volume of a gas at STP, we are working with real gases.

Deviations from the gas laws occur because ideal gases do not really exist. Under ordinary circumstances, the differences between ideal gas behaviour and real gas behaviour are extremely small. Under high pressures and low temperature conditions, however, the behaviour of gases deviates more significantly.

As you know, the compressed butane in a lighter is in a liquid state. So is compressed propane in a barbecue tank. Clearly, real gases do condense into liquids (or solids) and gas molecules do experience intermolecular forces, and thus they do attract each other.

At normal atmospheric pressures, there are relatively few collisions between the entities and therefore the gases behave fairly ideally. At high pressures, however, real

Table 1 Actual Molar Volumes of Various Real Gases at STP

Gas	Molar Volume at STP (L/mol)
any ideal gas	22.414
helium, He	22.426
oxygen, O_2	22.397
chlorine, Cl_2	22.063
ammonia, NH_3	22.097

gas entities experience more collisions. This increases the attraction between the entities. As well, when the volume of the gas is decreased to much smaller volumes, the volume of the gas entities is significant with respect to the size of the container.

At normal temperatures, gas entities have a lot of kinetic energy and therefore travel at relatively high speeds. These high speeds make the forces of attraction insignificant when the entities collide. At low temperatures, however, real gas entities travel relatively slowly. At these reduced speeds, real gas entities experience forces of attraction when they collide.

The behaviour of gases resembles ideal gas behaviour when pressures are fairly low, volumes are relatively large, and temperatures are relatively high. Under these conditions the rapid movement of the gas entities overcomes any attractive forces between them, and the volume of the entities is small compared with the size of the container.

Despite these differences, we can use the ideal gas law to predict the behaviour and properties of gases under most circumstances. It is used extensively in laboratories and industrial applications. In fact, many scientific theories consider ideal situations that have to be modified to be applied to the real world.

CAREER LINK

To find out more about how engineers, such as pneumatics engineers, work with real gas, and adapt the ideal gas law to real situations,

GO TO NELSON SCIENCE

12.2 Summary

- Ideal gas molecules behave ideally. They do not condense to a liquid when cooled, they have no volume, and they do not attract each other.
- The ideal gas law, $PV = nRT$, is a combination of Charles' law, Boyle's law, and Avogadro's law (**Figure 3**). It is a powerful relationship that describes the behaviour of ideal gases.
- The gas constant (R) can be determined experimentally and has a value of $8.314\ \text{kPa} \cdot \text{L} \cdot \text{mol}^{-1} \cdot \text{K}^{-1}$.
- Even though there are no ideal gases in the real world, we can use the ideal gas laws to predict the behaviour of gases under most situations.

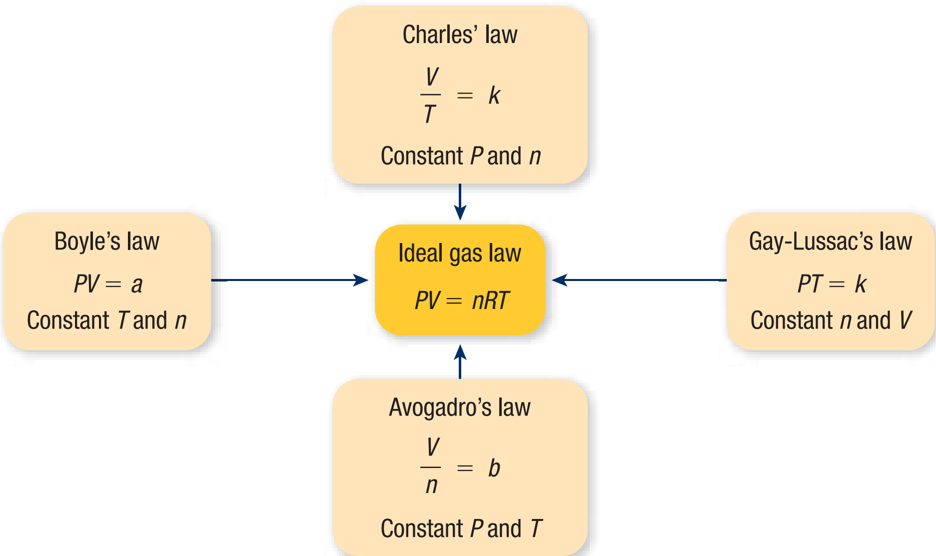


Figure 3 We can combine the individual gas laws to form the ideal gas law.

12.2 Questions

1. Describe what is meant by the term "ideal gas." K/U
2. Under what conditions is a gas closest to having the properties of an ideal gas? Explain why. K/U
3. (a) Distinguish between the molar mass of a gas and the molar volume of a gas. K/U
   (b) Does molar mass and/or molar volume change with temperature and pressure?
4. Use the ideal gas law to find the volume that 4.30 mol of oxygen gas occupies at 99.7 kPa and 35.0 °C. T/I
5. Laughing gas (dinitrogen monoxide), $N_2O(g)$, is sometimes used by dentists to keep patients relaxed during dental procedures. A 47.5 L cylinder of laughing gas has a pressure of 112.0 kPa and the temperature of the dental office is 22.0 °C. Determine the mass of laughing gas in the cylinder. T/I
6. A sample of helium gas occupies 4.52 L at a pressure of 45 kPa. If the gas has a mass of 43.5 g, calculate the temperature of the gas in degrees Celsius. T/I
7. A sample of 7.76 g of nitrogen gas, $N_2(g)$, is stored in a 3.50 L container at a temperature of 35 °C. Under what pressure is the gas being stored? T/I
8. A sample of gas has a mass of 7.60 g and occupies a volume of 2.45 L at STP. T/I
   (a) Calculate the molar mass of the gas.
   (b) Is it possible to identify the gas from these data? Explain.
9. (a) Calculate the molar volume of any gas at 112 kPa and 22 °C.
   (b) What assumptions did you make in your calculation? K/U T/I
10. Find the density of propane gas, C_3H_8,(g), at SATP. If the density of air (a mixture of gases) at SATP is approximately 1.2 g/L, will pure gaseous C_3H_8 sink or rise in air? Should propane gas detectors be placed up high or down low in a room? T/I A
11. Give three examples of how the gas laws apply to the operation and safety of automobiles. A
12. You are trying to experimentally determine the molar mass of a gas in order to identify it. What values must you be able to calculate to determine the molar mass of this mystery gas? T/I
13. (a) Under what circumstances do real gases deviate from ideal behaviour?
    (b) Research two applications where you think the differences between real gas and ideal gas behaviour would be significant. K/U A

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The Science of Cold

ABSTRACT
Understanding cold and how matter behaves at the lowest temperatures has been an epic journey of scientific discovery. The journey to achieve absolute zero experimentally has had great rivalries and it sparked the development of new technologies. Each breakthrough has been built on previous discoveries. All of this work has culminated in finding a new state of matter.

Absolute Zero

The thermometer was not readily available until the beginning of the eighteenth century. With the technology to measure temperature came the interest in the lower limit: the frontier of cold. Once the theoretical concept of absolute zero had been proposed, chemists began to research ways to achieve this temperature experimentally. The challenge would inspire two great competitions.

Dewar and Kamerlingh Onnes

At the beginning of the twentieth century James Dewar and Heike Kamerlingh Onnes both used liquified gases in their attempts to reach absolute zero. At the University of Cambridge, Dewar tried but failed to liquefy what he termed "permanent" gases: nitrogen, oxygen, and hydrogen. Discouraged, Dewar abandoned this project. It took the work of a Dutch physicist, Johannes Diderik van der Waals, to bring Dewar back to this research (**Figure 1**).

Figure 1 Heike Kamerlingh Onnes and Johannes Diderik van der Waals

Van der Waals, in 1873, was interested in the size of gas entities and the attractive forces between them. He determined that in order to liquefy permanent gases, they needed to be cooled under pressure. This discovery opened the floodgates in cold research. Oxygen, nitrogen, and hydrogen were successfully liquified over the next few years.

Dewar had competition from a young researcher in the Netherlands, Heike Kamerlingh Onnes, whose laboratory was supported by a small army of glass-blowers, instrument makers, and assistants. Kamerlingh Onnes published monthly updates of his progress. Dewar, by contrast, had a small lab and was very secretive about his progress. They both used similar methods to liquefy hydrogen: cooling one gas under pressure and using it to cool another gas in a series of steps. Each gas in turn caused the next to liquefy. The last gas, hydrogen, would finish the process at 180 times atmospheric pressure. The hydrogen was then released through a valve. The enormous pressure drop would trigger a large temperature drop. In 1898, Dewar's lab successfully used this method to liquefy hydrogen at 23 K. However, several of Dewar's assistants were injured in high-pressure explosions of equipment. Word of these explosions caused Dutch officials to close Kamerlingh Onnes' laboratory. Dewar wrote a letter of protest on his behalf, but the lab remained closed for two years.

When the Dutch lab reopened, it was apparent that the competition was not over. A recently discovered element, helium, had successfully been isolated. Liquefying helium presented a new challenge. Dewar worked in the lab next door to the discoverers of helium. Acquiring a sample of helium should have been an easy task. Unfortunately, Dewar had been critical of his colleagues and they refused to provide him with any helium. Kamerlingh Onnes also had difficulty obtaining helium. Eventually both labs acquired the gas, but an assistant in Dewar's lab inadvertently released all their stored helium into the air. In 1908, Kamerlingh Onnes was able to liquefy helium at 5 K. For his work, Kamerlingh Onnes was awarded the Nobel Prize in 1913. An infuriated Dewar moved on to other research.

The Bose–Einstein Condensate

Early in the twentieth century, Satyendra Bose, an Indian physicist, used mathematical models to propose that a new state of matter was possible at temperatures close to absolute zero. He sent his ideas to Albert Einstein in Germany. Together, the two scientists developed the idea of this new state of matter, which became known in 1925 as a Bose–Einstein condensate. Making this new state of matter became the next great challenge in the exploration of cold. The Bose–Einstein condensate is not like anything else in common experience. As atoms cool they stop behaving like individual atoms; rather, they exist in a state described as a "super atom."

At Massachusetts Institute of Technology (MIT) in the early 1990s, Dan Kleppner and Tom Greytak attempted to make a Bose–Einstein condensate. They chose hydrogen, which is light and has weak attractive forces between molecules, and used magnetic fields to create cold conditions. Unfortunately the atoms would occasionally interact in a way that increased the temperature, thwarting the researchers.

In 1995 at the University of Colorado, Eric Cornell and Carl Wieman took up the challenge to form the condensate. They used heavier atoms and a "laser trap" as a cooling system. When atoms are moving at a particular speed, they also have a certain frequency. Lasers can be set to exactly match the frequency of the atoms, which has the effect of slowing the atoms down. Wolfgang Ketterle and his research group at MIT also joined the race. Eventually all these groups were using similar magnetic fields and laser cooling.

The first Bose–Einstein condensate was made in 1995 at 170 billionths of a degree above absolute zero by Wieman and Cornell (**Figure 2**). Within weeks, Ketterle had produced a larger sample of the condensate. All three research groups shared the Nobel Prize for this accomplishment in 2001.

The Future of Cold

Research into the study of the limits of cold continues. In 2003, Deborah Jin of the University of Colorado created the first fermionic condensate: yet another state of matter. In 2003, Ketterle and his MIT group reached 0.45 nK (4.5×10^{-10} K), the coldest temperature—to date.

Each breakthrough was built on previous ideas and experiments, and has inspired new technologies. Competition and collaboration drove the progress toward absolute zero and the exploration of new states of matter.

Figure 2 A new state of matter: a Bose–Einstein condensate. This is a representation of a small cloud of super-cooled (to 180 billionths of a degree above absolute zero) rubidium atoms.

Further Reading

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Shachtman, T. (2000). *Absolute zero and the conquest of cold.* New York: Mariner Books.

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12.3 Questions

1. Dewar and Kamerlingh Onnes had different approaches to their research. Briefly describe these differences. K/U
2. In your opinion, should the Nobel Prize have been given jointly to both Kamerlingh Onnes and Dewar? Defend your position. A
3. Name two methods used to cool atoms to make the Bose–Einstein condensate. K/U
4. Briefly describe a Bose–Einstein condensate. Why is it named for two scientists? K/U
5. Why do you think scientists want to achieve a temperature of absolute zero? T/I
6. Do you think reaching absolute zero is an attainable goal? Defend your answer. A
7. Do you think competition between labs or scientists helps or hinders scientific progress? Explain your position and connect it to the information in this journal article. A
8. James Dewar invented the vacuum flask, also called the Dewar flask or Thermos flask. Research this discovery and how it works. Write a short summary of your findings. T/I C A

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12.4

Gas Mixtures and the Law of Partial Pressures

Most naturally occurring gases are mixtures. As you learned in Section 11.2, air is an example of a gaseous mixture of oxygen, nitrogen, and other gases. In this chapter's opening, you read about how volcanic eruptions are caused by the expansion of gases as they rise to the surface of Earth. Gas mixtures also play an important role in the destructive power of volcanoes. Pyroclastic flows, or *nuées ardentes* (French for glowing clouds), are hot mixtures of gases and volcanic fragments (**Figure 1**). The gases include water vapour, carbon dioxide, and sulfur dioxide with smaller quantities of other acidic compounds at 200 °C to 700 °C. These flows move close to the ground with speeds greater than 80 km/h. In 1902, Mont Pelée in Martinique experienced a pyroclastic flow that killed close to 30 000 inhabitants.

WEB LINK

To learn more about the gases that are released during a volcanic eruption,

GO TO NELSON SCIENCE

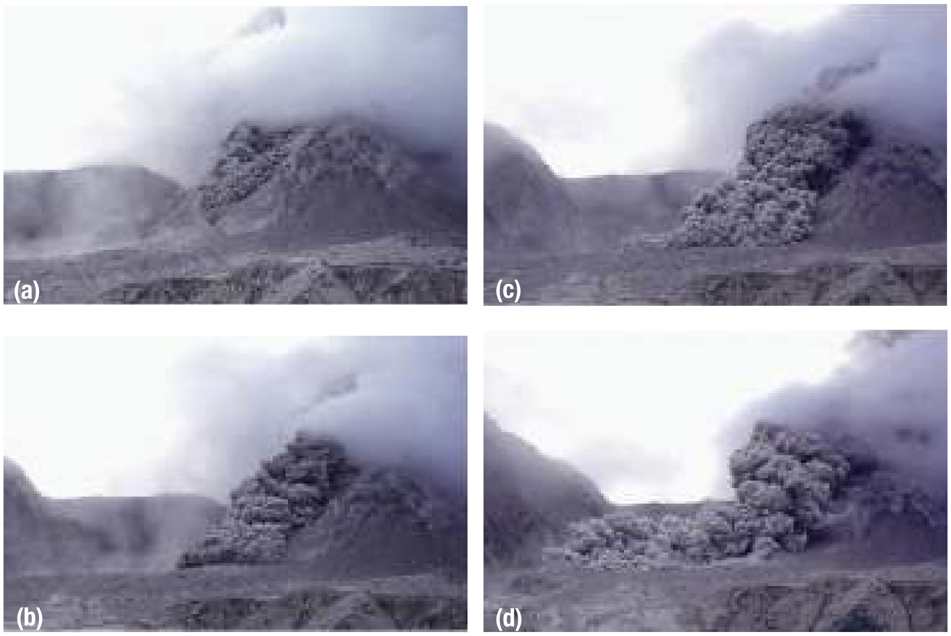

Figure 1 These images of the Montserrat volcano, taken over only a few seconds, show how fast a pyroclastic flow can move.

In our day-to-day lives we have frequent experience with gases. Human flatulence is a mixture of gases: mostly nitrogen and carbon dioxide. Methane, oxygen, and hydrogen are present in lesser amounts, along with a small concentration of sulfur-containing gases. It is gases such as hydrogen sulfide that cause the unpleasant odour. These gases are all by-products of bacterial metabolism. We all have bacteria living in our large intestines; they are necessary for digestion. Unfortunately, the bacteria produce gases that are released as flatulence.

Gas mixtures can be very useful. Anesthesiology is the branch of medicine that involves the administration of medications that cause the patient to lose awareness, so no pain is felt during a medical procedure. Many of these drugs are given to patients as gas mixtures. Anesthesiologists use anesthesia machines to ensure that accurate gas mixtures are prepared and administered.

CAREER LINK

Anesthesiologists are physicians who administer anesthetics to patients to help control pain, usually during surgery. To learn more about this profession,

GO TO NELSON SCIENCE

John Dalton was a scientist whose interests extended beyond his theory of atoms. He also researched the behaviour of gases extensively. He noticed that air at higher temperatures could "hold" greater concentrations of water vapour than air at lower temperatures. The work on water vapour led him to study the pressures of air and water vapour mixtures. He began with dry air and determined the pressure. He added water vapour, and noticed that the pressure increased. Dalton used the term **partial pressure** to refer to the pressure that was exerted by each of the gases in a mixture. He published his discovery in 1805. In the following excerpt, the term "fluid" refers to the gas state:

> I think a very important and fundamental position in the doctrine of such fluids;—namely, that the elastic or repulsive power of each particle is confined to those of its own kind; and consequently the force of such fluid, retained in a given vessel, . . . is the same in a separate as in a mixed state. . . . This principle accords with all experience, and I have no doubt will soon be perceived and acknowledged by chemists and philosophers . . .

The Philosophical Magazine, XXIII, (1805–1806), pp. 349–356.

partial pressure the pressure that a gas in a mixture would exert if it were the only gas present in the same volume and at the same temperature

WEB LINK

To read the original printing of John Dalton's essay, *Experimental Inquiry into the Proportion of the Several Gases or Elastic Fluids Constituting the Atmosphere,*

GO TO NELSON SCIENCE

The Law of Partial Pressures

Dalton's extensive studies led him to develop a law that has become known as **Dalton's law of partial pressures** (**Figure 2**).

Dalton's Law of Partial Pressures

The total pressure of a mixture of non-reacting gases is equal to the sum of the partial pressures of the individual gases.

$$P_{total} = P_1 + P_2 + P_3 + \ldots$$

Dalton's law of partial pressures the statement that the total pressure of a mixture of non-reacting gases is equal to the sum of the partial pressures of the individual gases

Figure 2 Dalton became a teacher at his local school when he was 12 years old. He was fascinated with meteorology, keeping a weather journal containing over 200 000 entries for over 50 years.

The law of partial pressures indicates that the total pressure of a mixture of gases is the sum of the pressures of each gas in the mixture. The gases must not react, and the pressure units must be consistent. Imagine a mixture of nitrogen and oxygen in a scuba diver's tank: the total pressure would be the sum of the pressure of the oxygen and the pressure of the nitrogen (**Figure 3**).

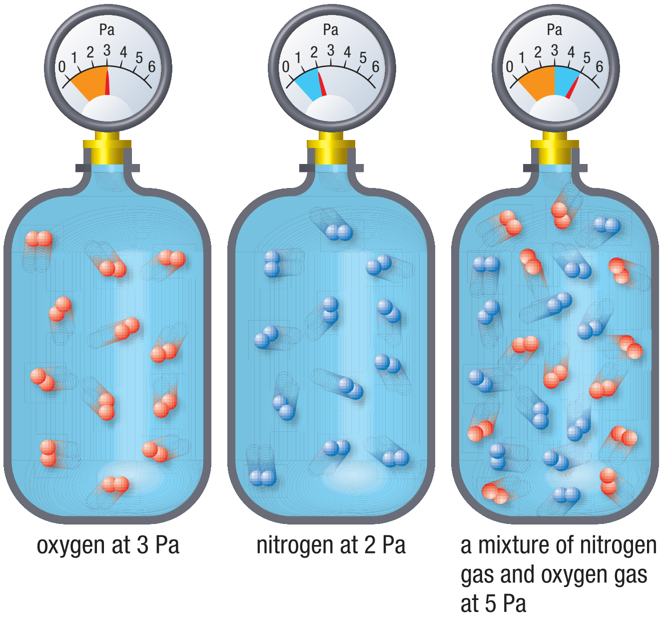

Figure 3 According to Dalton's law of partial pressure, the total pressure of the oxygen and nitrogen mixture (5 Pa) equals the sum of the pressure of each gas (3 Pa + 2 Pa).

We can explain the law of partial pressures by using the kinetic molecular theory (KMT). According to the KMT, the pressure of a gas is caused by the collisions of molecules with the walls of the container, and gas entities act independently of each other. The total pressure of a mixture of gases should equal the sum of their individual pressures because the gas particles do not interact with one another. The pressure of this mixture would be the combined pressures of the individual gas entities striking the container.

Tutorial 1 Using the Law of Partial Pressures

We can use Dalton's law of partial pressures to solve problems involving mixtures of two or more gases that do not react. Remember that the partial pressures must all be in the same unit.

Sample Problem 1: Calculating the Total Pressure

A compressed air tank for scuba diving, designed for depths of 30 m, has oxygen gas at a partial pressure of 2.8×10^3 kPa and nitrogen gas at a partial pressure of 1.1×10^4 kPa. What is the total pressure in the tank in atm and kPa?

Given: partial pressure of oxygen, $P_1 = 2.8 \times 10^3$ kPa;
partial pressure of nitrogen, $P_2 = 1.1 \times 10^4$ kPa

Required: total pressure, P_{total}

Analysis: Apply Dalton's law of partial pressures.

$$P_{\text{total}} = P_1 + P_2$$

Solution: Substitute known values into the equation and solve.

$$\begin{aligned} P_{\text{total}} &= P_1 + P_2 \\ &= 2.8 \times 10^3 \text{ kPa} + 1.1 \times 10^4 \text{ kPa} \\ &= 1.38 \times 10^4 \text{ kPa} \end{aligned}$$

Statement: The pressure in the scuba tank is 1.4×10^4 kPa.

We can also use the law of partial pressures and the percentages of a gas in a mixture to determine the pressure exerted by a single gas.

Sample Problem 2: Calculating the Pressure of One Gas in a Mixture

If the atmosphere is 21 % oxygen by volume and the atmospheric pressure is 101.3 kPa, what is the pressure of the atmospheric oxygen?

$$\begin{aligned} P_{\text{oxygen}} &= \frac{21}{100} \times 101.3 \text{ kPa} \\ &= 21 \text{ kPa} \end{aligned}$$

The pressure of oxygen in the atmosphere is 21 kPa.

Practice

SKILLS HANDBOOK A6

1. A mixture of gases is composed of nitrogen with a partial pressure of 300 kPa, oxygen with a partial pressure of 200 kPa, and carbon dioxide with a partial pressure of 400 kPa. What is the total pressure of this mixture? T/I [ans: 900 kPa]
2. The total pressure of a mixture of oxygen and water vapour is 380 kPa. If the water vapour accounts for 162 kPa of pressure, how much pressure does the remaining oxygen exert? T/I [ans: 218 kPa]
3. A container holds three gases: oxygen, carbon dioxide, and helium. The partial pressures of the oxygen and carbon dioxide are 104 kPa and 242 kPa, respectively. If the total pressure inside the container is 545 kPa, what is the partial pressure of the helium gas? T/I [ans: 199 kPa]
4. If a gas sample is $\frac{1}{3}$ neon and $\frac{2}{3}$ krypton with a total pressure of 120 kPa, what is the pressure of each gas? T/I [ans: P_{Ne} = 40 kPa, P_{Kr} = 80 kPa]

Research This

The Hazards of Scuba Diving

Skills: Researching, Communicating

SKILLS HANDBOOK A5.1

Scuba diving—diving while breathing compressed gases—is both a career and a sport (**Figure 4**). Whatever their reasons for diving, divers must follow certain safety rules. These rules were developed as a result of divers' unpleasant experiences, such as decompression sickness ("the bends") and nitrogen narcosis. Most divers breathe air (a mixture of nitrogen gas, oxygen gas, and carbon dioxide gas) from their scuba tanks. Advanced divers, however, will mix helium in with the air.

1. Research the bends and nitrogen narcosis.

A. Explain decompression sickness using your knowledge of gas laws. A

B. How can decompression sickness be prevented? Why are these precautions effective? A

C. Describe the symptoms of nitrogen narcosis and explain what causes them. A

D. How might the addition of helium reduce the dangers of advanced dives to great depths? A

Figure 4 Scuba diving is a popular activity. It is important to be well trained and informed about the possible dangers.

CAREER LINK

Professional scuba divers are extremely careful about the mixture of gases in their tanks. There are many aspects of the gas laws that affect their work. To find out more about their training and workday,

GO TO NELSON SCIENCE

GO TO NELSON SCIENCE

Collecting Gas over Water

In the chemistry lab, you have used a technique for collecting gases that involves the displacement of water from a vessel (Investigation 1.5.1). How does this technique relate to the law of partial pressures? In this technique, a gas collection bottle, completely full of water, is placed in a wide container of water. The gas is bubbled into the gas collection bottle. As the gas bubbles rise in the bottle, the water is forced out (**Figure 5**).

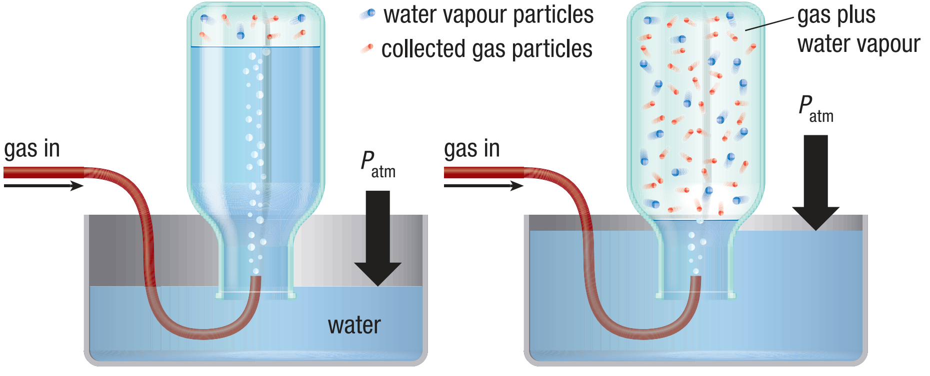

Figure 5 Gas collection by the downward displacement of water actually collects a gaseous mixture of the desired gas and water vapour.

A gas collected over a container of water is actually a mixture of the desired gas and water vapour. If the mixture is warm, it will contain a greater percentage of water than if it is cold. In other words, the partial pressure of water in the mixture depends in part on temperature. The partial pressure of water vapour in a gas mixture is

Table 1 Vapour Pressure of Water at Various Temperatures

Temperature (°C)	Vapour pressure (kPa)
17.0	1.94
18.0	2.06
19.0	2.20
20.0	2.34
21.0	2.49
22.0	2.64
23.0	2.81
24.0	2.98
25.0	3.17
26.0	3.36
27.0	3.57
28.0	3.78
29.0	4.01
30.0	4.24

Investigation 12.4.1

Determining the Molar Mass of Butane (p. 605)

In this investigation you will capture a sample of butane over water, and determine its mass and volume. From these experimental data you will determine the molar mass of butane.

known as the vapour pressure of water. Scientists have experimentally determined the vapour pressure of water at various temperatures (**Table 1**). It is important to state that we assume that the solubility of the collected gas in water is negligible. All liquids exist with a certain quantity of water vapour above their surfaces as a result of evaporation, therefore all liquids have an associated vapour pressure at a given temperature.

Dalton's law of partial pressures tells us that the total pressure of the mixture of the gases is equal to the partial pressure of the collected gas plus the partial pressure of the water vapour. In this case, the total pressure is also equal to the atmospheric pressure, which can easily be measured with a barometer.

$$P_{\text{total}} = P_{\text{atm}} = P_{\text{gas collected}} + P_{\text{water vapour}}$$

We can find the pressure of the pure gas collected by subtracting the partial pressure of water (at the appropriate temperature) from the total pressure.

$$P_{\text{gas collected}} = P_{\text{atm}} - P_{\text{water vapour}}$$

Tutorial 2 Analyzing Gas Collection Data

Whenever you collect gases over water you should consider the partial pressure of water vapour. For this, you will need to measure the temperature at which the investigation is taking place. You can then look up the partial pressure of water at this temperature in Table 1.

Sample Problem 1: Applying Dalton's Law to Gas Collection Data

Hydrogen gas, H_2, is collected using downward water displacement. If the atmospheric pressure is 100.50 kPa and the temperature is 20.0 °C, determine the pressure of the hydrogen gas.

Given: atmospheric pressure, $P_{\text{atm}} = 100.5$ kPa

temperature, $t = 20.0$ °C

vapour pressure of water at 20.0 °C (Table 1), $P_{H_2O} = 2.34$ kPa

Required: partial pressure of hydrogen gas, P_{H_2}

Analysis: Use the equation for Dalton's law of partial pressures

$$P_{\text{total}} = P_{\text{atm}} = P_{H_2} + P_{H_2O}$$

Solution:

$$P_{H_2} = P_{\text{atm}} - P_{H_2O}$$
$$= 100.50 \text{ kPa} - 2.34 \text{ kPa}$$
$$P_{H_2} = 98.16 \text{ kPa}$$

Statement: The pressure of the hydrogen gas is 98.16 kPa.

Practice

SKILLS HANDBOOK A6

1. 80.0 L of oxygen is collected over water at 30.0 °C. The atmospheric pressure in the room is 96.00 kPa. T/I
   (a) What is the total pressure of this mixture? [ans: 96.00 kPa]
   (b) What is the partial pressure of the oxygen gas? [ans: 91.76 kPa]
2. If 60.0 L of nitrogen is collected over water at SATP (25 °C and 100 kPa), what is the partial pressure of the nitrogen in kPa? T/I [ans: 96.8 kPa]
3. A volume of hydrogen gas is collected over water in an investigation. If the atmospheric pressure is 102.1 kPa and the pressure of the hydrogen gas is 100.1 kPa, what temperature is the water? T/I [ans: 17.5 °C]

12.4 Summary

- Gas mixtures are very common. Understanding how to work with gas mixtures is useful in many situations.
- The partial pressure of a gas in a mixture is the pressure that gas would exert if it were the only gas present.
- Dalton's law of partial pressures states that the total pressure of a mixture of non-reacting gases is equal to the sum of the partial pressures of the individual gases:

$$P_{total} = P_1 + P_2 + P_3 + \ldots$$

- Dalton's law of partial pressures has many applications and is essential for investigations that require collecting gases over water.

12.4 Questions

1. Give three examples of gas mixtures that exist in the natural world. K/U
2. When applying Dalton's law of partial pressure, are you assuming ideal or real gas behaviour? Explain briefly. T/I
3. A scuba tank holds a mixture of pressurized nitrogen, $N_2(g)$, and oxygen, $O_2(g)$. The partial pressure of the oxygen is 145 kPa and the total pressure of the mixture is 375 kPa. Determine the partial pressure of the nitrogen. T/I
4. A 550 mL sample of methane, $CH_4(g)$, is collected over water by downward displacement. The temperature is 25 °C and the atmospheric pressure is 100.7 kPa. K/U T/I
   (a) Determine the partial pressure of methane gas.
   (b) What important assumption is made when gas is collected by water displacement?
5. Under SATP, magnesium was reacted with acid and 500 mL of hydrogen gas was collected over water. T/I
   (a) What is the vapour pressure of water?
   (b) What is the pressure of the hydrogen gas?
   (c) What amount of hydrogen was produced?
6. Hydrogen gas is generated by the reaction of magnesium and ethanoic acid (in vinegar), and is collected by the displacement of water. If the pressure of the hydrogen is 101.35 kPa and the temperature is 18 °C, what is the total pressure of the gas mixture? T/I
7. Some athletes feel that altitude training or using low-pressure tents improves their performance. Others feel that inhaling oxygen-enhanced air is helpful. Research and compare these two approaches. Is it possible that they are both beneficial? Prepare a brief report, in any format, entitled "Oxygen: More or Less?" T/I C A
8. Oxygen bars are trendy establishments that sell oxygen-enhanced air (**Figure 6**). Customers typically pay by the minute for a supply of air. Oxygen bars operate in urban centres around the world. Research the history and operation of oxygen bars and the potential benefits or drawbacks of frequenting such a place. Remember to use reputable, unbiased sources of information. T/I A

Figure 6 Recreational oxygen use in an oxygen bar

GO TO NELSON SCIENCE

12.5 Reactions of Gases and Gas Stoichiometry

Figure 1 The combustion reaction in a motorbike's engine produces a large volume of gases: the exhaust.

The roar of a motorcycle engine turns heads as it picks up speed and races past you. (**Figure 1**). What causes that sound and results in that impressive speed? A gas reaction!

Gases are involved in chemical reactions all around us, from medical applications to the gases used to heat our homes and power our vehicles. Gases interact chemically with each other and with other forms of matter. These interactions can be represented as balanced chemical reactions. The chemical reaction in the motorcycle engine involved gasoline vapour and oxygen, and produced mostly carbon dioxide and water. In this section we will examine chemical reactions involving gases and the volumes of those gases.

Gas Stoichiometry

You learned a number of key concepts about stoichiometric relationships in chemical reactions in Unit 3. You know that chemical reactions involve mole ratios of chemicals. When working with chemical reactions where some of the reactants and/or products are gases, it is more likely that you will be measuring volumes rather than masses. We therefore use slightly different techniques to solve these stoichiometry problems.

Recall Gay-Lussac's law of combining volumes (Section 12.1). This law explains that, when gases react, the volumes of the reactants and products react in whole-number ratios if the temperatures and pressures are constant. This is useful for determining unknown volumes when working on gas stoichiometry problems.

WEB **LINK**

To learn more about how catalytic converters convert carbon monoxide into carbon dioxide,

GO TO NELSON SCIENCE

Figure 2 A catalytic converter is part of the exhaust system in an automobile.

Tutorial 1 Volume-to-Volume Stoichiometry

When solving stoichiometry problems you often know the volumes of gaseous reactants and/or products and are asked to determine other volumes in the reaction. (This assumes that the temperature and pressure remain constant throughout the problem.) For this type of problem you will need to start by writing the balanced chemical equation. You will find it helpful to list any known quantities, with the appropriate units, underneath the equation. Make sure that similar variables have identical units. You can then use a conversion factor derived from the mole ratio of the two gases to calculate the required volume.

Sample Problem 1: Using a Volume Ratio to Determine a Volume

A catalytic converter in the exhaust system of a car uses oxygen (from the air) and a catalyst to convert carbon monoxide to carbon dioxide (**Figure 2**). If the temperature and pressure remain the same, what volume of oxygen is required to react with 65.0 L of carbon monoxide produced during a road trip?

Given: volume of carbon monoxide, $V_{CO} = 65.0$ L

Required: volume of oxygen, V_{O_2}

Solution:

Step 1. Write the balanced chemical equation for the reaction, listing the given value(s) and required values below.

$$2\,CO(g) + O_2(g) \rightarrow 2\,CO_2(g)$$

65.0 L $\quad V_{O_2}$

Step 2. Convert the volume of the given substance, V_{CO}, to volume of the required substance, V_{O_2}. To do this, multiply by a conversion factor derived from the mole ratio of the given substance to the required substance. Since we want an answer involving oxygen, the conversion factor is $\frac{1\ mol_{O_2}}{2\ mol_{CO}}$.

$$V_{O_2} = V_{CO} \times \frac{1\ \text{mol}_{O_2}}{2\ \text{mol}_{CO}}$$

$$= 65.0\ \text{L} \times \frac{1\ \cancel{\text{mol}_{O_2}}}{2\ \cancel{\text{mol}_{CO}}}$$

$$V_{O_2} = 32.5\ \text{L}$$

Statement: The volume of oxygen required is 32.5 L.

Practice

1. What volume of oxygen will be required for the complete combustion of 54 L of hydrogen if both gases are measured at the same temperature and pressure?

   $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$ T/I [ans: 27 L]

2. What volume of hydrogen gas, $H_2(g)$, is required to react with nitrogen gas, $N_2(g)$, to produce 34.5 L of ammonia, $NH_3(g)$? Assume that the temperature and pressure remain constant throughout the reaction. T/I [ans: 51.8 L]

3. Gunpowder is a mixture of potassium nitrate (commonly known as saltpetre), $KNO_3(s)$, charcoal, C(s), and sulfur, $S_8(s)$. When heated or struck by a sharp blow, the potassium nitrate decomposes to produce oxygen, which reacts rapidly with the charcoal and sulfur. The decomposition of saltpetre is shown by the following balanced chemical equation:

   $4\ KNO_3(s) \rightarrow 2\ K_2O(s) + 2\ N_2(g) + 5\ O_2(g)$

   Both gases are measured at the same temperature and pressure. What volume of oxygen is produced along with 15.0 L of nitrogen? T/I [ans: 37.5 L]

More commonly, temperature and pressure vary during a chemical reaction. In this case, we must use the ideal gas law equation ($PV = nRT$) to calculate the volume of a gas given the amount determined by stoichiometric techniques.

Airbags are a dramatic example of how a gas can save human lives (**Figure 3**). During a collision, sensors activate a chemical reaction that produces a volume of nitrogen gas that immediately inflates the airbag. After cushioning the driver, the airbag quickly deflates as nitrogen molecules escape through the permeable cover.

Figure 3 A safety airbag uses gas stoichiometry to save lives.

The nitrogen gas is produced from a series of chemical reactions. The gas generator contains an electrical igniter and a precise mixture of three compounds: sodium azide, NaN_3; iron(III) oxide, Fe_2O_3; and silicon dioxide, SiO_2. This mixture is placed in a porous folded pouch. When ignited, the sodium azide decomposes very quickly to produce sodium metal and nitrogen gas:

$$2\ NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g)$$

Almost instantly, the sodium metal reacts with the iron(III) oxide to produce solid sodium oxide and iron:

$$6\ Na(s) + Fe_2O_3(s) \rightarrow 3\ Na_2O(s) + 2\ Fe(s)$$

Finally, the thermal energy released by these reactions melts the solid products and the silicon dioxide to form small pieces of a safe, unreactive solid similar to glass.

The nitrogen gas in the airbag must inflate to approximately 67.0 L at a certain pressure in order for the airbag to be safe and effective. Automobile designers and engineers need to determine what quantities of the various reactants are required to achieve this inflation volume. Airbags are designed to slow the forward motion of the average adult male. Children are generally smaller, so they are at risk of serious injury if they are in the front seat when an airbag inflates.

WEB LINK

To learn more about how air bags work,

GO TO NELSON SCIENCE

The same stoichiometric principles that you developed in Unit 3, combined with an understanding of the ideal gas law equation, will allow you to solve problems similar to the airbag situation.

Tutorial 2 Mass or Volume-to-Mass or Volume Stoichiometry

In many stoichiometry problems you are given the mass, amount, or volume of gaseous reactants and/or products and asked to determine a corresponding quantity. Again, you will need the balanced chemical equation to determine the mole ratios. You may have to use molar mass, molar volume, the ideal gas law, or other appropriate concepts or equations as part of the solution. Make sure that your answer makes sense and is stated in the required units.

Sample Problem 1: Determining the Volume of a Gaseous Product

What volume of carbon dioxide is produced when 6.40 g of methane gas, $CH_4(g)$, reacts with excess oxygen? All gases are at 35.0 °C and 100.0 kPa.

Given: mass of methane, $m_{CH_4} = 6.40$ g

molar mass of methane, $M_{CH_4} = 16.05$ g/mol

$t = 35.0$ °C and

$P = 100.0$ kPa.

Required: volume of carbon dioxide, V_{CO_2}

Solution:

Step 1. Convert the temperature value(s) to kelvins.

$T = 35.0 + 273$

$T = 308$ K

Step 2. Write the balanced chemical equation, listing the given and required quantities, with the appropriate units, underneath.

$CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$

6.40 g $\qquad V_{CO_2}$

Step 3. Convert the mass of the given substance, m_{CH_4}, into an amount, n_{CH_4}, using a conversion factor derived from the molar mass of methane, CH_4.

$$n_{CH_4} = 6.40\ \cancel{g} \times \frac{1\ \text{mol}}{16.05\ \cancel{g}}$$

$$n_{CH_4} = 0.3988\ \text{mol (extra digits carried)}$$

Step 4. Determine the amount of carbon dioxide produced from the amount of methane used using the appropriate mole ratio derived from the balanced chemical equation.

$$n_{CO_2} = 0.3988\ \cancel{\text{mol}_{CH_4}} \times \frac{1\ \text{mol}_{CO_2}}{1\ \cancel{\text{mol}_{CH_4}}}$$

$$n_{CO_2} = 0.3988\ \text{mol}_{CO_2}$$

Step 5. Determine the volume of the required substance, V_{CO_2}, using the ideal gas law equation, $PV = nRT$. Remember that $R = 8.314\ \text{kPa}\cdot\text{L}\cdot\text{mol}^{-1}\cdot\text{K}^{-1}$.

$$PV = nRT$$

$$PV_{CO_2} = n_{CO_2}RT$$

$$V_{CO_2} = \frac{n_{CO_2}RT}{P}$$

$$= \frac{(0.3988\ \cancel{\text{mol}})\ (8.314\ \cancel{\text{kPa}}\cdot\text{L}\cdot\cancel{\text{mol}^{-1}}\cdot\cancel{\text{K}^{-1}})\ (308\ \cancel{\text{K}})}{100.0\ \cancel{\text{kPa}}}$$

$$V_{CO_2} = 10.2\ \text{L}$$

Statement: When 6.40 g of methane reacts with excess oxygen, 10.2 L of carbon dioxide is produced (under the given conditions of temperature and pressure).

Sample Problem 2: Determining the Mass of a Gaseous Reactant

What mass of sodium azide is required to produce the 67.0 L of nitrogen gas that is required to fill a safety airbag in an automobile? Assume that the gas is produced at a temperature of 32 °C and a pressure of 105 kPa.

Given: volume of nitrogen gas, $V_{N_2} = 67.0$ L

temperature, $t = 32$ °C

pressure, $P = 105$ kPa

Required: mass of sodium azide, m_{NaN_3}

Solution:

Step 1. Convert the temperature value(s) to kelvins.

$$T = 32 + 273$$

$$T = 305 \text{ K}$$

Step 2. Write the balanced chemical equation, listing the given and required quantities, with the appropriate units, underneath.

$$2\ NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g)$$

$$m_{NaN_3} \qquad\qquad\qquad 67.0 \text{ L}$$

Step 3. Determine the amount of the given substance, n_{N_2}, produced in the reaction using the ideal gas law equation, $PV = nRT$.

$$PV = nRT$$

$$PV_{N_2} = n_{N_2}RT$$

$$n_{N_2} = \frac{PV_{N_2}}{RT}$$

$$= \frac{(105\ \cancel{\text{kPa}})\ (67.0\ \cancel{\text{L}})}{(8.314\ \cancel{\text{kPa}} \cdot \cancel{\text{L}} \cdot \text{mol}^{-1} \cdot \cancel{\text{K}^{-1}})(305\ \cancel{\text{K}})}$$

$$n_{N_2} = 2.774 \text{ mol (extra digits carried)}$$

Step 4. Determine the amount of sodium azide used to produce the amount of nitrogen gas determined in Step 3, using the appropriate mole ratio derived from the balanced chemical equation:

$$n_{NaN_3} = 2.774\ \cancel{\text{mol}_{N_2}} \times \frac{2\ \text{mol}_{NaN_3}}{3\ \cancel{\text{mol}_{N_2}}}$$

$$n_{NaN_3} = 1.849\ \text{mol}_{NaN_3}$$

Step 5. Determine the mass of sodium azide produced from the amount of sodium azide determined in Step 4 by multiplying this amount by a conversion factor derived from the molar mass of sodium azide.

$$M_{NaN_3} = 65.0 \text{ g/mol}$$

$$m_{NaN_3} = n_{NaN_3} \times \frac{65.0 \text{ g}}{1 \text{ mol}}$$

$$= 1.849\ \cancel{\text{mol}} \times \frac{65.0 \text{ g}}{1\ \cancel{\text{mol}}}$$

$$m_{NaN_3} = 120 \text{ g}$$

Statement: 120 g of sodium azide is required to produce 67.0 L of nitrogen gas at these conditions.

Practice

SKILLS HANDBOOK A6

1. A 128 g sample of oxygen gas reacts completely with excess nitrogen monoxide according to the following balanced chemical equation:

   $2\ NO(g) + O_2(g) \rightarrow 2\ NO_2(g)$

   What volume of nitrogen dioxide will be produced at STP? T/I [ans: 179 L]
2. Sodium carbonate and hydrochloric acid react to give sodium chloride, carbon dioxide, and water. Calculate the mass of sodium carbonate that would be required to produce 2.00 L of carbon dioxide at 300.0 K and 101.3 kPa. T/I [ans: 8.61 g]
3. The following chemical equation shows the complete combustion of ethane:

   $2\ C_2H_6(g) + 7\ O_2(g) \rightarrow 6\ H_2O(g) + 4\ CO_2(g)$

   How many litres of oxygen would be required for the complete combustion of 82.0 L of ethane at 123 °C and 105 kPa? T/I [ans: 287 L]

UNIT TASK BOOKMARK

As you work on your Unit Task, on page 616, think about how small-scale methods of capturing carbon dioxide be scaled up to be used on an industrial level.

Investigation 12.5.1

Gas Stoichiometry: Determining the Mass of Hydrogen Gas (p. 606)
In this investigation you will set up a reaction that produces hydrogen gas. You will measure the volume of the gas, and the experimental conditions. You will use stoichiometry to determine the amount of gas produced. You will then combine these data to calculate the molar volume of hydrogen, and compare this value to the accepted molar volume.

Mini Investigation

Capturing Carbon Dioxide on a Small Scale

Skills: Planning, Performing, Observing, Analyzing, Communicating — SKILLS HANDBOOK A1, A2.4

Limewater, or calcium hydroxide solution, $Ca(OH)_2(aq)$, reacts to "capture" carbon dioxide gas. Carbon capture is a hot topic these days as global efforts unfold to control the emission of greenhouse gases. Many different technologies and solutions are currently being developed to sequester, capture, and convert carbon dioxide. Try your hand at carbon capture.

Equipment and Materials: chemical safety goggles; apron; plastic straw; small beaker containing limewater, $Ca(OH)_2(aq)$; dropper bottle containing bromothymol blue indicator

Do not suck up the limewater. If you get any limewater in your mouth, spit it into the sink immediately and rinse your mouth thoroughly.

1. Put on your chemical safety goggles and lab apron.
2. Place the plastic straw in the limewater. Gently exhale through the straw to bubble your breath into the limewater (**Figure 4**). Record your observations.
3. Dispose of your limewater according to your teacher's directions.
4. Obtain a fresh sample of limewater. Add 2 or 3 drops of bromothymol blue to the limewater and repeat Steps 2 and 3.

A. Write the balanced chemical equation for the reaction between calcium hydroxide and carbon dioxide. K/U C

B. What experimental evidence do you have to support your proposed chemical equation? T/I

C. Bromothymol blue is an acid–base indicator. Why did the colour change as carbon dioxide was added to the limewater? T/I

D. Would this be an effective method to remove carbon dioxide from the atmosphere? Explain. A

E. Design an experiment, based on this test, to determine the difference in the percentage of carbon dioxide in exhaled air compared with the percentage in atmospheric air.

Figure 4 Always blow out through limewater.

12.5 Summary

- The law of combining volumes for gas reactions states that volumes of gaseous reactants and products of a chemical reaction are always in simple ratios of whole numbers, when measured at the same temperature and pressure. This can be used to solve simple gas stoichiometry problems.
- Stoichiometric principles and the use of the ideal gas law equation can be used to solve a variety of gas stoichiometry problems.

12.5 Questions

1. Butane gas, C_4H_{10}, undergoes complete combustion with excess oxygen to produce carbon dioxide and water vapour. T/I C
   (a) Write the equation for the total combustion of butane.
   (b) Determine the volume of carbon dioxide gas that would be produced by the complete combustion of 5.60 L of butane. Assume that all gases are at STP.
2. Potassium chlorate decomposes according to the following equation:

   $2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$

   What volume of a gas can be produced by the decomposition of 122.6 g of potassium chlorate measured under the following conditions? T/I
   (a) at STP
   (b) at SATP
3. Sodium metal with a total mass of 20.0 g is dropped into a beaker of water to produce hydrogen gas and sodium hydroxide. What volume of hydrogen gas will be produced if the temperature is 30 °C and the pressure is 125 kPa? T/I
4. Most cars today are powered by the combustion of gasoline. The main component of gasoline is octane, $C_8H_{18}(g)$, which burns as follows:

   $2\ C_8H_{18}(g) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$

   One of the products is carbon dioxide, a major cause of climate change. T/I
   (a) Calculate the mass and volume of carbon dioxide gas produced by the combustion of a tank of gasoline. The mass of gasoline in the tank is 64.2 kg. The reaction takes place at a temperature of 80 °C and a pressure of 101.3 kPa. Assume that there is excess oxygen available and that only the complete combustion reaction occurs.
   (b) Plants effectively change carbon dioxide to glucose and oxygen gas through the process of photosynthesis. A single mature tree can absorb carbon dioxide at a rate of approximately 22 kg per year (**Figure 5**):

   $6\ CO_2(g) + 6\ H_2O(g) \rightarrow C_6H_{12}O_6(g) + 6\ O_2(g)$

   Calculate the number of trees required to absorb, in one year, the carbon dioxide produced from one tank of gas (calculated in (a)).

Figure 5 Trees provide many benefits, including carbon dioxide removal.

5. Calcium carbonate reacts with hydrochloric acid as follows:
   $CaCO_3(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$ T/I
   (a) What volume of carbon dioxide, measured at 20 °C and 100.0 kPa, would be produced from 1025 g of calcium carbonate?
   (b) Calculate the mass of calcium carbonate required to produce 1550 L of carbon dioxide at 22.5 °C and 100.0 kPa.
6. Ammonium sulfate reacts with potassium hydroxide solution as follows:

   $(NH_4)_2SO_4(aq) + 2\ KOH(aq) \rightarrow 2\ NH_3(g) + K_2SO_4(aq) + 2\ H_2O(l)$

   Calculate the volume of ammonia gas, measured at 23 °C and 64 kPa, that could be produced from 264.0 g of ammonium sulfate and 280.0 g of potassium hydroxide. T/I
7. Propane gas burns in air:

   $C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$

   Determine the volume of carbon dioxide that will be formed when 54.0 g of propane, C_3H_8, reacts with excess oxygen at 25 °C and 202.1 kPa. T/I

12.6 Explore Applications of Gases

SKILLS MENU

- Researching
- Performing
- Observing
- Analyzing
- Evaluating
- Communicating
- Identifying Alternatives

Figure 1 Can snow burn?

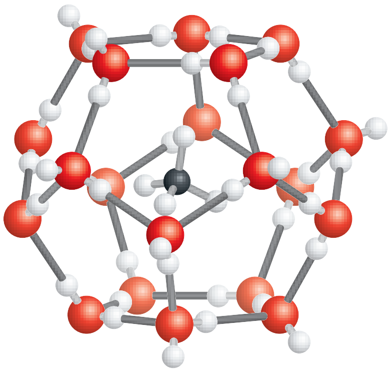

Figure 2 A methane molecule trapped in a cage of water molecules. This unusual structure is called a gas hydrate.

WEB LINK

To start your research,

GO TO NELSON SCIENCE

Burning Snowballs

It seems unlikely . . . a snowball of fire (**Figure 1**). What you see is a gas hydrate: molecules of gases, including methane, trapped inside ice (**Figure 2**). As you know, methane is a non-polar molecule and water is polar, so normally the two do not mix. Gas hydrates only form under low temperatures and high pressures. The gases are produced by the breakdown of organic matter. Some estimates predict that more organic carbon is contained in gas hydrates than in all other known fossil fuel sources combined: perhaps 10 000 G (gigatonne)!

The Application

Our ever-increasing demand for fossil fuels is encouraging exploration for unlikely substances. The sea floor has large deposits of gas hydrates. The extraction of gas hydrates is quite controversial because of the potential environmental impact and our continued reliance on fossil fuels.

You are a researcher with a new Canadian oil and gas exploration company. The company is considering "mining" gas hydrates as a source of natural gas. You are to conduct some initial research to help the company decide whether to pursue this idea. As the company is trying to be environmentally responsible, they also want information on the environmental impacts of extracting and using gas hydrates.

Your Goal

To provide the oil and gas company with objective, factual information to help them decide whether to explore for gas hydrates

Research

SKILLS HANDBOOK A5.1

Focus your research on the following topics:

- gas hydrate deposit locations around Canada
- cost of gas hydrate recovery, compared with other fossil fuels
- challenges of gas hydrate mining in Arctic waters
- environmental effects of gas hydrate recovery
- effects on climate change of release of gas hydrates from Arctic ice
- impact on air quality of burning methane compared to burning other fuels

Summarize

Extract the main points of your research, ensuring that you address all of the topics listed above.

Communicate

Summarize your findings in a brief, factual, unbiased report for the executives of the oil and gas company. Your report should be supported by references to reputable sources of information. Be prepared to also present an oral summary of your research results.

Plan for Action

While gas hydrate extraction might have its benefits, there is no denying the fact that burning methane releases carbon dioxide, a greenhouse gas. What can you do to reduce the need to burn fossil fuels? Is your school or your home heated by natural gas? Make a plan involving at least three strategies to reduce the quantity of natural gas used for heating.

GO TO NELSON SCIENCE

Investigation 12.4.1 OBSERVATIONAL STUDY

Determining the Molar Mass of Butane

SKILLS MENU

• Questioning	• Planning	• Observing
• Researching	• Controlling Variables	• Analyzing
• Hypothesizing	• Performing	• Evaluating
• Predicting		• Communicating

In this investigation you will determine the molar mass of butane, C_4H_{10}, using two methods. First, you will use the ideal gas law to determine the molar mass of butane experimentally. Second, you will use the periodic table to calculate it mathematically. You will then evaluate the design of this investigation by comparing the experimentally determined molar mass of butane with the accepted value.

Purpose

To determine the molar mass of butane experimentally

Equipment and Materials

SKILLS HANDBOOK A1.1, A1.2

- chemical safety goggles
- lab apron
- butane lighter (with flint removed) or butane cylinder with tubing
- plastic bucket
- 100–500 mL graduated cylinder
- retort stand with utility clamp
- balance
- thermometer
- barometer

Butane is highly flammable. There must be no sparks or open flames in the laboratory. The laboratory must be well ventilated.

Procedure

SKILLS HANDBOOK A2.4

1. Put on your chemical safety goggles and lab apron.
2. Determine the mass of the butane lighter.
3. Fill the bucket with tap water until it is two-thirds full. Fill the graduated cylinder with water. Invert the graduated cylinder into the bucket. Ensure that no air remains in the graduated cylinder. Secure the inverted graduated cylinder with a clamp and retort stand or hold it in place manually.
4. Hold the butane lighter in the water directly under the graduated cylinder. Press the button to release the gas from the lighter. Collect the gas in the graduated cylinder. Ensure that all the gas bubbles enter the cylinder (**Figure 1**). Fill half to three-quarters of the graduated cylinder with gas.

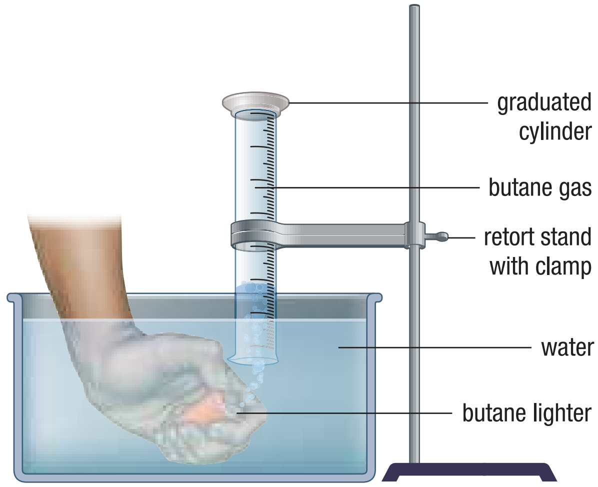

Figure 1

5. Equalize the pressure inside and outside the cylinder by adjusting the position of the cylinder until the water levels inside and outside the cylinder are identical.
6. Record the volume of the gas collected as indicated by the water level in the graduated cylinder. Record the ambient room temperature and pressure. Record your data as accurately as possible given the equipment available.
7. Thoroughly dry the butane lighter. Determine the final mass of the lighter.
8. Release the butane gas from the cylinder outdoors or in the fume hood away from open flames.

Analyze and Evaluate

SKILLS HANDBOOK A6.5

(a) Using the mass of the butane lighter before and after the investigation, determine the mass of butane released into the cylinder. Remember to correct for vapour pressure of water. T/I

(b) Using the data collected, the molar mass equation, and the ideal gas law equation, determine the experimental molar mass of butane. K/U T/I

(c) Determine the theoretical or accepted molar mass of butane using data from the periodic table. T/I

(d) Compare the molar mass determined experimentally to the accepted value by determining the percentage difference. T/I

(e) Use the percentage difference to evaluate the procedure. T/I

(f) Identify any problems that you encountered in this investigation. Suggest any improvements to this procedure. T/I A

Apply and Extend

(g) Collecting a gas over water affects the results. Research the effects of the use of water in gas collection. What effect, if any did the method of gas collection have on your results? T/I A

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Investigation 12.5.1 OBSERVATIONAL STUDY

Gas Stoichiometry: Determining the Mass of Hydrogen Gas

SKILLS MENU

- Questioning
- Researching
- Hypothesizing
- Predicting
- Planning
- Controlling Variables
- Performing
- Observing
- Analyzing
- Evaluating
- Communicating

The purpose of this investigation is to use stoichiometry to predict the molar volume of hydrogen. The gas will be produced in a reaction between zinc and hydrochloric acid:

$$Mg(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

The volume of hydrogen gas produced by this reaction will be calculated using stoichiometry and the ideal gas law. These results will be used to calculate the molar volume of hydrogen at STP.

The reaction will take place inside an inverted graduated cylinder or gas collection tube. The cylinder will contain water and hydrochloric acid. A piece of magnesium ribbon (the limiting reagent) will be held in place by copper wire. Hydrochloric acid is denser than water. Although it is initially placed in the bottom of the graduated cylinder, when the tube is inverted the acid will flow down to meet the magnesium. The gas produced by the reaction will collect in the top of the graduated cylinder.

Purpose

To determine the molar volume of hydrogen at STP using experimental data

Equipment and Materials

SKILLS HANDBOOK A1.1, A1.2

- chemical safety goggles
- lab apron
- disposable plastic gloves
- steel wool
- centigram or analytical balance
- gas collection tube or 100 mL graduated cylinder
- 20 mL pipette with pipette pump
- 2-hole stopper to fit cylinder
- large beaker (600 or 1000 mL)
- 250 mL beaker
- thermometer
- barometer
- magnesium ribbon (30 to 50 mm)
- piece of fine copper wire (100 to 150 mm)
- hydrochloric acid (3.00 mol/L)
- water

Hydrochloric acid is corrosive. If you get acid on your skin or clothes, or in your eyes, wash the affected area with plenty of cool water, and inform your teacher.

Hydrogen gas, produced in this reaction, is flammable. Ensure that the room is well ventilated. There must be no sparks or open flames in the laboratory.

Procedure

SKILLS HANDBOOK A2.4

1. Put on your chemical safety goggles, lab apron, and gloves.
2. Clean the magnesium ribbon by rubbing it with steel wool. Measure and record the mass of the magnesium ribbon.
3. Coil or fold the magnesium ribbon so that it can be held by the copper-wire cage (**Figure 1**).
4. Wrap the copper wire around the magnesium, making a cage to contain the magnesium. Leave 30 to 50 mm of wire at each end to attach the wire to the stopper.
5. Fill the gas collection tube or graduated cylinder with water to within 15 mL of the top.
6. Using a pipette pump, fill a pipette with 10 to 15 mL hydrochloric acid. Insert the pipette to the bottom of the graduated cylinder and slowly pump the acid out. The liquid at the top should be primarily water and the hydrochloric acid should remain at the bottom.
7. Half-fill the large beaker with tap water.
8. Bend the copper wire through the holes in the stopper so that the magnesium bundle is positioned about 10 mm from the bottom of the stopper (Figure 1).

Figure 1

9. Over a sink, insert the stopper into the graduated cylinder. The liquid may overflow a little. Cover the holes on the stopper with your gloved fingers and quickly invert the cylinder into the beaker. Once the stopper is below the surface of the water, remove your fingers from the stopper holes.
10. Observe the reaction for approximately 5 min or until gas production stops. Allow the contents of the graduated cylinder to reach room temperature.
11. Adjust the graduated cylinder by raising or lowering it until the level of liquid inside the graduated cylinder is the same as the level of liquid in the beaker. This ensures that the gas pressure in the cylinder is equal to the air pressure in the room.
12. Measure and record the volume of the gas produced in the graduated cylinder.
13. Measure and record the ambient temperature and pressure in the laboratory.
14. Cover the holes in the stopper with your gloved fingers and lift the cylinder out of the beaker.
15. Carefully pour the liquids from this investigation down the drain. Rinse the sink with plenty of water.
16. All the magnesium should have been consumed as magnesium is the limiting reagent. If any magnesium remains, rinse and dry the magnesium strip and record the mass remaining. Determine the mass of magnesium consumed in the reaction.

Analyze and Evaluate

SKILLS HANDBOOK A6.5

(a) The gas in the cylinder is a combination of the hydrogen gas produced and water vapour. The total pressure of this gas mixture is equal to the atmospheric pressure. To determine the pressure of hydrogen, take the total pressure and subtract the vapour pressure of water (according to Dalton's law of partial pressure). Table 1 in Section 12.4 shows the vapour pressure of water at various temperatures. T/I

(b) Determine the amount of magnesium consumed in this reaction, then predict the amount of hydrogen gas that should be produced. T/I

(c) Determine the experimental molar volume of hydrogen gas produced during this investigation at your experimental conditions. Use your collected data and the ideal gas law. T/I

(d) Using the combined gas law, convert the experimental molar volume of hydrogen gas produced at your experimental conditions to a molar volume of hydrogen gas at STP. T/I

(e) How does your answer in (d) compare with the theoretical molar volume of any gas at STP (22.4 L/mol)? T/I

(f) What are some sources of error in this investigation? What could be done to improve the design? T/I A

(g) Why was the graduated cylinder held in water? T/I

Apply and Extend

(h) Predict what would happen if you used a stopper without holes in the graduated cylinder. T/I

CHAPTER 12 SUMMARY

Summary Questions

1. Create a study guide based on the points listed in the margin on page 574. For each point, create sub-points that provide further information, relevant examples, explanatory diagrams, or general equations.
2. Look back at the Starting Points questions on page 574. Answer these questions using what you have learned in this chapter. Compare your latest answers with those that you wrote at the beginning of the chapter. Note how your answers have changed.
3. Prepare a list of assessment questions based on the material in this chapter. Your questions should include a variety of multiple-choice questions, fill-in-the-blank questions, short-answer questions (including diagrams), and problems.

Vocabulary

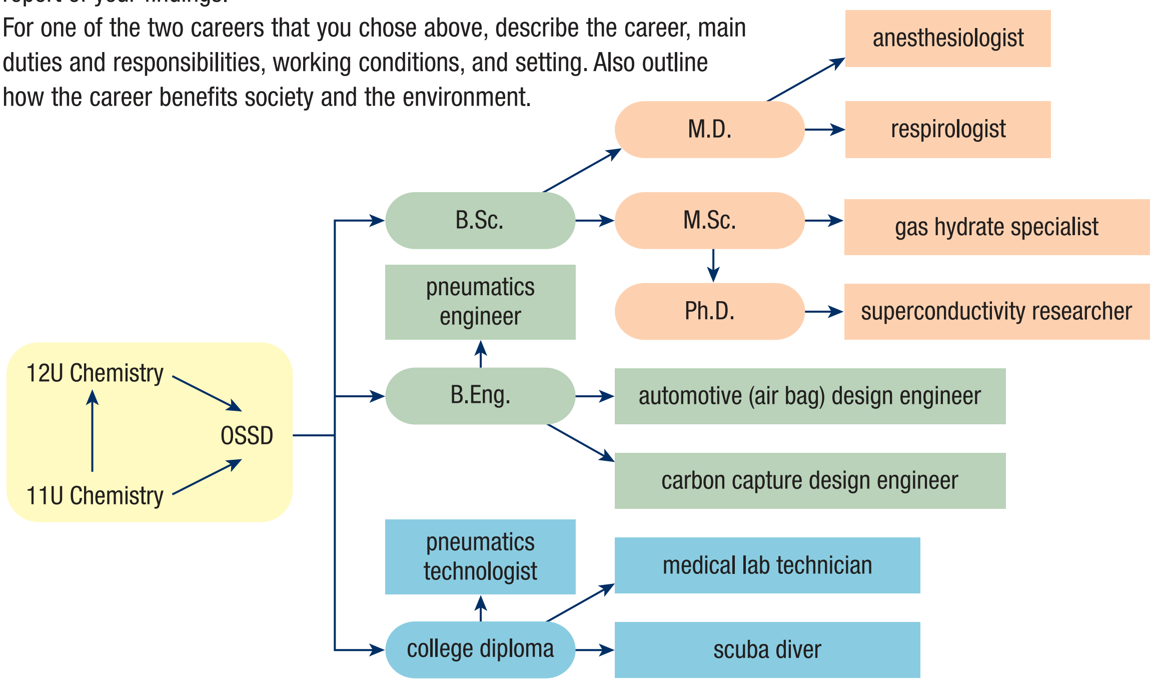
law of combining volumes (p. 576)
Avogadro's law (p. 578)
molar volume (p. 579)
ideal gas (p. 582)
universal gas constant (*R*) (p. 583)
ideal gas law (p. 583)
partial pressure (p. 593)
Dalton's law of partial pressures (p. 593)

CAREER PATHWAYS

SKILLS HANDBOOK A7

Grade 11 Chemistry can lead to a wide range of careers. Some require a college diploma or a B.Sc. degree. Others require specialized or post-graduate degrees. This graphic organizer shows a few pathways to careers mentioned in this chapter.

1. Select two careers, related to the behaviour of gases, that you find interesting. Research the educational pathways that you would need to follow to pursue these careers. What is involved in the required educational programs? Prepare a brief report of your findings.
2. For one of the two careers that you chose above, describe the career, main duties and responsibilities, working conditions, and setting. Also outline how the career benefits society and the environment.



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K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

For each question, select the best answer from the four alternatives.

1. According to Gay-Lussac's law, how many litres of hydrogen gas, H_2, react with 2 L of nitrogen gas, N_2, to produce 4 L of ammonia gas, NH_3? (12.1) K/U
   (a) 2 L (c) 4 L
   (b) 3 L (d) 6 L
2. According to Avogadro's law, which of the following factors must be constant for equal volumes of gases to contain equal numbers of particles? (12.1) K/U
   (a) mass and pressure
   (b) temperature and mass
   (c) temperature and pressure
   (d) mass, pressure, and temperature
3. How do intermolecular forces among entities in an ideal gas compare to those among entities in a real gas? (12.2) K/U
   (a) Intermolecular forces are greater among entities in an ideal gas.
   (b) The intermolecular forces for real gases equal those for ideal gases.
   (c) Intermolecular forces exist among entities of real gases, but do not exist among entities of ideal gases.
   (d) The intermolecular forces are similar, but not equal.
4. According to the ideal gas law, the product of the pressure and the volume of a gas
   (a) is proportional to the product of the amount and the absolute temperature of the gas.
   (b) equals the universal gas constant, *R*.
   (c) equals the product of the amount of gas and its temperature.
   (d) is proportional to the product of the amount and the Celsius temperature of the gas. (12.2) K/U
5. Which of the following scientists liquefied gases in their attempts to reach absolute zero? (12.3) K/U
   (a) Kelvin and Dewar
   (b) Dewar and Onnes
   (c) Bose and Einstein
   (d) Lambert and van der Waals
6. Hydrogen gas was collected over water at standard atmospheric pressure and 23 °C. The vapour pressure of the water is 2.81 kPa. What is the pressure of the hydrogen gas? (12.4) K/U
   (a) 2.81 kPa
   (b) 97.19 kPa
   (c) 98.49 kPa
   (d) 104.1 kPa
7. What causes the pressure exerted by a gas? (12.4) K/U
   (a) the mass of the gas molecules only
   (b) the speed of the gas molecules only
   (c) the collisions of gas molecules with each other
   (d) the collisions of gas molecules with the walls of their container
8. Which gas law applies when collecting gases over water? (12.4) K/U
   (a) Charles' law
   (b) Gay-Lussac's law
   (c) Boyle's law
   (d) Dalton's law of partial pressures
9. Where are many large deposits of gas hydrates found? (12.6) K/U
   (a) on the sea floor
   (b) on the Moon
   (c) in the Arctic ice
   (d) in the tundra

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

10. Hydrogen and oxygen gases are produced when electrical energy passes through water. (12.1) K/U
11. Collisions of entities of an ideal gas are not elastic. (12.2) K/U
12. Gases behave most ideally at low pressures and low temperatures. (12.2) K/U
13. A Bose–Einstein condensate is a rare state of matter reached at very low temperatures. (12.3) K/U
14. Dalton's law of partial pressure can be explained by the kinetic molecular theory. (12.4) K/U
15. To find the total pressure of a mixture of water vapour and carbon dioxide, find the pressure of the water vapour and subtract the pressure of the carbon dioxide. (12.4) K/U
16. To do volume-to-volume stoichiometry problems, at least one of the reactants and/or products must be a gas. (12.5) K/U
17. To do any stoichiometry problem, you need a balanced chemical equation. (12.5) K/U
18. Gas hydrates form under low temperatures and low pressures. (12.6) K/U

To do an online self-quiz,
GO TO NELSON SCIENCE

K/U Knowledge/Understanding T/I Thinking/Investigation C Communication A Application

Knowledge

For each question, select the best answer from the four alternatives.

1. Which of the following laws describes the observations that Gay-Lussac made about volumes of gases in chemical reactions? (12.1) K/U
   (a) Avogadro's law
   (b) the law of combining volumes
   (c) Boyle's law
   (d) the law of molar volume
2. What volume is occupied by 2.0 mol of a gas at STP? (12.1) K/U
   (a) 22.4 L
   (b) 24.8 L
   (c) 44.8 L
   (d) 49.6 L
3. How does the volume of an entity of an ideal gas compare to the volume of its container? (12.2) K/U
   (a) The volumes are equal.
   (b) The volumes are similar, but not equal.
   (c) The volume of the gas particle is smaller, but it is measurable.
   (d) The volume of the gas particle is so small as to be insignificant.
4. One form of the universal gas constant has a numerical value of 8.314. For this value of *R*, what are the units for pressure, volume, amount, and temperature, respectively? (12.2) K/U
   (a) atm, mL, mol, K
   (b) kPa, L, mol, K
   (c) atm, L, mol, °C
   (d) kPa, mL, mol, °C
5. The search for absolute zero was inspired by which of the following factors? (12.3) K/U
   (a) an increase in funding for research
   (b) recent expeditions to the Arctic and Antarctic
   (c) the discovery of the noble gases
   (d) the development of the thermometer
6. Nitrogen gas was collected over water at 100.5 kPa and 22 °C. The vapour pressure of the water is 2.64 kPa. What is the pressure of the nitrogen gas? (12.4) K/U
   (a) 2.64 kPa
   (b) 97.86 kPa
   (c) 101.3 kPa
   (d) 103.1 kPa
7. Which of the following statements is true about most naturally occurring gases? (12.4) K/U
   (a) They are mixtures.
   (b) They are elements.
   (c) They are compounds.
   (d) They are ions.
8. For the volume of a gaseous reactant in a chemical reaction and the volume of a gaseous product to be in a whole-number ratio, what must be true? (12.5) K/U
   (a) The reaction occurs at STP.
   (b) Temperature and pressure must remain constant.
   (c) All reactants and products must be gases.
   (d) The volumes of the gaseous reactant and the gaseous product are in a 1:1 ratio.
9. What is the name given to a structure formed when a methane molecule is trapped inside a cage of water molecules? (12.6) K/U
   (a) fossil fuel
   (b) organic matter
   (c) gas hydrate
   (d) natural gas

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

10. One mole of a gas has a volume of 22.4 L at ambient temperature and pressure. (12.1) K/U
11. One mole of nitrogen gas and one mole of oxygen gas contain the same number of entities under the same conditions of temperature and pressure. (12.1) K/U
12. Ideal gases do not condense into liquids when cooled. (12.2) K/U
13. A gas at high temperature behaves more like a real gas than does a gas at low temperature. (12.2) K/U
14. Entities in a Bose–Einstein condensate behave like individual atoms. (12.3) K/U
15. For Dalton's law of partial pressure to apply, the gases must not react. (12.4) K/U
16. When temperature and pressure vary during a chemical reaction, Dalton's law of partial pressure can be used to calculate volumes. (12.5) K/U
17. The first step to solving any stoichiometry problem is to write a balanced chemical equation for the reaction. (12.5) K/U
18. Canadian oil and gas companies are considering mining gas hydrates as a source of natural gas. (12.6) K/U

Match the scientist on the left with what that person is known for on the right.

19. (a) Gay-Lussac
    (b) Avogadro
    (c) Cornell and Wieman
    (d) Dalton

    (i) relating V and n for a gas if T and P are constant
    (ii) Bose–Einstein condensate
    (iii) law of partial pressures
    (iv) law of combining volumes (12.1, 12.3, 12.4) K/U

Write a short answer to each question.

20. What did Gay-Lussac's observations of chemical reactions involving gases show? (12.1) K/U
21. State Avogadro's law. (12.1) K/U
22. Why is molar volume at ambient temperature and pressure greater than molar volume at STP? (12.1) K/U
23. Describe the motion of the entities in an ideal gas. (12.2) K/U
24. What does the following statement mean? "The collisions among entities in an ideal gas are perfectly elastic." (12.2) K/U
25. What three gas laws combine to form the ideal gas law? (12.2) K/U
26. What is a Bose–Einstein condensate? In your definition, include how this state of matter differs from others. (12.3) K/U
27. Has absolute zero ever been reached? Explain your answer. (12.3) K/U
28. A sample of air was analyzed. The oxygen in the air exerted a pressure of 20.0 kPa, and the nitrogen exerted a pressure of 78.0 kPa. Other gases, such as argon and carbon dioxide, added a combined pressure of 3.0 kPa. What was the total pressure of the air? (12.4) K/U
29. What are two requirements of a mixture of gases for Dalton's law of partial pressures to apply to it? (12.4) K/U
30. Why is the following statement *not* true according to Dalton's law of partial pressures? "In a mixture of argon and neon, if neon has a pressure of 12 kPa and argon has a pressure of 15 mm Hg, the total pressure of the gases is 27 kPa." (12.4) K/U
31. Why can you use volume to solve stoichiometry problems involving gases but not those involving only other states? (12.5) K/U
32. Why must temperature and pressure remain constant when solving stoichiometry problems involving volume? (12.5) K/U
33. Why do airbags contain a chemical that reacts quickly to produce a gas? (12.5) K/U
34. Why might you be surprised that methane and water together form a gas hydrate? (12.6) K/U
35. What environmental problem might be the result of burning gas hydrates? (12.6) K/U
36. Why are a gas hydrate and a compound such as copper(II) sulfate pentahydrate, $CuSO_4•5H_2O$, both called hydrates? (12.6) K/U

Understanding

37. Ammonia gas, NH_3, is found in many common products, including fertilizers and cleaners. The commercial production of ammonia involves making the compound from hydrogen gas, H_2, and nitrogen gas, N_2. (12.1) K/U T/I C
    (a) Write a balanced chemical equation for the formation of ammonia gas from hydrogen gas and nitrogen gas.
    (b) According to Gay-Lussac's observations, what volume of nitrogen will react with 4.5 L of hydrogen?
    (c) What volume of ammonia is formed from these volumes of reactants?
38. Helium is commonly used to fill balloons. After 1.5 mol of helium gas is added to a balloon, the volume of the balloon is 27.3 L. What will be the volume of the balloon if another 2.4 g of helium is added to the balloon, and temperature and pressure remain constant? (12.1) T/I
39. Neon is a noble gas that produces the red glow in "neon" lights. What is the volume of 6.4 mol of neon gas at STP? (12.1) K/U T/I
40. Argon has uses that range from lighting to welding. What mass of argon has a volume of 12.3 L at standard ambient temperature and pressure? (12.1) K/U T/I
41. Examine the molar volumes in **Table 1**. (12.2) K/U T/I A

Table 1 Actual Molar Volumes of Various Real Gases at STP

Gas	Molar volume at STP (L/mol)
any ideal gas	22.414
helium, He	22.426
oxygen, O_2	22.397
chlorine, Cl_2	22.063
ammonia, NH_3	22.097

    (a) Why do you think the molar volume of helium is higher than the molar volume of the other real gases that are listed?
    (b) Would you expect the molar volume of neon to be closer to that of helium or closer to that of the other listed real gases? Explain your answer.

42. The ideal gas law is $PV = nRT$. Show how this equation can be used to find each of the following variables. In your description, include any additional information you need to know to make the calculation. (12.2) K/U T/I
   (a) the mass, m, of a gas
   (b) the molar mass, M, of a gas
   (c) the density, d, of a gas
43. A student poured hydrochloric acid on a piece of magnesium and then collected the 27 mL of hydrogen gas produced. The temperature in the lab was 24 °C, and the pressure was 102 kPa. What amount of hydrogen gas was produced? (12.2) K/U T/I
44. Valuable documents might react with oxygen or other reactive components of air. These reactions could damage the ink, paper, and binding. Nitrogen gas is only slightly reactive and it sometimes is used as the gas in sealed containers for storing such valuable documents. Determine the density of 2.5 mol of pure nitrogen gas at STP. (12.2) K/U T/I
45. If you read the label on a bottle of chlorine bleach, you will find cautions about how the bleach is to be used. If used improperly, chlorine bleach decomposes, releasing toxic chlorine gas, Cl_2. What amount of chlorine gas is released from chlorine bleach if the gas occupies 64 mL at 22 °C and 102 kPa? (12.2) K/U T/I
46. Use what you know about intermolecular forces to explain why hydrogen was the last of the permanent gases to be liquefied. (See Section 3.4 if you need to review intermolecular forces.) (12.3) K/U T/I
47. (a) In terms of the kinetic molecular theory, what happens to the motion of entities as they approach absolute zero?
   (b) How does your answer to Part (a) help to explain why researchers noticed that atoms stop behaving as individual atoms at very cold temperatures? (12.3) K/U T/I
48. Why does Dalton's law of partial pressures specify that the mixture of gases must be non-reacting gases? (12.4) T/I
49. Propane is a by-product of the processing of natural gas. It is readily available in Canada. As a fuel, propane burns completely only when it is mixed with the right proportion of air. Suppose that, in a mixture of three gases in a closed container, propane exerts a pressure of 5.0 kPa, argon exerts a pressure of 19 kPa, and the pressure exerted by nitrogen is 79 kPa. What is the total pressure of the gas mixture? (12.4) K/U A
50. Electrical energy is used to decompose water in a beaker. The hydrogen and oxygen produced are collected in two separate test tubes (**Figure 1**). (12.4) K/U T/I A

Figure 1

   (a) In the "hydrogen" test tube, the partial pressure of hydrogen is 98.5 kPa, and the temperature is 23 °C. What is the total pressure in the test tube? (Refer to Table 1, Vapour Pressure of Water at Various Temperatures, in Section 12.4.)
   (b) Predict how the total pressure in the oxygen test tube compares to the total pressure in the test tube containing hydrogen.
51. A closed, half-full bottle of pop contains carbon dioxide gas and water vapour, in addition to the pop. The temperature is 17.0 °C. If the pressure of the carbon dioxide in the bottle is 152 kPa, what is the total pressure of the gases in the bottle? (Refer to Table 1, Vapour Pressure of Water at Various Temperatures, in Section 12.4.) (12.4) K/U T/I A
52. Butane, C_4H_{10}, is a hydrocarbon that is sometimes used as a fuel for cooking. It also is used as a propellant in certain aerosol cans. (12.5) K/U T/I C
   (a) Write a balanced chemical equation for the combustion of butane to produce water vapour and carbon dioxide.
   (b) What volume of oxygen reacts completely with 4.0 L of butane, assuming that the temperature and pressure are constant?
   (c) What volume of water vapour is produced?
   (d) How many litres of carbon dioxide are also produced?

53. Alberta is a major producer of ethane. Ethane is used primarily to make other chemicals. It is also used as a fuel. Like all hydrocarbons, complete combustion produces water vapour and carbon dioxide. (12.5) K/U T/I C
   (a) Write a balanced chemical equation for the complete combustion of ethane gas, $C_2H_6(g)$.
   (b) How does the total volume of the reactants in Part (a) compare to the total volume of the products?
   (c) What volume of water vapour is produced from 56.0 g of ethane at STP? (Assume that excess oxygen is present.)
   (d) What volume of carbon dioxide forms from 78 g of oxygen at 98 kPa and 24.0 °C? (Assume that oxygen is the limiting reagent in a complete combustion reaction.)

Analysis and Application

54. A student has a Mylar balloon filled with helium. Another student has an identical balloon filled with air. The temperature and the pressure inside both balloons are the same. (12.1) K/U T/I
   (a) How do the numbers of entities in the two balloons compare?
   (b) How does the kinetic molecular theory explain your answer to Part (a)?
55. Radon gas is produced by the radioactive decay of uranium. Radon can contaminate soil and water and seep into buildings, where it is a health hazard because it is carcinogenic. Imagine the unrealistic situation of a basement completely filled with radon gas at standard ambient temperature and pressure. The dimensions of the basement are 10.0 m × 7.0 m × 3.0 m. What amount of radon is in the basement? (12.1) K/U T/I A
56. Copy **Figure 2** into your notebook and modify the diagram so that it also could be used to relate the mass of a gas to its volume at STP. (12.1) K/U T/I C

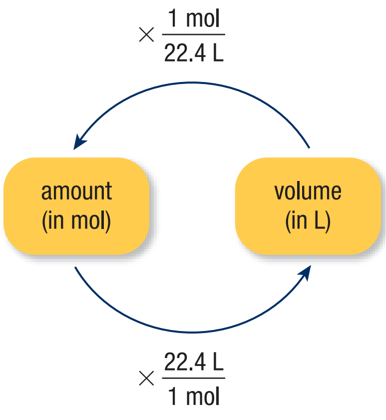


Figure 2

57. Before the party, a birthday balloon had a volume of 20.4 L at 25.0 °C and 101 kPa. During the party, the temperature in the room increased to 28.0 °C and the pressure fell to 98.0 kPa. Assuming that no gas leaked out of the balloon, what was the volume of the balloon during the party? (12.2) K/U T/I A
58. A closed container has a volume of 250 mL. Inside the container is 25.0 mL of liquid bromine, Br_2. The remaining volume is bromine vapour. The temperature is 20.0 °C, and the pressure is 102 kPa. What mass of bromine vapour is in the container? (12.2) K/U T/I
59. Use **Figure 3** to answer the following questions: (12.2) T/I A

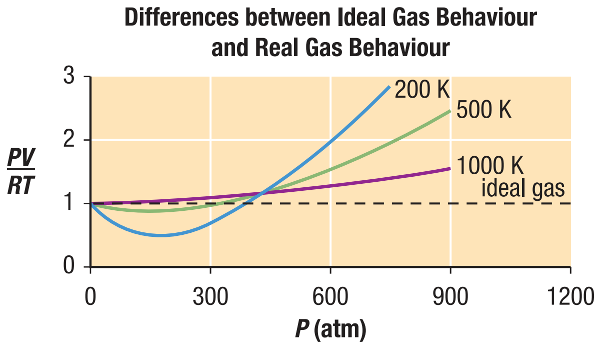


Figure 3

   (a) Predict the shape of a line representing the behaviour of a real gas at 100 K.
   (b) Use the graph to make a general statement of when the behaviour of a real gas most closely approaches the behaviour of an ideal gas. Make your statement in terms of the temperature and the pressure of the gas.
60. The label fell off a tank of compressed gas. Although the colour of the tank provides information about the identity of the gas, the company selling the gas wanted to confirm what gas is in the tank. They filled a 500.0 mL container with 1.55 g of the gas. The pressure of the gas was 240 kPa, and the temperature was 25.0 °C. (12.2) K/U T/I A
   (a) Assuming the gas acts as an ideal gas, what is the molar mass of the gas?
   (b) From the molar mass of the gas, identify the gas. Assume that the gas is an element, not a compound or a mixture.
61. Scuba divers often use a mixture of gases called nitrox in their tanks for shallow dives. Nitrox is a mixture of nitrogen and oxygen that might also contain other gases. The concentration of oxygen in nitrox for divers is greater than the concentration of oxygen in air. In terms of a tank of nitrox, explain Dalton's law of partial pressures. (12.4) K/U A

62. The following chemical equation for photosynthesis shows how green plants, in the presence of chlorophyll and sunlight, change carbon dioxide and water into oxygen and glucose:

$$6\ CO_2(g) + 6\ H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6\ O_2(g)$$

(12.5) K/U T/I A

(a) If 12.7 L of carbon dioxide is consumed during photosynthesis, what volume of oxygen is produced? (Assume that pressure and temperature remain constant and that there is an excess of water available.)

(b) From the answer to Part (a), a student infers that the masses of carbon dioxide and oxygen must also be in the same ratio. Evaluate this statement and justify your answer.

63. Yeast and some bacteria feed on glucose, $C_6H_{12}O_6(aq)$, during alcoholic fermentation. The products of the reaction are ethanol, $CH_3CH_2OH(l)$, and carbon dioxide gas. Energy is also released. (12.5) T/I C A

(a) Write a balanced chemical equation for alcoholic fermentation.

(b) What volume of carbon dioxide is produced from 2.00 mol of glucose at STP?

(c) If 82 g of ethanol is produced, what mass of carbon dioxide is also produced?

(d) From your answer to Part (b), why do you think alcoholic beverages are fermented before they are bottled or canned? What would happen if any fermentation took place in a closed container?

Evaluation

64. Suppose the weather station at Yellowknife in the Northwest Territories uses two identical weather balloons to send instruments into the atmosphere to measure air conditions. At ground level the balloons are kept at the same temperature and pressure. Before the balloons are released, a meteorologist adds 140.0 mol of gas to one balloon until it has a volume of 3.10 kL. Another meteorologist wants to fill the other balloon with the same gas to a volume of 2.60 kL. (12.1) K/U T/I A

(a) The second meteorologist calculates that she needs 167 mol of gas to fill the balloon to the desired volume. Does this calculation make sense? Explain your answer.

(b) What, if any, mistakes did the meteorologist make in her calculations?

(c) What is the correct amount of gas that should be added to the second balloon?

65. A student examined the following chemical equation:

$$HCl(g) + NH_3(g) \rightarrow NH_4Cl(s)$$

He stated that, according to the law of combining volumes, equal volumes of these three compounds are in the reaction. Explain whether the student was correct or not and justify your answer. (12.1) K/U T/I

66. A student wanted to use the ideal gas law to confirm that the molar volume of a gas is 22.4 L at STP. He substituted 1.00 atm for P, 1.00 mol for n, 8.314 kPa•L•mol^{-1}•K^{-1} for R, and 273 K for T. Instead of the calculation giving a result of 22.4 L, it was 2270 L. What mistake did the student make? (12.2) T/I

67. Suppose a mixture of 20.0 % argon and 80.0 % helium in a closed container exerts a total pressure of 158 kPa. (12.4) K/U T/I

(a) What pressure is exerted by each of the gases in the mixture?

(b) The gas mixture is moved to another closed container. This new container has twice the volume of the first container. The temperature remains constant. How do you determine what the pressure of each gas is in the new container?

(c) Calculate the pressure of each gas in the new container.

68. The oxygen gas produced during photosynthesis is consumed during the process of cellular respiration. This process produces most of the energy used by cells for growth, movement, and other processes. The chemical equation for cellular respiration is as follows:

$$C_6H_{12}O_6(aq) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(l) + \text{energy}$$

(12.5) K/U T/I A

(a) How does the chemical equation for cellular respiration compare to the chemical equation for photosynthesis? To review the equation for photosynthesis, see Question 62.

(b) After you exercise, your body is warmer. What can you infer about the volume of oxygen inhaled and the volume of carbon dioxide exhaled before exercising compared to the volumes inhaled and exhaled during and after exercising? Use the chemical equation for cellular respiration to help you.

Reflect on Your Learning

69. Some balloons are made of higher-quality materials than others. Suppose for a birthday party you purchased 10 inflated helium balloons of varying qualities. The balloons are kept at the same temperature and pressure. Two days later, some balloons are the same size as on the day you purchased them, but some are smaller. Use Avogadro's law to explain why the sizes of the balloons are now different. K/U T/I A

70. People with breathing problems may be helped by inhaling oxygen from a tank. The oxygen in the tanks is in liquid form. How does this example illustrate that oxygen is a real gas, not an ideal gas? K/U T/I A

71. How does the history of the work of James Dewar emphasize the importance of communication, cooperation, and safety in scientific research? K/U A

72. Canada is considered to be a world leader in fuel cell technologies. More than 80 companies across Canada are involved in developing this potential source of energy. In a hydrogen fuel cell, hydrogen gas combines with oxygen gas. The product of the reaction is water. T/I A
    (a) What do you think happens to the concentration of gases in this fuel cell over time?
    (b) What do you think might be done to the fuel cell to restore it to its original concentration of gases?

73. Probably the most important chemical reaction in the world is the one that occurs during photosynthesis. Green plants, in the presence of chlorophyll and sunlight, change carbon dioxide and water into oxygen and glucose according to the following chemical equation:

    $$6\ CO_2(g) + 6\ H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6\ O_2(g)$$

    K/U T/I A
    (a) How do carbon dioxide and oxygen affect your environment?
    (b) How do you think the reduction of the size of the world's rainforest affects what happens in this equation for photosynthesis?

Research

GO TO NELSON SCIENCE

74. Alberta is one of the world's largest suppliers of natural gas. This province produces almost 80 % of all the natural gas produced in Canada. Research answers to the following questions and prepare a short summary of your findings: K/U T/I C
    (a) Natural gas is composed of several gases. What compound is the main component of natural gas?
    (b) What are the products of burning this gas in oxygen? Write a balanced chemical equation for this reaction.
    (c) If 2 mol of this gas burns in a complete combustion reaction, what amount of oxygen is used? What amount of each product forms?

75. The preparation of super-cold condensates has important scientific implications. Explore Bose–Einstein and fermionic condensates. Compile your answers to the following questions into a presentation. You can choose your preferred format. T/I C A
    (a) What is a fermionic condensate?
    (b) Compare and contrast Bose–Einstein condensates and fermionic condensates. Include the different types of entities in each.
    (c) What practical uses do scientists predict for these super-cold states of matter?

76. Most common anesthetics used in hospitals contain mixtures of gases. Assemble your answers to the following questions into a poster that will be displayed in the waiting room in a hospital. K/U C A
    (a) What gases are most commonly used as the anesthetic agents?
    (b) What gas is used as the carrier for the anesthetic agents?
    (c) Why is an effective anesthetic a mixture of gases instead of just an anesthetic agent with no carrier gas?
    (d) According to the Canadian Centre for Pollution Prevention, anesthetic gases contribute to the emission of greenhouse gases into the atmosphere. Why do these researchers think anesthetic gases are a problem to the environment?

77. Methane is not the only gas that forms a gas hydrate. Another gas hydrate that is receiving attention is carbon dioxide hydrate. Research the following questions:
    - What benefit to the environment would result from making large quantities of carbon dioxide hydrate?
    - What location in Canada is being considered as a storage site for this hydrate?
    - How can the formation of carbon dioxide hydrate be related to obtaining supplies of gas hydrate containing methane?

    Prepare a web page to communicate the results of your research. K/U C A

What to Do with Carbon

What do you do when your room contains too much stuff? Do you throw your old gear into the nearest closet and hope it goes away? Do you sell your gently used belongings in a yard sale or donate them to charity? Or do you dismantle your possessions and construct different objects?

As you are probably aware, the concentration of carbon dioxide in our atmosphere is increasing. Our atmosphere now contains an average of 385 parts per million (ppm) of carbon dioxide. This is higher than at any time over the past 400 000 years. Most scientists agree that carbon dioxide is having significant effects on our climate. Furthermore, our activities continue to emit approximately 2.7×10^{10} t of carbon dioxide each year. As a result, the polar ice caps are melting and ecosystems are going through tremendous change. Scientists are warning that we need to both cut the emission of carbon dioxide and remove some of the carbon dioxide from our atmosphere. But how can we do this?

Just as you have to make difficult decisions about what to do with your unwanted stuff, we, as a society, have to decide what to do with the excess carbon dioxide in our environment (**Figure 1**). There are a number of options available.

Figure 1 What is the best way to trap carbon?

First, we must consider how to capture the carbon dioxide before it is released into the atmosphere. Carbon capture and storage (CCS) will likely become a boom industry in the years to come. This strategy involves capturing the carbon dioxide at the source of emission. The captured material is then isolated, or sequestered, underground. Scientists and engineers are working on technologies such as specially designed sponges that soak up carbon dioxide from smokestacks and tailpipes. These sponge-like materials have pores that trap carbon dioxide molecules. They can remove up to 83 times their own volume of carbon dioxide. Once these sponges are filled, they can be emptied and used again.

Another solution to the problem of carbon dioxide is to find alternative uses for it. For example, commercial greenhouses routinely pump carbon dioxide over their young plants to accelerate their growth (**Figure 2**).

Figure 2 A Toyota car manufacturing facility in Japan provides carbon dioxide that helps these young plants grow.

Strangely, many oil companies are eager to purchase carbon dioxide emissions from nearby factories. In a process known as enhanced oil recovery, water and compressed carbon dioxide are pumped into oil wells to extract oil that cannot otherwise be removed. Once all the oil is removed, the carbon dioxide remains stored underground.

How do we remove carbon dioxide once it has been released into the atmosphere? One alternative is to convert carbon dioxide into other carbon-based materials. The carbon cycle, which includes photosynthesis and respiration, continually moves carbon between the atmosphere, ocean, biosphere, and sediments. Mimicking some of these natural processes might help to decrease the concentration of atmospheric carbon dioxide. In Section 12.5, you may have performed an activity to capture carbon dioxide on a small scale.

Carbon dioxide removal may appear to be too expensive and technologically challenging. The process of decreasing emissions or converting carbon may use too much energy and not be feasible—today. In addition, there are concerns about the viability of long-term carbon storage and the possible environmental impacts. However, the consumption and manipulation of carbon is fast becoming economically desirable. Governments and industries are coming to terms with the economic realities of this issue.

In this task, you will choose a specific industry that emits carbon dioxide. For this industry, you will decide on the best plan to either capture waste carbon dioxide before it is emitted, or capture it in the atmosphere. You will prepare your plan in an original format. Finally, you will share the plan with your class and send the plan to the actual facility or industry.

The Issue

What should we do with the carbon dioxide in our emissions and/or in the atmosphere?

Goal

In this Unit Task you will develop, communicate, and present a plan for how to deal with carbon dioxide in industrial emissions and/or with the accumulation of carbon dioxide in our atmosphere.

Research

SKILLS HANDBOOK A5.1, A5.2

In pairs or small groups, conduct research on an industry that has significant carbon dioxide emissions. Some examples of industries are fossil fuel extraction, trucking, automobile assembly, oil refining, and chemical manufacturing. Get your teacher's approval on your chosen industry before you begin.

Research technological methods to deal with carbon dioxide emissions and/or carbon dioxide released in the atmosphere. The technologies could be already in operation or still in the planning stages. Technologies could include

- carbon capture and storage technologies (CCS)
- alternative uses for carbon dioxide
- the reaction of carbon dioxide to produce other carbon-based materials

As you research, consider the practicality and the economic cost of the proposed technologies as well as the long-term implications to human health and environmental sustainability. Also, consider how the gas laws and other concepts introduced in this unit relate to carbon removal.

Evaluate

Develop a strategy for decreasing carbon dioxide emissions at the source or for decreasing the concentration of carbon dioxide in the atmosphere. Your strategy, based on your research, should

- include an overview of recent emissions information and how your chosen industry has contributed to these emissions
- compare at least two technologies for dealing with carbon dioxide
- evaluate the technologies and select the one that you think is best, giving your reasons
- make connections with the properties and behaviours of gases that you learned about in this unit (such as gases under pressure, gas stoichiometry)
- discuss the economic implications of imposing your plan on the industry
- discuss the environment costs and ethical issues surrounding your selected strategy

Communicate

Communicate your carbon strategy in a presentation format of your choice, such as a pamphlet, poster, formal written report, audiovisual presentation, or video.

Your presentation should contain

- an attention-grabbing introduction
- an overview of the information about your industry
- a clear description of the two alternative technologies and how they relate to the properties of gases, including the pros and cons of each technology
- your recommendation, with an explanation of your choice

You will present your plan to your classmates. Then, based on their suggestions and questions, you will make any necessary additions or changes to your presentation. Finally, send your plan to a company in the industry you researched and await their response!

GO TO NELSON SCIENCE

ASSESSMENT CHECKLIST

Your completed Unit Task will be assessed according to the following criteria:

Knowledge/Understanding

- ✓ Describe how a chosen industry contributes to increasing the carbon dioxide concentration in the atmosphere.
- ✓ Demonstrate an understanding of the properties of gases as applied to two carbon capture technologies.
- ✓ Describe two carbon capture technologies that could be appropriate for the chosen industry.

Thinking/Investigation

- ✓ Compare the pros and cons of two carbon capture technologies.

Communication

- ✓ Prepare and deliver an informative, clear, and engaging presentation that addresses all the specified points.

Application

- ✓ Recommend one of the technologies for the chosen industry. Provide a rationale for your recommendation.

For each question, select the best answer from the four alternatives.

1. Which of the following descriptions is characteristic only of a liquid? (11.1) K/U
   (a) It takes the shape and volume of its container.
   (b) It takes the shape of its container, but has a definite volume.
   (c) It has a definite shape and volume, unaffected by its container.
   (d) It flows readily.
2. Which of the following sources is the primary contributor to greenhouse gas emissions in Canada? (11.2) K/U
   (a) combustion of fossil fuels
   (b) eruptions of volcanoes
   (c) decay of plant matter in forests
   (d) decomposition of garbage in landfills
3. If mixed with pure nitrogen gas, which of the following gases would produce a mixture with a greater density than that of nitrogen alone? (11.3) K/U
   (a) hydrogen, H_2
   (b) helium, He
   (c) methane, CH_4
   (d) argon, Ar
4. Which of the following substances is a VOC? (11.4) K/U
   (a) ozone
   (b) gasoline
   (c) nitrogen monoxide
   (d) soot
5. After a snowstorm, you should check for obstructions in the exhaust vents of your furnace, wood stove, fireplace, or gas dryer to help avoid an indoor buildup of which of the following pollutants? (11.5) K/U
   (a) radon
   (b) carbon monoxide
   (c) methanal
   (d) mould spores
6. Which of the following substances may be found in car exhaust? (11.6) K/U
   (a) carbon monoxide
   (b) nitrogen oxides
   (c) VOCs
   (d) all of the above
7. Which of the following pressures is closest to 0.50 atm? (11.7) K/U
   (a) 380 Torr
   (b) 1530 mm Hg
   (c) 202.6 kPa
   (d) 103 000 N/m^2
8. If 40 cm^3 of a gas is warmed from 25 °C to 50 °C at constant pressure, the new volume would be approximately
   (a) 20 cm^3
   (b) 37 cm^3
   (c) 43 cm^3
   (d) 80 cm^3 (11.8) T/I
9. If 150 cm^3 of neon gas at 1.5 atm were allowed to expand to 300 cm^3 at constant temperature, what would be the sample's new pressure? (11.9) K/U
   (a) 0.38 atm
   (b) 0.75 atm
   (c) 3.0 atm
   (d) 6.0 atm
10. A cylinder of welding gas is outside on a hot day. The temperature of the gas increases from 20.0 °C to 35.0 °C but its volume does not change. If the original pressure was 13 000 kPa, what is the approximate value of the new pressure? (11.9) T/I
    (a) 7400 kPa
    (b) 12 400 kPa
    (c) 13 700 kPa
    (d) 22 800 kPa
11. 2.60 L of hydrogen gas is collected at 680 mm Hg and 25.0 °C. What is the predicted volume of the gas sample at STP? (11.9) T/I
    (a) 2.13 L
    (b) 2.54 L
    (c) 2.66 L
    (d) 2.86 L
12. Which of the following gas samples would contain the same amount of gas as 200 mL of helium, He(g), at 25 °C and 1 atm? (12.1) T/I
    (a) 400 mL of hydrogen, $H_2(g)$, at 25 °C and 1 atm
    (b) 40 mL of methane, $CH_4(g)$, at 25 °C and 1 atm
    (c) 200 mL of neon, Ne(g), at 25 °C and 1 atm
    (d) all of the above

13. Which of the following sets of conditions for a gas, when substituted into the ideal gas law equation, would give an answer numerically equal to the molar volume of an ideal gas at STP? (12.2) T/I
    (a) 1 atm, 1 mol, 273 K
    (b) 0.5 atm, 0.5 mol, 273 K
    (c) 2 atm, 2 mol, 273 K
    (d) all of the above
14. A mixture of hydrogen, $H_2(g)$, and ammonia, $NH_3(g)$, has a total pressure of 90 kPa. If the partial pressure of the ammonia is 30 kPa, the partial pressure of the hydrogen would be
    (a) 30 kPa
    (b) 60 kPa
    (c) 90 kPa
    (d) 120 kPa (12.4) T/I
15. What volume of oxygen is produced when 30 L of dinitrogen monoxide completely decomposes according to the chemical equation given below? (Assume that all gas volumes are measured at the same pressure and temperature.) (12.5) T/I
    $2\ N_2O(g) \rightarrow 2\ N_2(g) + O_2(g)$
    (a) 7.5 L
    (b) 15 L
    (c) 30 L
    (d) 45 L
16. What mass of carbon is needed to produce 25.0 L of ethene, C_2H_4, at 20.0 °C and 85.0 kPa according to the chemical equation given below? (12.5) T/I
    $2\ C(s) + 2\ H_2(g) \rightarrow C_2H_4(g)$
    (a) 20.9 g
    (b) 48.8 g
    (c) 154 g
    (d) 307 g

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

17. Temperature is a measure of the average kinetic energy of the entities in a sample. (11.1) K/U
18. In the last 200 years, increases in the atmospheric concentrations of greenhouse gases can be attributed mainly to climate change. (11.2) K/U
19. The discovery of oxygen forced Dmitri Mendeleev to add another column to his periodic table. (11.3) K/U
20. Ground-level ozone in photochemical smog is produced by the reaction of sulfur oxides and VOCs in the presence of sunlight. (11.4) K/U
21. Biological pollutants originate from organisms that tend to live in warm, dry places in homes. (11.5) K/U
22. Sealing cracks in exterior walls, especially around pipes and drains, helps to prevent radon from entering a home. (11.5) K/U
23. Idling cars tend to produce more air pollution than cars being driven at highway speeds. (11.6) K/U
24. Pressure is defined as force per unit volume. (11.7) K/U
25. Volume and temperature are directly proportional when the temperature is expressed in degrees Celsius. (11.8) K/U
26. In theory, the volume of a gas would shrink to zero at –273.15 K. (11.8) K/U
27. In an inverse proportion such as Boyle's law, when one variable increases, the other variable increases. (11.9) K/U
28. When Gay-Lussac's law is rearranged to solve for T_2, the result is
    $T_2 = \frac{P_2}{P_1}T_1$ (11.9) T/I
29. When the combined gas law is rearranged to solve for P_2, the result is
    $P_2 = \frac{P_1V_1}{T_2} \times \frac{T_1}{V_2}$ (11.9) T/I
30. When water is decomposed by electrical energy, the volume of hydrogen gas produced will equal the volume of oxygen gas produced. (Assume that the gases are measured at the same temperature and pressure.) (12.1) K/U
31. The collisions between ideal gas molecules are assumed to be perfectly elastic. (12.2) K/U
32. The molar volume of an ideal gas is 22.4 L at standard temperature and pressure. (12.2) K/U
33. Johannes van der Waals was the first scientist to liquefy hydrogen. (12.3) K/U
34. The partial pressure of a gas collected over water is equal to the atmospheric pressure plus the partial pressure of the water vapour mixed with the gas. (12.4) K/U
35. The chemical equation given below indicates that equal masses of hydrogen and fluorine will react to produce hydrogen fluoride. (Assume that all gases are measured at the same temperature and pressure.)
    $H_2(g) + F_2(g) \rightarrow 2\ HF(g)$ (12.5) T/I
36. A gas hydrate is formed when a methane molecule is trapped in a cage of water molecules. (12.6) K/U

Knowledge

For each question, select the best answer from the four alternatives.

1. Which of the following motions carries entities through space on a linear path? (11.1) K/U
   (a) vibrational
   (b) rotational
   (c) translational
   (d) all of the above
2. In which layer of the atmosphere does ozone absorb ultraviolet rays from the Sun that could be very harmful to living organisms if they reached Earth's surface? (11.2) K/U
   (a) thermosphere
   (b) mesosphere
   (c) troposphere
   (d) stratosphere
3. Riding a bicycle instead of driving in a car would help to fight climate change by
   (a) reducing greenhouse gas emissions
   (b) saving money
   (c) sequestering carbon
   (d) increasing nitrogen in the air (11.2) K/U
4. When attempting to determine the densities of hydrogen and oxygen, Lord Rayleigh had to develop special techniques for measuring
   (a) the mass of a gas sample
   (b) the volume of a gas sample
   (c) the temperature of a gas sample
   (d) the pressure of a gas sample (11.3) K/U
5. Which source is responsible for the majority of the carbon monoxide emissions in Ontario? (11.4) K/U
   (a) smelting operations
   (b) transportation
   (c) furnaces in buildings
   (d) controlled forest burns
6. Which is a symptom of carbon monoxide poisoning? (11.5) K/U
   (a) chest pain (c) dizziness
   (b) headache (d) all of the above
7. Where would a "no idling" sign be appropriate? (11.6) K/U A
   (a) at the drive-through window of a fast food restaurant
   (b) at the student drop-off and pick-up location of a school
   (c) at the waiting line for a car wash
   (d) all of the above
8. Which of the following gas laws represents an inverse proportion? (11.8, 11.9) K/U
   (a) Charles' law (c) Gay-Lussac's law
   (b) Boyle's law (d) all of the above
9. In an experiment, one volume of fluorine, $F_2(g)$, reacts with one volume of xenon, Xe(g), at STP. This observation indicates that the simplest chemical formula of the xenon–fluorine compound that forms is
   (a) XeF (c) Xe_2F
   (b) XeF_2 (d) Xe_2F_2 (12.1) T/I
10. Under which of the following sets of conditions would a real gas behave most like an ideal gas? (12.2) K/U
    (a) low pressure, low temperature
    (b) high pressure, high temperature
    (c) low pressure, high temperature
    (d) high pressure, low temperature
11. James Dewar and Heike Kamerlingh Onnes competed to be the first to
    (a) liquefy permanent gases
    (b) isolate a noble gas
    (c) prepare a Bose–Einstein condensate
    (d) identify the relationship between temperature and pressure (12.3) K/U
12. Of the scientists listed below, who first understood the concept of the partial pressure of a gas and developed a mathematical description of partial pressures? (12.4) K/U
    (a) Jacques Charles (c) John Dalton
    (b) Antoine Lavoisier (d) Amedeo Avogadro
13. A sample of dry air at 28 °C in a sealed flask has a pressure of 90 kPa. Which of the following values is most likely the partial pressure of oxygen, $O_2(g)$, in the flask? (12.4) T/I
    (a) 70 kPa (c) 3 Pa
    (b) 19 kPa (d) 0.8 kPa
14. Which of the following statements is a correct interpretation of the chemical equation shown below?
    $2\ CO(g) + O_2(g) \rightarrow 2\ CO_2(g)$ (12.5) K/U
    (a) 2 molecules of carbon monoxide react with 1 molecule of oxygen to form 2 molecules of carbon dioxide.
    (b) 2 mol of carbon monoxide reacts with 1 mol of oxygen to form 2 mol of carbon dioxide.
    (c) 2 L of carbon monoxide reacts with 1 L of oxygen to form 2 L of carbon dioxide, all gases at STP.
    (d) all of the above

Indicate whether each statement is true or false. If you think the statement is false, rewrite it to make it true.

15. When a liquid freezes, its molecules have only translational motion. (11.1) K/U
16. In order of increasing altitude, the layers of Earth's atmosphere are troposphere, stratosphere, mesosphere, and thermosphere. (11.2) K/U
17. Of the gases in the atmosphere, nitrogen, $N_2(g)$, is the most abundant by volume. (11.2) K/U
18. The density of a mixture of argon and nitrogen gases is less than the density of a sample of pure nitrogen gas at the same pressure. (11.3) K/U
19. Just prior to the 2008 Olympic Games in Beijing, China, the Chinese government spent a great deal of money planting trees, blocking traffic, and relocating factories to improve the architectural appeal of the city. (11.4) K/U
20. An AQHI of 7 or more indicates a high to very high degree of air pollution, so people should refrain from outdoor activities. (11.4) K/U
21. The dangerous gas that might seep into a home's basement, through cracked exterior walls, is methane. (11.5) K/U
22. Adults have a greater risk of respiratory irritation and health effects from air pollution because they breathe more rapidly and inhale more pollutants per kilogram of body mass. (11.6) K/U
23. A cubic metre of air at an altitude of 3 km would have a greater mass than the same volume at 300 km. (11.7) K/U
24. When two variables are directly proportional to each other, adding them yields a constant answer. (11.8) K/U
25. According to Avogadro's law, equal volumes of gas at the same temperature and pressure must consist of equal masses of entities. (12.1) K/U
26. A 1 mol sample of any ideal gas has a volume of 22.4 L at 101.325 kPa and 0 °C. (12.2) K/U
27. Hydrogen can be cooled sufficiently to liquefy it by suddenly increasing its pressure greatly. (12.3) K/U
28. When a gas is collected over water, the partial pressure of the gas is greater than the pressure of the atmosphere. (12.4) K/U
29. Gas hydrates form under low temperatures and low pressures. (12.6) K/U

Match each term on the left with the most appropriate description on the right.

30. (a) ideal gas law
    (b) combined gas law
    (c) Charles' law
    (d) Boyle's law
    (e) Gay-Lussac's law

    (i) constant temperature and amount of gas
    (ii) constant pressure and amount of gas
    (iii) pressure, volume, temperature, and amount of gas may all vary
    (iv) constant volume and amount of gas
    (v) constant amount of gas (11.8, 11.9, 12.2) K/U

Write a short answer to each question.

31. Give the state(s) of matter that apply in each case below. (11.1) K/U
    (a) The material takes the shape of the entire container.
    (b) The material cannot be compressed into a smaller volume even with great pressure.
    (c) The entities of the material can undergo vibrational, rotational, and translational motion.
    (d) The material retains its shape and volume independent of any container into which it is placed.
    (e) The entities have a high degree of order.
    (f) The material can easily be compressed into a smaller volume.
    (g) The material can be readily poured.
    (h) The entities of the material can only undergo vibrational motion.
    (i) The material takes the shape of the container but has a definite volume.
32. In what layer of the atmosphere is more than 99 % of the atmospheric water vapour found? (11.2) K/U
33. What gaseous air pollutant that contributes to the formation of acid rain is produced by the burning of fossil fuels containing sulfur impurities? (11.4) K/U
34. (a) What substances are needed for the formation of photochemical smog?
    (b) What other factor is required for the formation of photochemical smog? (11.4) K/U
35. Distinguish between chemical pollutants and biological pollutants that can be found indoors. (11.5) K/U
36. (a) What is high-altitude athletic training?
    (b) How effective is high-altitude training? (11.7) K/U
37. Why is there no such thing as a negative absolute temperature? (11.8) K/U
38. Why do car tires have a lower than usual air pressure on cold mornings? (11.9) K/U A

39. State the law of combining volumes and provide an example that illustrates this law. (12.1) K/U
40. At ordinary room conditions, butane is a gas. How does the liquid butane in a lighter demonstrate that butane is not an ideal gas? (12.2) K/U
41. How can lasers help to cool atoms? (12.3) K/U
42. What element did Kleppner and Greytak choose in their attempts to create a Bose–Einstein condensate? Why did they choose this element? (12.3) K/U
43. What two conditions must be met for mixed gases to illustrate the law of partial pressures? (12.4) K/U
44. (a) Explain why water and methane do not usually mix.
    (b) How do water and methane form a gas hydrate? (12.6) K/U
45. Give two reasons why the proposal to extract gas hydrates from the sea floor is controversial. (12.6) K/U

Understanding

46. Use the kinetic molecular theory to explain why gases mix quickly. (11.1) K/U
47. Greenhouse gases are both bad and good for Earth. Explain why this is the case. (11.2) K/U
48. What volume of oxygen, $O_2(g)$, could be isolated from 200 cm^3 of air? Show your calculations. (11.2) K/U T/I
49. Why would important documents be stored in sealed chambers, in an atmosphere of pure argon gas? (11.3) K/U
50. The concentration of ozone, $O_3(g)$, in smog tends to peak on sunny days near midday. Explain why, using relevant chemical equations in your answer. (11.4) K/U C
51. Why is it unwise to use a propane camp stove to heat a tent at night? (11.5) K/U A
52. Explain why an idling car or bus engine produces increased air pollution. (11.6) K/U A
53. If you drop a water balloon from a short height onto a single nail sticking up from a board, the balloon will break. However, if you drop the water balloon onto a bed of nails, the balloon survives. Explain how this is possible. (11.7) K/U
54. An average adult takes in 450 cm^3 of air with each breath. Suppose the temperature of the air prior to being inhaled is 22 °C and when in the lungs it warms to 37 °C. When the air is exhaled, what will be its volume? (11.8) T/I
55. Imagine that you have a 45 cm^3 sample of hydrogen gas at 30.0 °C. To what Celsius temperature must this sample be cooled to reduce its volume to 35 cm^3? Assume constant pressure. (11.8) T/I
56. The fabric envelope of a hot air balloon contains 3.00×10^3 m^3 of air at 95 °C. If the air in the balloon were to cool to 20 °C, what volume would it occupy? (11.8) T/I
57. A 30.0 cm^3 sample of hydrogen gas is collected at 18 °C and 90.7 kPa. Calculate the volume the gas would have at a pressure of 96.0 kPa and a temperature of 28 °C. (11.9) T/I
58. Two students have different ideas about gas pressure. One student says, "Equal volumes of two gases with very different molar masses exert the same pressure at the same temperature." The other student questions, "Wouldn't the gas with the greater molar mass exert more pressure because its entities would bounce off the walls of the container with more force?" Which student is correct? Explain your thinking. (12.1) K/U T/I
59. An aqueous solution of ammonia, $NH_3(aq)$, can be broken down by electrolysis into pure nitrogen gas, $N_2(g)$, and pure hydrogen gas, $H_2(g)$. (12.1) K/U C
    (a) Write the balanced chemical equation for the decomposition of ammonia.
    (b) Predict the ratio of the volumes of nitrogen gas to hydrogen gas produced, assuming the same temperature and pressure for each gas. Explain your prediction.
60. What volume would 10.0 g of neon, Ne(g), occupy at STP? (12.1) T/I
61. What volume, in litres, would 8.40 g of sulfur trioxide gas, $SO_3(g)$, occupy at a pressure of 90.0 kPa and a temperature of 27 °C? (12.2) T/I
62. Calculate the density, in g/L, of pure ammonia, $NH_3(g)$, at 90.2 kPa and 30 °C. (12.2) T/I
63. How does the competition to liquefy helium illustrate that the progress of science can be greatly affected by human nature and human shortcomings? (12.3) K/U
64. A gaseous mixture of carbon dioxide, carbon monoxide, and nitrogen is held in a container with a total pressure of 100 kPa. If the partial pressures of carbon dioxide and carbon monoxide are equal and the partial pressure of nitrogen is 40 kPa, what is the partial pressure of each of the other two gases? Show calculations to support your answer. (12.4) K/U T/I
65. A tank contains a mixture of methane, ethane, and propane gases. The partial pressures of the gases are, respectively, 80 kPa, 40 kPa, and 30 kPa. (12.4) K/U T/I
    (a) Determine the total pressure in the tank.
    (b) Some additional methane gas is injected into the tank and the total pressure becomes 200 kPa. What are the partial pressures of each of the gases inside the tank after the additional methane is injected? Explain your answer.
66. A sample of oxygen gas is collected over water at 26 °C. The total pressure is 85 kPa. What is the pressure of the oxygen alone? (12.4) T/I
67. Why is it important for airbags to have a cover that is permeable to gases? (12.5) K/U A

68. A 100 g sample of potassium chlorate, $KClO_3(s)$, is completely decomposed by heating:

$$2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$$

The oxygen is collected at 89 kPa and 23 °C. Calculate the volume, in litres, of oxygen that would be produced during the decomposition of this sample. (12.5) T/I

69. Calculate the volume of ammonia, $NH_3(g)$, that would be produced when 60 L of hydrogen reacts completely with excess nitrogen:

$$3\ H_2(g) + N_2(g) \rightarrow 2\ NH_3(g)$$

Assume that all gases are measured at 22 °C and 100 kPa. (12.5) T/I

Analysis and Application

70. When potassium permanganate, $KMnO_4(s)$, dissolves in water it produces a violet solution. (11.1) K/U T/I
    (a) Predict how the spread of the colour would compare when a single crystal of potassium permanganate is dropped into 50 mL of very cold water and another crystal is dropped into 50 mL of very hot water.
    (b) Use the KMT to explain your prediction.
71. A student is using a microscope to observe the Brownian motion of tiny pollen grains in a drop of water on a slide. The microscope's light gradually warms the water. (11.1) K/U T/I
    (a) Predict what will happen to the motion of the pollen grains as the temperature of the water rises.
    (b) Explain your prediction in terms of the kinetic molecular theory.
72. Carbon dioxide levels in Earth's atmosphere have fluctuated for millennia. Why, then, do scientists believe that the increase in atmospheric carbon dioxide over the past 200 years is the result of human activities, not just a natural fluctuation? (11.2) K/U T/I
73. Dmitri Mendeleev did not add the noble gases to his periodic table until 1902, quite some time after many of the other elements in the periodic table had been found. Suggest two reasons to explain why it took chemists so long to discover the noble gases. (11.2, 11.3) T/I
74. A family buys a home that has recently been renovated. Newly varnished hardwood floors, freshly painted walls, and new window coverings have greatly improved the appearance of the living space. (11.5) K/U T/I A
    (a) What advice would you give to the family with regard to potential health risks when they take occupancy of the home? Explain your thinking.
    (b) Would your advice include the basement even though no renovations have been done there? Why or why not?
    (c) It is possible that the new paint has been applied over mould. Why would this be a problem and what would you advise the new residents?
75. You put 100 mL of hot water into a 2 L pop bottle and cap it tightly. Then you put the bottle into the refrigerator. Predict what will happen and explain your prediction. (11.1, 11.7) K/U T/I A
76. Use the kinetic molecular theory to explain how the air inside an inflated tire can support the weight of the car. (11.1, 11.7) K/U T/I A
77. A scientist records volume versus temperature data for an investigation involving three noble gases (**Table 1**).

Table 1 Molar Masses of Three Gases

Gas	*M* (g/mol)
neon	20.18
argon	39.95
krypton	83.80

The same amount of each gas was used and the pressure was held constant during the measurements. However, the value at which the pressure was maintained was different for each gas. The scientist's data are plotted in **Figure 1**. (11.8, 11.9) K/U T/I A

   (a) Based on Figure 1, which gas was at the lowest constant pressure and which was at the highest constant pressure? Explain your answers.
   (b) Assuming that experimental errors were insignificant, predict how the value for absolute zero obtained by extrapolation would compare for each set of data. Explain your prediction.
   (c) For each gas, would a in $V/T = a$ have the same value? Give evidence to support your answer.

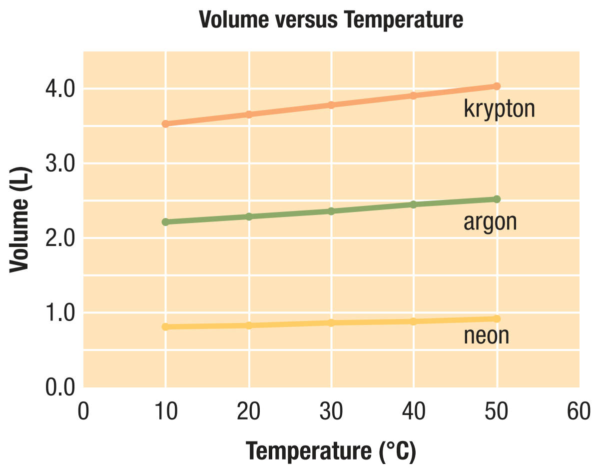


Figure 1

78. An unopened can of pop is never completely filled with liquid. There is a small volume of space that contains carbon dioxide gas at a pressure of several hundred kilopascals. The pressure of the carbon dioxide gas keeps the drink properly carbonated. (11.1, 11.9) T/I A
    (a) Does the pressure of the gas in the pop can compress the liquid into a smaller volume than the value stated on the label? Why or why not?
    (b) A certain brand of pop is sold in cans with 15 cm^3 of air space filled with carbon dioxide gas. The desired pressure in the air space is 405 kPa to 425 kPa. If 90 cm^3 of carbon dioxide at standard pressure is injected into each can during sealing at the factory, will the gas pressure in the air space be at the correct value? Assume that the temperature and amount of gas remain constant. Show calculations to support your answer.

79. An automobile racing team is trying to determine which tires would be best for their car. They are interested in one model that has an interior volume of 25 L. The specifications state that before the car is driven, the tires should be inflated to a pressure of 220 kPa, assuming an outside air temperature of 22 °C. Out on the track, though, the tires typically warm up to 90 °C, which causes the tire pressure to rise. If the pressure exceeds 290 kPa, these particular tires become too rigid and the car handles poorly. Is this model of tire a good choice? Support your answer with calculations. Assume that the amount and volume of gas remain constant. (11.9) T/I A

80. Many scientific supply companies sell self-contained lab burners that have a barrel and valve attached to a pressurized can of butane fuel that is about the size of a typical aerosol can. A full can burns for 5.5 hours while fuelling a typical laboratory flame. Suppose you are a teacher and you need to know how much burning time is left in a set of partly used burners before having a class do a 15-minute investigation with them. Describe mass measurements you could make on a used burner that would provide you with the burning time remaining for this type of burner. Assume that new fuel cans are all filled with the same amount of butane and that the empty burners have the same mass. (12.1) K/U T/I

81. An inventor is perfecting a machine that will rapidly pump precisely measured, pre-set masses of nitrogen, $N_2(g)$, into inflatable devices. In a test of the machine, the inventor wants to inflate a 50.0 L airbag to SATP conditions. Calculate the mass of nitrogen that the machine should deliver into the airbag. (12.1) T/I

82. **Figure 2** shows 10 flasks containing various volumes and masses of gases. The gases may be the same substance, or different. (12.1) T/I A
    (a) Identify which flasks contain equal numbers of gaseous entities.
    (b) Explain the reasoning that led you to your answer for Part (a).
    (c) What is the ratio of the mass of an entity of gas in vessel C to the mass of an entity of gas in vessel A? Explain your answer.
    (d) What is the mass ratio of an entity of gas in vessel A to an entity of gas in vessel B? Explain your answer.

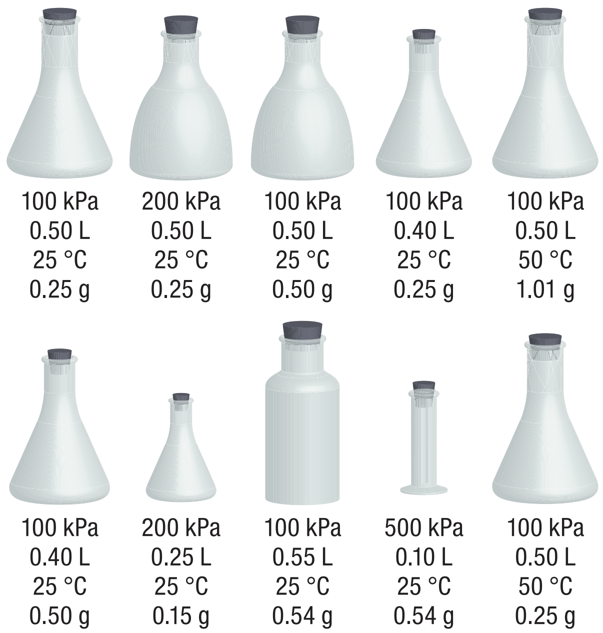


Figure 2

83. Under the same conditions of temperature and pressure, equal volumes of fluorine, $F_2(g)$, and an unknown gas have respective masses of 0.76 g and 0.34 g. Calculate the molar mass of the unknown gas. (12.1) T/I

84. An analytical chemist believes that an unknown gas is dinitrogen pentoxide, $N_2O_5(g)$. As a confirmatory test, she determines the density of the gas to be 3.46 g/L at 80.0 kPa and 27.0 °C. (12.2) T/I
    (a) Calculate the molar mass of the unknown gas from the data given.
    (b) Does your result from Part (a) support the chemist's conclusion that the unknown gas is N_2O_5? Explain why or why not.

85. Natural gas is mostly methane, $CH_4(g)$, and ethane, $C_2H_6(g)$. The average molar mass of a natural gas mixture indicates the approximate concentrations of methane and ethane. The density of a natural gas sample is found to be 0.714 g/L at 93.2 kPa and 22 °C. Follow the steps below to find the methane and ethane concentrations in the gas. (12.2) T/I
   (a) Calculate the average molar mass of the natural gas sample using the given density.
   (b) Using atomic masses, calculate the molar masses of pure methane and pure ethane.
   (c) The average molar mass of a mixture of two gases, A and B, is given by the equation
   $M_{avg} = xM_A + (1 - x)M_B$
   in which x is the decimal fraction (by volume) of the mixture that is gas A. Use your answers to Parts (a) and (b) to find x.
   (d) Use your value of x from Part (c) to determine the percentage by volume of methane and ethane in the natural gas.

86. A small business has obtained an inflatable balloon for advertising. The nearest supplier of helium only has 60.0 kg available. If the balloon has a volume of 283 000 L, is 60.0 kg enough helium to inflate the balloon to a desired pressure of 100 kPa at 18 °C? Support your answer. (12.2) T/I

87. Real gases do not behave exactly like ideal gases. In which case might we expect a greater deviation of the sum of the partial pressures from the actual total pressure: a mixture of two gases composed of polar molecules or a mixture of two gases composed of non-polar molecules? Explain. (12.4) K/U T/I

88. Three gases are mixed in a 2.0 L vessel. The vessel is connected by a valve to a 4.0 L chamber. This second chamber contains no gas at all; it is a vacuum. The three gases in the 2.0 L vessel—helium, neon, and argon—have the following partial pressures: 45.0 kPa, 20.0 kPa, and 30.0 kPa. (12.4) K/U T/I
   (a) What is the total pressure in the 2 L chamber?
   (b) When the valve between the chambers is opened, what will the partial pressure of each gas become when the gases have evenly filled both chambers, assuming no temperature change?
   (c) What is the total pressure in the two-chamber system after the valve is opened?

89. 1.0 mol of nitrogen, $N_2(g)$, 0.20 mol of argon, $Ar(g)$, and 1.2 mol of oxygen, $O_2(g)$, are pumped into a 10.0 L tank at a temperature of 22 °C. (12.2, 12.4) T/I
   (a) Use the ideal gas law to determine the partial pressure, in kPa, of each gas in the tank.
   (b) What is the total pressure in the tank?

90. A scientist is calibrating equipment that will measure the energy released by chemical reactions. The equipment consists of a steel chamber surrounded by water held in an insulated jacket (**Figure 3**). The steel chamber has a volume of 0.100 L and contains an electrical device for initiating reactions. The scientist plans to detonate 1.00 g of sodium azide, $NaN_3(s)$, in the chamber. He knows that the explosion will instantly warm the gas in the chamber by 585 °C. The chamber can withstand 1720 kPa of pressure and is initially filled with air at 101.3 kPa of pressure and 25 °C. Work through the following questions to determine whether detonating 1.00 g of sodium azide will result in too much pressure and damage the equipment. (12.2, 12.4, 12.5) T/I

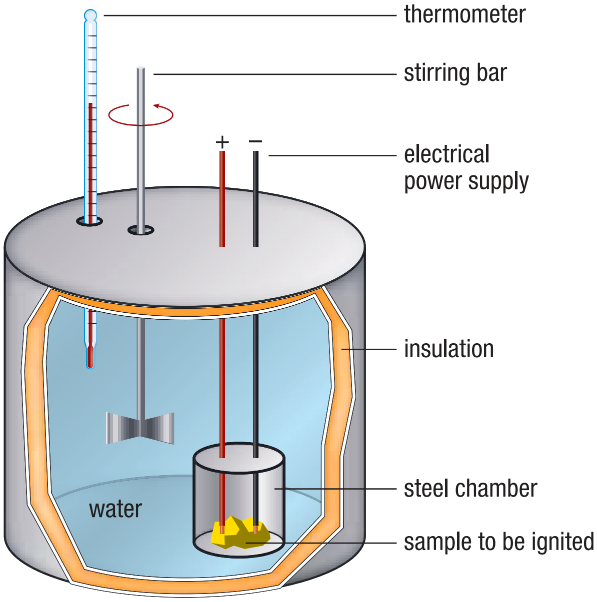


Figure 3

   (a) Calculate the amount of air in the chamber at the start.
   (b) Sodium azide decomposes explosively:
   $2\ NaN_3(s) \rightarrow 2\ Na(s) + 3\ N_2(g) + \text{energy}$
   Calculate the amount of nitrogen that is produced when 1.00 g of sodium azide decomposes.
   (c) Calculate the total amount of gas in the chamber after the sodium azide is detonated.
   (d) Calculate the gaseous pressure inside the chamber at the temperature produced by the detonation of the sodium azide.
   (e) Will the equipment be damaged by detonating 1.00 g of sodium azide in the chamber? Explain.

91. A chemistry teacher has asked her students to determine whether a granulated, shiny, silvery metal is zinc or aluminum. The teacher provides a procedure whereby a known mass of the unknown metal is reacted with hydrochloric acid to produce a dissolved ionic compound and hydrogen gas. The students collect the gas at a pressure of 92 kPa and a temperature of 21 °C. One set of data is given in **Table 2**. (12.2, 12.4, 12.5) K/U C A

Table 2 Reaction of a Metal with Hydrochloric Acid

Mass of metal reacted (g)	Volume of $H_2(g)$ produced (mL)
0.25	104

(a) Write balanced chemical equations for the reactions of zinc with hydrochloric acid and aluminum with hydrochloric acid.
(b) Based on the mass of the metal reacted, use your equations from Part (a) and the ideal gas law to predict the volume of hydrogen gas that each reaction would produce at the pressure and temperature of the experiment.
(c) Based on your calculations, what can you conclude about the identity of the metal?
(d) The experimental volume of hydrogen collected is a little larger than the theoretical volume. A student suspects that it may be related to the fact that the hydrogen was collected over water. Explain this student's hypothesis and outline how the student may correct for the discrepancy.

Evaluation

92. Scientists predict that climate change will cause temperate vegetation to spread 160–640 km north into Canada's boreal, or northern, forest. A friend says that a warmer climate would be a good thing. It would make life much more comfortable and convenient for people living in the boreal forest. Growing seasons will be longer, so more crops can be raised and less fossil fuel will be used to heat homes. Your friend thinks that climate change should be welcomed. Do you agree? Cite specific examples to defend your position. (11.2) K/U T/I A

93. A study of methane emissions from mid-boreal forests in northern Ontario showed that beaver ponds are significant sources of methane, a greenhouse gas. What would you have to know to assess whether or not methane emissions from beaver ponds are a major contributor to climate change? Be specific about the information you would need and calculations to be performed. (11.2) K/U T/I A

94. A tree sequesters carbon as it grows. When it dies and decays, however, its carbon is recycled to the environment as carbon dioxide, $CO_2(g)$, and methane, $CH_4(g)$. (11.2) K/U T/I A
(a) What biological process involving sunlight allows trees to sequester carbon? Explain.
(b) How does the sequestration of carbon through the process you described in Part (a) help to fight climate change?
(c) Explain why trees harvested for lumber do not recycle carbon to the environment immediately.
(d) Canada is the world's largest exporter of forest products. The Canadian forest industry emphasizes environmentally sound forest management, and many Canadians find jobs working for companies that specialize in planting new trees to replace harvested trees. What advantages would young trees have over mature trees with regard to sequestering carbon?
(e) Is growing and harvesting trees for lumber a perfect way to sequester carbon? Why or why not?

95. Two new homeowners are debating what kind of carbon monoxide detectors to get for their two-storey home to protect themselves and their children. They can afford two plug-in detectors or four battery-powered detectors. One partner wants to buy the two plug-in detectors to avoid having to deal with the expense and inconvenience of battery replacement. He wants to place one detector next to the gas-fired furnace and one next to the gas-fired water heater, both of which are in the basement. The other partner wants to buy the battery-powered detectors to be able to put additional detectors in the hall just outside the master bedroom and the children's bedroom, both of which are upstairs. Which plan do you believe is best? Do you have other suggestions for a better plan overall? Explain your reasoning. (11.5) K/U T/I A

96. A student performs an experiment that uses an electronic probe to measure the pressure of a sample of air sealed in a rubber balloon at various temperatures. The student holds the bottom of the balloon in trays of water at different temperatures to change the temperature of the air sample. The student's goal is to obtain an experimental value for absolute zero by extrapolating a plot of the pressure versus volume data. Another student performs a similar experiment with the same goal, except that the pressure probe is attached to a sealed test tube that is entirely immersed in water baths at various temperatures. Whose method is more trustworthy and likely to yield an accurate result? Explain why you have reached this conclusion, describing significant sources of error in the poorer method that are less likely to occur in the better method. (11.8, 11.9) K/U T/I

97. A meteorologist is assembling a weather balloon. The balloon's envelope has a volume of 60 000 L and a mass of 4.0 kg. The balloon must lift 20.0 kg of instruments in addition to its own mass. When properly inflated, the balloon must have a pressure of 90 kPa at a temperature of 15 °C. The mass that the balloon can lift is the mass of the air it displaces minus the mass of the gas inside its envelope. The meteorologist wants to use methane, $CH_4(g)$, to fill the balloon. His colleague, however, believes that helium, despite its greater cost, is the only gas that will work. Assuming that the average molar mass of air is 28.90 g/mol, which gas is better: methane or helium? Show calculations to support your answer. (12.2) K/U T/I A

98. A student plans to measure the density of carbon dioxide gas at ambient conditions. First the student will determine the mass of an effervescent tablet, a flask containing 100 mL of water, and a rubber stopper fitted with rubber tubing. Then the student will place the tablet into the water, quickly stopper the flask, and hold the tubing under a water-filled, inverted graduated cylinder clamped so that its opening is under water. As the tablet dissolves, it will produce carbon dioxide gas that will collect in the inverted graduated cylinder. When the bubbling has stopped, the student will read the volume of the carbon dioxide and then find the mass of the stopper, tubing, flask, and contents of the flask. Subtracting this mass from the mass of everything before the reaction will give the mass of the carbon dioxide. With the mass and volume of the carbon dioxide sample in hand, the student will find the density. Evaluate the student's procedure. (12.4) T/I

99. Canada, working collaboratively with NASA, was one of the first nations to become involved in launching Earth-orbiting satellites. Launching satellites requires rockets. A good rocket fuel, among other things, must generate high pressure per mass within the combustion chamber as it burns. Pressures are higher when the combustion reaction produces greater amounts of gaseous products. Chemical equations for three common rocket fuel systems are given below.

Fuel System I: Aluminum and ammonium perchlorate

$$10\ Al(s) + 6\ NH_4ClO_4(s) \rightarrow 4\ Al_2O_3(s) + 2\ AlCl_3(s) + 12\ H_2O(g) + 3\ N_2(g) + \text{energy}$$

Fuel System II: Hydrogen combustion

$$2\ H_2(l) + O_2(l) \rightarrow 2\ H_2O(g) + \text{energy}$$

Fuel System III: Hydrazine decomposition

$$N_2H_4(l) \rightarrow N_2(g) + 2\ H_2(g) + \text{energy}$$

Which fuel system is best with respect to the amount of gas produced per kilogram of fuel? Support your answer with calculations. (12.5) K/U T/I

Reflect on Your Learning

100. Describe something in your everyday life that illustrates any one of the gas laws that you learned about in this unit. K/U A

101. State and explain something that you learned in this unit about the atmosphere that you had not known before. K/U

102. What topic from this unit would you like to know more about? Explain why and state a plan for finding more information. K/U A

Research

GO TO NELSON SCIENCE

103. The planet Mercury has no atmosphere, whereas Venus and Earth have relatively thick atmospheres. What allows a planet to keep an atmosphere? Find the answer to this question through research. Report your findings, incorporating concepts of the kinetic molecular theory into your report. K/U A C

104. Dr. Donald Stedman of the University of Denver, Colorado, pioneered the development of remote sensing equipment that can measure the pollutants in car exhaust as the cars drive by. Immediately after it was developed, the sensing system allowed researchers to acquire data about the pollution coming from cars under actual driving conditions. Research the history, design, operation, and findings of Stedman's remote sensing system. Report what you learn, including diagrams of the components and roadside operation of the sensing system. T/I C A

105. There are several types of barometers: devices that measure atmospheric pressure. Research how mercury barometers, aneroid barometers, and water barometers work. Prepare a poster that has a diagram of each type of barometer with accompanying text that explains how it works. Also research how barometric readings can be used to predict the weather and include a summary of this information on your poster. K/U C A

106. Avogadro's hypothesis, which was proposed in 1811, resolved a great deal of confusion about atomic masses in the nineteenth century, much of which stemmed from incorrect assumptions by John Dalton. However, it took until 1860, four years after Avogadro's death, for the scientific community to realize that Avogadro's ideas were the key to the puzzle. Research the history of the determination of atomic masses (weights) in the nineteenth century, focusing on the ideas of Avogadro, Dalton, and Gay-Lussac. In particular, find out what role Stanislao Cannizzaro played. Present your findings in a brief written report. K/U C

Appendices

CONTENTS

SKILLS HANDBOOK LINKS

Throughout this textbook you will see links to the Skills Handbook, where appropriate, in each activity. These links identify supporting material in the Skills Handbook that will help you with that activity.

A1 Safety

A1.1 Safety Conventions and Symbols

Although we make every effort to make the science experience a safe one, there are some inherent risks. These risks are generally associated with the materials and equipment used, and with disregard of safety instructions when conducting investigations. Most of these risks pose no more danger than we normally experience in everyday life. We can reduce these risks by doing the following: being aware of the possible hazards, knowing the rules, behaving appropriately, and using common sense.

Remember, you share the responsibility not only for your own safety, but also for the safety of those around you (**Figure 1**). Always alert your teacher in case of an accident. In this textbook, chemicals, equipment, and procedures that are hazardous are indicated by the appropriate Workplace Hazardous Materials Information System (WHMIS) symbol or by .

Figure 1 Taking the proper safety precautions will ensure a safe environment for you and your classmates.

WHMIS SYMBOLS AND HHPS

The Workplace Hazardous Materials Information System (WHMIS) provides workers and students with complete and accurate information about hazardous products. All chemical products supplied to schools, businesses, and industries must contain standardized labels and be accompanied by a Material Safety Data Sheet (MSDS). The MSDS provides detailed information about the product. Clear and standardized symbols are an important component of WHMIS (**Table 1**, p. 630). These symbols must be present on the product's original container and shown on other containers if the product is transferred.

The *Canadian Hazardous Products Act* requires manufacturers of consumer products containing chemicals to include a symbol that specifies the nature of the hazard and whether it is the contents or the container that is dangerous. In addition, the label must state any secondary hazards, first-aid treatment, storage, and disposal. Household Hazardous Product Symbols (HHPS) are used to show the type of hazard. The shape of the frame around the symbol indicates whether the hazard is due to the contents or the container (**Figure 2**).

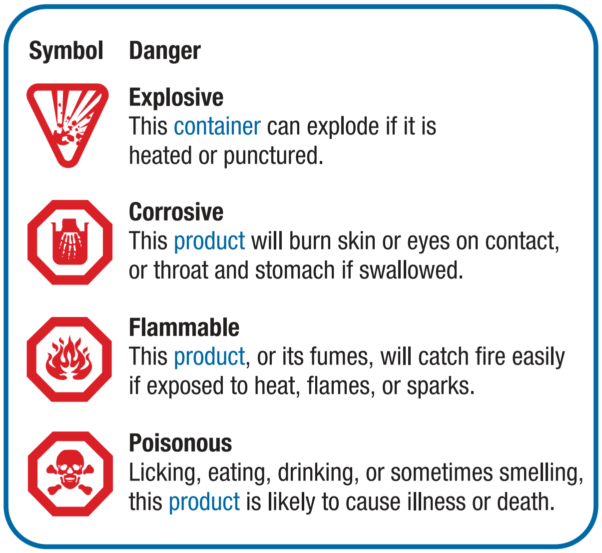


Figure 2 Household Hazardous Products Symbols (HHPS) appear on many products. A triangular frame indicates that the container is potentially dangerous. An octagonal frame indicates that the contents pose a hazard.

Table 1 The Workplace Hazardous Materials Information System (WHMIS)

Class and type of compounds	WHMIS symbol	Risks	Precautions
Class A: *Compressed Gas* Material that is normally gaseous and kept in a pressurized container		• could explode due to pressure • could explode if heated or dropped • possible hazard from both the force of explosion and the release of contents	• ensure container is always secured • store in designated areas • do not drop or allow to fall
Class B: *Flammable and Combustible Materials* Materials that will continue to burn after being exposed to a flame or other ignition source		• may ignite spontaneously • may release flammable products if allowed to degrade or when exposed to water	• store in designated areas • work in well-ventilated areas • avoid heating • avoid sparks and flames • ensure that electrical sources are safe
Class C: *Oxidizing Materials* Materials that can cause other materials to burn or support combustion		• can cause skin or eye burns • increase fire and explosion hazards • may cause combustibles to explode or react violently	• store away from combustibles • wear body, hand, face, and eye protection • store in container that will not rust or oxidize
Class D: *Toxic Materials* Poisons and potentially fatal materials that cause immediate and severe harm		• may be fatal if ingested or inhaled • may be absorbed through the skin • small volumes have a toxic effect	• avoid breathing dust or vapours • avoid contact with skin or eyes • wear protective clothing, and face and eye protection • work in well-ventilated areas and wear breathing protection
Class D: *Toxic Materials* Materials that have a harmful effect after repeated exposures or over a long period		• may cause death or permanent injury • may cause birth defects or sterility • may cause cancer • may be sensitizers causing allergies	• wear appropriate personal protection • work in well-ventilated areas • store in appropriate designated areas • avoid direct contact • use hand, body, face, and eye protection • ensure respiratory and body protection is appropriate for the specific hazard
Class D: *Biohazardous Infectious Materials* Infectious agents or a biological toxin causing a serious disease or death		• may cause anaphylactic shock • includes viruses, yeasts, moulds, bacteria, and parasites that affect humans • includes fluids containing toxic products • includes cellular components	• special training is required to handle materials • work in designated biological areas with appropriate engineering controls • avoid forming aerosols • avoid breathing vapours • avoid contamination of people and/or area • store in special designated areas
Class E: *Corrosive Materials* Materials that react with metals and living tissue		• eye and skin irritation on exposure • severe burns/tissue damage on longer exposure • lung damage if inhaled • may cause blindness if contacts eyes • environmental damage from fumes	• wear body, hand, face, and eye protection • use breathing apparatus • ensure protective equipment is appropriate • work in well-ventilated areas • avoid all direct body contact • use appropriate storage containers and ensure nonventing closures
Class F: *Dangerously Reactive Materials* Materials that may have unexpected reactions		• may react with water • may be chemically unstable • may explode if exposed to shock or heat • may release toxic or flammable vapours • may vigorously polymerize • may burn unexpectedly	• handle with care avoiding vibration, shocks, and sudden temperature changes • store in appropriate containers • ensure storage containers are sealed • store and work in designated areas

A1.2 Safety in the Laboratory

Safety in the laboratory is an attitude and a habit more than it is a set of rules. It is easier to prevent accidents than to deal with the consequences of an accident. Most of the following rules are common sense:

- Do not enter a laboratory unless a teacher or other supervisor is present, or you have permission to do so.
- Know your school's safety regulations.
- Tell your teacher about any allergies or medical problems you may have.
- Wear eye protection, a lab apron, and safety gloves when instructed by your teacher. Wear closed shoes.
- Tie back long hair and wear a protective lab coat over loose clothing. Remove any loose jewellery and finger rings.
- Keep yourself and your work area tidy and clean. Keep aisles clear.
- Never eat, drink, or chew gum in the laboratory.
- Do not taste any substance in a laboratory.
- Know the location of MSDS information, exits, and all safety equipment, such as the first-aid kit, fire blanket, fire extinguisher, and eyewash station, and be familiar with their contents and operation.
- Avoid moving suddenly or rapidly in the laboratory, especially near chemicals or sharp instruments.
- If you are not sure what to do, ask your teacher for directions.
- Never change anything, or start an activity or investigation on your own, without your teacher's approval.
- Before you start an investigation that you have designed yourself, get your teacher's approval.
- Never attempt unauthorized experiments.
- Never work in a crowded area or alone in the laboratory.
- Always stand up when doing laboratory practical work. Do not sit down.
- Use stands, clamps, and holders to secure any potentially dangerous or fragile equipment that could be tipped over.
- Never smell chemicals unless specifically instructed to do so by your teacher. Do not inhale the vapours, or gas, directly from the container. Take a deep breath to fill your lungs with air, then waft or fan the vapours toward your nose.
- Report all accidents.
- Inform your teacher of any spills and follow your teacher's instructions on how to clean up the spill. Clean up all spills, even water spills, immediately.
- Wash your hands with soap and warm water when you finish an investigation, and before you leave the laboratory.
- Remember safety procedures when you leave the laboratory. Accidents can also occur outdoors, at home, and at work.

EYE, EAR, AND FACE SAFETY

- Always wear approved eye protection in a laboratory. Keep the safety glasses or goggles over your eyes, not on top of your head. For certain experiments, full face protection may be necessary.
- If you must wear contact lenses in the laboratory, be extra careful; whether or not you wear contact lenses, do not touch your eyes without first washing your hands. If you do wear contact lenses, make sure that your teacher is aware of it. Carry your lens case and a pair of glasses with you.
- If you wear prescription eyeglasses, you must still wear the appropriate eye protection on top of them.
- Do not stare directly at any bright source of light (for example, burning magnesium ribbon). You will not feel any pain if your retina is being damaged by intense radiation. You cannot rely on the sensation of pain to protect you.
- If a piece of glass or other foreign object enters your eye, seek immediate medical attention.

HANDLING GLASSWARE SAFELY

- Never use glassware that is broken, cracked, or chipped. Give such glassware to your teacher or dispose of it as directed. Do not put the item back into circulation.
- Never pick up broken glassware with your fingers. Use a broom and dustpan.
- Dispose of glass fragments in special containers marked "Broken Glass."
- Check with your teacher before heating any glassware. Heat glassware only if it is approved for heating.
- Be very careful when cleaning glassware. There is an increased risk of breakage from dropping when the glassware is wet and slippery.
- If you cut yourself, inform your teacher immediately and get appropriate first aid. Embedded glass or continued bleeding requires medical attention.

USING SHARP INSTRUMENTS SAFELY

- Make sure that your instruments are sharp. Dull cutting instruments require more pressure than sharp instruments and are, therefore, much more likely to slip.
- Select the appropriate instrument for the task. Never use a knife when scissors would work best.
- Always cut away from yourself and others.
- If you cut yourself, inform your teacher immediately and get appropriate first aid.

HEAT SAFETY

- Make sure that heating equipment, such as the burner, hot plate, or electric heater, is secure on the bench and clamped in place when necessary.
- Do not use a laboratory burner near wooden shelves, flammable liquids, or any other item that is combustible.
- Do not allow overheating if you are performing an experiment in a closed area.
- Always assume that hot plates and electric heaters are hot and use protective gloves when handling.
- In a laboratory where burners or hot plates are being used, never pick up a glass or metal object without first checking the temperature by placing your hand near but not touching it. It may be hot enough to burn you.
- Do not touch a light source that has been on for some time. It may be hot and cause burns.
- If you burn yourself, *immediately* run cold water gently over the burned area or immerse the burned area in cold water and inform your teacher.
- Never look down the barrel of a laboratory burner.
- Always pick up a burner by its base, never by its barrel.
- Never leave a lighted burner unattended.
- Any metal powder can be explosive. Do not put these in a flame.
- When heating a test tube over a laboratory burner, use a test-tube holder or a utility clamp. Holding the test tube at an angle, with the open end pointed away from you and others, gently move the test tube back and forth through the flame.
- To heat a beaker, put it on the hot plate and secure with a ring support attached to a utility stand. (Placing a wire gauze under the beaker is optional.)
- Remember to include a cooling time in your experiment plan; do not put away hot equipment.

FIRE SAFETY

To use a burner:

- Tie back long hair and tie back or roll up long sleeves or other loose clothing.
- Secure the burner to a stand using a metal clamp.
- Check that the rubber hose is properly connected to the gas valve.
- Close the air vents on the burner. Use a sparker to light the burner.
- Open the air vents just enough to get a blue flame.
- Control the size of the flame using the gas valve.
- Immediately inform your teacher of any fires. A very small fire in a container may be extinguished by covering the container with a wet paper towel or a ceramic square to cut off the supply of air. Alternatively, sand may be used to smother small fires. A bucket of sand with a scoop should be available in the laboratory.
- If anyone's clothes or hair catch fire, tell the person to drop to the floor and roll. Then use a fire blanket to smother the flames. Never wrap the blanket around a standing person on fire.
- For larger fires, immediately evacuate the area. Call the office or sound the fire alarm. Do not try to extinguish larger fires. As you leave the classroom, make sure that the windows and doors are closed.
- If you use a fire extinguisher, direct the extinguisher at the base of the fire and use a sweeping motion, moving the extinguisher nozzle back and forth across the front of the fire's base.
- Different extinguishers are effective for different classes of fires. The fire classes are outlined in **Table 2**. Fire extinguishers in the laboratory are 2A10BC. They extinguish class A, B, and C fires.

Table 2 Classes of Fires

Class	Description	Firefighting strategy
Class A	• ordinary combustible materials that leave coals or ashes (e.g., wood, paper, cloth)	• water or dry chemical extinguishers
Class B	• flammable liquids (e.g., gasoline or solvents)	• carbon dioxide or dry chemical extinguishers
Class C	• live electrical equipment (e.g., appliances, photocopiers, computers, laboratory equipment)	• carbon dioxide or dry chemical extinguishers • do not use water
Class D	• burning metals (e.g., sodium, potassium, magnesium, aluminum)	• sand, salt, or graphite • do not use water
Class E	• radioactive substances	• requires special consideration at each site

ELECTRICAL SAFETY

- Do not operate electrical equipment near running water or a large container of water. Water or wet hands should never be near electrical equipment such as a hot plate, a light source, or a microscope.
- Check the condition of electrical equipment. Do not use it if wires or plugs are damaged, or if the ground pin has been removed.
- Make sure that electrical cords are not placed where someone could trip over them.
- When unplugging equipment, remove the plug gently from the socket. Do not pull on the cord.

WASTE DISPOSAL

Waste disposal at school, at home, and at work is a societal issue. Most laboratory waste can be washed down the drain or, if it is in solid form, placed in ordinary garbage containers. However, some waste must be treated more carefully. It is your responsibility to follow procedures and to dispose of waste in the safest possible manner according to your teacher's instructions.

FIRST AID

The following guidelines apply in case of an injury, such as a burn, cut, chemical spill, ingestion, inhalation, or splash in the eyes:

- Always inform your teacher immediately of any injury.
- If the injury is a minor cut or abrasion, wash the area thoroughly. Using a compress (for example, clean paper towels), apply pressure to the cut to stop the bleeding. When bleeding has stopped, replace the compress with a sterile bandage. If the cut is serious, apply pressure and seek medical attention immediately.
- If you get a chemical in your eye, quickly use the eyewash or nearest running water. Continue to rinse the eye with water for at least 15 min. Unless you have a plumbed eyewash system, you will also need assistance in refilling the eyewash container. Have another student inform your teacher of the accident. The injured eye should be examined by a doctor.
- If the injury is a burn, immediately immerse the affected area in cold water, or run cold water gently over the burned area. This will reduce the temperature and prevent further tissue damage.
- In case of electric shock, unplug the appliance and do not touch it or the victim. Inform your teacher immediately.
- If a classmate's injury has rendered him/her unconscious, notify your teacher immediately. Your teacher will perform CPR if necessary. Do not administer CPR unless under specific instructions from your teacher. Call the school office and request emergency medical help.

A2 Scientific Inquiry

As we observe the natural world, we encounter questions, mysteries, or events that are not readily explainable. To develop explanations, we investigate using inquiry investigations. An important aspect of inquiry investigations is that science is only one of many ways in which people explore, explain, and come to understand the world around them. Inquiry investigations generally involve the following:

- formulating questions that can be answered through investigation
- using appropriate tools and techniques to gather, analyze, and interpret data
- developing descriptions, explanations, predictions, and models using evidence
- thinking critically and logically to make the relationships between evidence and explanations
- recognizing and analyzing alternative explanations and predictions
- communicating scientific procedures and explanations

The methods used in inquiry investigations depend, to a large degree, on the purpose of the inquiry. There are four common types of inquiry investigations: (1) the controlled experiment, (2) the correlational study, (3) the observational study, and (4) the field study. These types of inquiry investigations require specific skills. The skills are discussed below, followed by a detailed description of how they relate to each of the four types of inquiry investigations.

A2.1 Skills of Scientific Inquiry

Certain skills are important in the process of conducting a scientific investigation. These skills can be organized into four categories: initiating and planning, performing and recording, analyzing and interpreting, and communicating.

INITIATING AND PLANNING

(1) Questioning: Most inquiry investigations begin with a question. It is important to ask the right questions. In certain types of inquiry investigations, the question must be testable. This means that it must ask about a possible cause-and-effect relationship. A cause-and-effect relationship is one in which a change in one variable (see #3) causes a change in another variable. A testable question might start in one of the following ways: What is the relationship between . . . and . . . ? How does . . . affect . . . ? If . . . , what happens to . . . ?

(2) Researching: This is a skill that occurs across all four categories of scientific investigation and includes preparing for research, accessing resources, processing information, and transferring learning. The process involves identifying the type of information that is required, using strategies to locate and access the information, recording the information, synthesizing findings, and formulating conclusions.

(3) Identifying variables: Considering the variables involved in an investigation is an important step in designing an effective investigation. Variables are any factors that could affect the outcome of an investigation. There are three kinds of variables in a controlled experiment: the manipulated variable, the responding variable, and the controlled variables.

- The manipulated variable (also known as the independent variable or cause variable) is the variable that is deliberately changed by the investigator.
- The responding variable (also known as the dependent variable or effect variable) is the variable that the investigator believes will be affected by a change in the manipulated variable.
- The controlled variables are variables that may affect the responding variable, but that are held constant so that they cannot affect the responding variable. A controlled experiment is a test of whether (and how) a manipulated variable affects a responding variable. To make the test fair, all other variables that may affect the responding variable are kept constant (unchanging).

(4) Hypothesizing: A hypothesis is a predicted answer to the testable question. It proposes a possible explanation based on an already known scientific theory, law, or other generalization.

A hypothesis may be written in the form of an "If . . . , then . . . because . . ." statement. If the manipulated (independent) variable is changed in a particular way, then we predict that the responding (dependent) variable will change in a particular way, and we provide a theoretical explanation for the prediction. For example,

If an air-filled balloon is placed in a freezer and its temperature is decreased, then its volume will decrease because, according to the kinetic molecular theory, atoms and molecules slow down and occupy less space at lower temperatures.

You may create more than one hypothesis from the same testable question.

When you conduct an investigation, your observations do not always support the prediction in your hypothesis. When this happens, you may re-evaluate and modify your hypothesis and design a new experiment.

(5) Planning: Planning an inquiry investigation involves developing an experimental design, identifying variables, selecting necessary equipment and materials, addressing safety concerns, and writing a step-by-step procedure.

PERFORMING AND RECORDING

(1) Conducting the inquiry investigation: As you perform an investigation, follow the steps in the procedure carefully and thoroughly. Check with your teacher if you find that you need to make significant alterations to your procedure. Use all equipment and materials safely, appropriately, and with precision.

(2) Making observations: When you conduct an inquiry investigation, you should make accurate observations at regular intervals and record them carefully and accurately. Record exactly what you observe. Observations from an experiment may not always be what you expect them to be. Qualitative (descriptive) and quantitative (measured) observations may be made during an investigation. Some observations may also be provided for you during an investigation.

Quantitative observations are based on measured quantities, such as temperature, volume, and mass. They are usually recorded in data tables.

Qualitative observations describe characteristics that cannot be expressed in numbers, such as texture, smell, and taste. They can be recorded using words, pictures, tables, or labelled diagrams.

(3) Collecting, organizing, and recording data: During an investigation you should collect and record all data and observations, and organize these into formats that are easily interpreted (such as tables, charts, etc.).

ANALYZING AND INTERPRETING

(1) Analyzing: Analyzing involves looking for patterns and relationships that will help explain your results and give new information about the question you are investigating. Your analysis will tell you whether your observations support your hypothesis.

(2) Evaluating: It is very important to evaluate the evidence that is obtained through observations and analysis. When evaluating the results of an investigation, here are some aspects you should consider:

- Experimental design: Were there any problems with the way you planned your experiment? Did you control all the variables except for the manipulated variable?
- Equipment and materials: Was the equipment adequate? Would other equipment have been better? Was equipment used incorrectly? Did you have difficulty with a piece of equipment?
- Skills: Did you have the appropriate skills for the investigation? Did you have to use a skill that you were just beginning to learn?
- Observations: Did you accurately record all the relevant observations?

COMMUNICATING

(1) It is important to share both your process and your results. Other people may want to repeat your investigation, or they may want to use or apply your results in another situation. Your write-up or report should reflect the process of scientific inquiry that you used in your investigation.

(2) At this stage, you should be prepared to extend insights and opinions from your findings, suggest areas for further investigation, and relate research findings to the world around you.

In the following sections, we will detail the components of the four types of inquiry investigation: controlled experiments, correlational studies, observational studies, and field studies.

A2.2 Controlled Experiments

A controlled experiment is an inquiry investigation in which a manipulated variable is intentionally changed to determine its effect on a responding variable. All other variables are controlled (kept constant). Controlled experiments are performed when the purpose of the inquiry is to create, test, or use a scientific concept.

The common components of controlled experiments are outlined below. Note that there are normally many cycles through the steps during an actual experiment.

TESTABLE QUESTION

A testable question forms the basis for your controlled experiment: the investigation is designed to answer the question. Controlled experiments are about relationships among variables, so your question could inquire about changes to variable A that may occur as a result of changes to variable B.

VARIABLES

The primary purpose of a controlled experiment is to determine whether a change in a manipulated variable causes a noticeable change in a responding variable while all other variables remain constant. Therefore, you must identify all major variables that you will measure and/or control in your investigation. What is the manipulated (independent) variable? What is the responding (dependent) variable? What are the controlled variables?

When conducting a controlled experiment, change only one manipulated variable at a time, holding all the others (except the responding variable) constant. This way, you can assume that the results are caused by the manipulated variable and not by any of the other variables.

HYPOTHESIS

When formulating a hypothesis, first read the testable question, the experimental design, and the procedure, if provided. Then, try to identify the manipulated variable, the responding variable, and the controlled variables. Your hypothesis will be a predicted answer to your testable question accompanied by a theoretical explanation for your prediction.

EXPERIMENTAL DESIGN

The design of a controlled experiment shows how you plan to answer your question. The design outlines how you will change the manipulated variable, measure any variations in the responding variable, and control all the other variables. It is a summary of your plan for the experiment.

EQUIPMENT AND MATERIALS

Make a detailed list of all equipment and materials used, including sizes and quantities where appropriate. Be sure to include safety equipment, such as eye protection, lab apron, protective gloves, and tongs, where needed. Draw a diagram to show any complicated setup of apparatus.

PROCEDURE

Write a step-by-step description of how you will perform your investigation. It must be clear enough for someone else to follow and it must explain how you will deal with each of the variables in your investigation. The first step in a procedure usually refers to any safety precautions that need to be addressed, and the last step relates to any cleanup that needs to be done.

OBSERVATIONS

There are many ways you can gather and record observations during your investigation. It is helpful to plan ahead and think about what data you will need to answer the question and how best to record them (for example, data tables, pictures, or labelled diagrams may be helpful). This helps to clarify your thinking about the question posed at the beginning, the variables, the number of trials, the procedure, the materials, and your skills. It will also help you organize your evidence for easier analysis.

ANALYZE AND EVALUATE

You will need to analyze and interpret your observations—this may include graphing your data and analyzing any patterns or trends that may be evident in your graphs. After thoroughly analyzing your observations, you may have sufficient and appropriate evidence to enable you to answer the testable question posed at the beginning of the investigation.

You must evaluate the processes that you followed to plan and perform the investigation. Identify and take into account any sources of error and uncertainty in your measurements. Note that experimental error does not include human error, such as careless measurements or spilled reactants or products. Rather, it includes contaminated chemicals, unwanted reactions, and the unavailability of ideal equipment.

You will also evaluate the outcome of the investigation, which involves evaluating your hypothesis/prediction. You must identify and take into account any sources of error and uncertainty in your measurements.

Finally, compare your hypothesis/prediction with the evidence. Is your hypothesis supported by the evidence?

APPLY AND EXTEND

Reflect on how your investigation relates to the world around you: how can you use your findings in everyday life?

REPORTING ON THE INVESTIGATION

Your lab report should describe your planning process and procedure clearly and in enough detail that the reader could duplicate your experiment. You should present your observations, your analysis, and your evaluation of your experiment clearly, accurately, and honestly.

A2.3 Correlational Studies

When the purpose of an inquiry investigation is to test a suspected relationship between two variables, but a controlled experiment is not possible, a correlational study is conducted. In a correlational study, the investigator tries to determine whether one variable is affecting another without purposely changing or controlling any of the variables. Instead, variables are allowed to change naturally. It is often difficult to isolate cause and effect in correlational studies. A correlational inquiry requires large sample numbers and many replications to increase the certainty of the results.

A correlational study does not require experiments or fieldwork; for example, the investigator can use databases prepared by other researchers to find relationships between two or more variables. The investigator can, however, choose to also make observations and measurements through fieldwork, interviews, and surveys.

A hypothesis or prediction may be made in a correlational study but is not as useful as it is in a controlled experiment. Correlational studies are not intended to establish cause-and-effect relationships. However, the results of a correlational study can be used to formulate a hypothesis about the causal relationship between the variables.

The common components of a correlational study are outlined below. Even though the sequence is presented as linear, there are normally many cycles through the steps during an actual study.

PURPOSE

When planning a correlational study, it is important that you pose a question about a possible statistical relationship between variable A and variable B. Choose a topic that interests you. Determine whether you are going to replicate or revise a previous study, or create a new one. Indicate your decision in the statement of the purpose.

VARIABLES

In a correlational study you must determine whether two variables are related, without controlling any of the variables. You must identify all the major variables that will be measured and/or observed in your investigation.

STUDY DESIGN

When designing your correlational study you must identify the setting and the methods you will use in carrying out your investigation. You should describe the type(s) and source(s) of data that you plan to collect. Your design should address questions such as the following: Will you be conducting a survey? If so, where will the survey be conducted? Who will answer your questionnaire? When will the survey be conducted? How often will the survey be administered? If you are obtaining information from an existing database, then describe the source of the information and your plans for analyzing the information.

EQUIPMENT AND MATERIALS

Make a detailed list of all equipment and materials used, including sizes and quantities where appropriate. Be sure to include safety equipment, such as eye protection and ear protection, where needed. Draw a diagram to show any complicated setup of apparatus.

PROCEDURE

Write a step-by-step description identifying how you will gather data on the variables under study. You will also need to identify potential sources of data. There are two possible sources: observations made by you, the investigator; and existing data (databases, etc.).

OBSERVATIONS

If you are collecting your own data through observation then you will need to plan ahead and think about what data you will need and how best to record them. There are many ways to gather and record your observations (such as data tables, pictures, or labelled diagrams). This is an important step because it helps to clarify your thinking about the question posed at the beginning, the variables, the procedure, the materials, and your skills. It will also help you organize your observations for easier analysis.

ANALYZE AND EVALUATE

Analyze and interpret your observations or sourced data. This will usually include graphing the data and analyzing any patterns or trends in your graphs. Identify any relationship between your two variables. A positive correlation indicates a direct relationship between variables: an increase in one variable corresponds to an increase in another (**Figure 1(a)**). A negative correlation indicates an inverse relationship: an increase in one variable corresponds to a decrease in the other (**Figure 1(b)**). If there is no relationship between the variables then there is no correlation (**Figure 1(c)**).

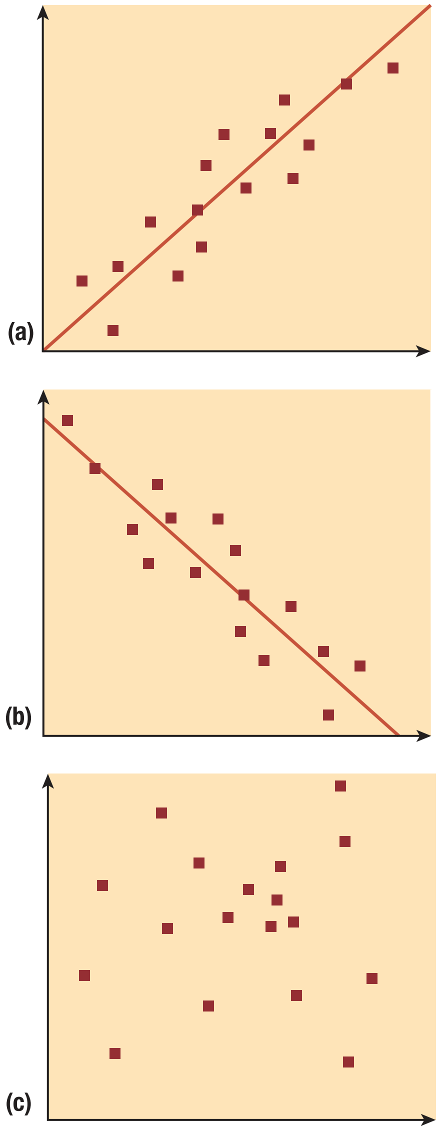


Figure 1 (a) In a positive correlation, variable *y* increases as variable *x* increases. (b) In a negative correlation, *y* decreases as *x* increases. (c) If there is no correlation, there is no pattern.

When you know that two variables are correlated, you can predict the value of one variable on the basis of the other. Generally, the stronger the correlation, the more probable it is that your prediction will be correct. After analyzing your observations, revisit the question posed at the beginning of the investigation. Is there a relationship between variable A and variable B?

Evaluate the processes that you followed to plan and perform the investigation. Also evaluate the outcome of the investigation, which involves evaluating any prediction you made at the beginning of the investigation. You must identify and take into account any sources of error and uncertainty in your measurements. Note that experimental error does not include human error, such as careless measurements or spilled reactants or products. Rather, it includes contaminated chemicals, unwanted reactions, and the unavailability of ideal equipment.

APPLY AND EXTEND

Reflect on how your investigation relates to the world around you: how can you use what you have learned in your everyday life?

REPORTING ON THE INVESTIGATION

In preparing your report, your objectives should be to describe your design and procedure accurately and to report your observations, analyses, and evaluations accurately and honestly.

A2.4 Observational Studies

Often, the purpose of an inquiry is simply to study a natural phenomenon with the intention of gaining scientifically significant information to answer a question. Observational studies involve observing a subject or phenomenon in an unobtrusive manner, usually with no specific hypothesis. The inquiry does not start off with a hypothesis, but a hypothesis may be generated as information is collected.

The stages and processes of observational studies are summarized below. Even though the sequence is presented as linear, there are normally many cycles through the steps during an actual study.

PURPOSE

In planning an observational study, it is important to pose a general question about the natural world. Choose a topic that interests you. Determine whether you are going to replicate or revise a previous study, or create a new one. Indicate your decision in a statement of the purpose.

EQUIPMENT AND MATERIALS

Make a detailed list of all equipment and materials used, including sizes and quantities where appropriate. Be sure to include safety equipment, such as eye protection, lab apron, protective gloves, and tongs, where needed. Draw a diagram to show any complicated setup of apparatus.

PROCEDURE

Write a step-by-step description of how you will make your observations. It must be clear enough for someone else to follow. The first step in a procedure usually refers to any safety precautions that need to be addressed, and the last step relates to any cleanup that needs to be done.

OBSERVATIONS

There are many ways you can gather and record observations—including quantitative observations—during an observational study. All observations should be objective and unambiguous. Organize your information for easier analysis.

ANALYZE AND EVALUATE

After thoroughly analyzing your observations, you may have sufficient and appropriate evidence to enable you to answer the question posed at the beginning of the investigation. You may also have enough observations and information to form a hypothesis for a controlled experiment.

At this stage of the investigation, you will evaluate the processes used to plan and perform the investigation. Evaluating the processes includes evaluating the materials, the design, the procedure, and your skills. The results of most such investigations will suggest further studies, perhaps controlled experiments or correlational studies, to explore tentative hypotheses you may have developed.

APPLY AND EXTEND

At this stage you should reflect on how your investigation relates to the world around you: how can you use what you have learned in your everyday life?

REPORTING ON THE INVESTIGATION

In your report, describe your design and procedure accurately, and report your observations accurately and honestly.

A2.5 Field Studies

A field study is a special case of observational study. The investigation is being conducted outside of the classroom setting and will require special planning and considerations. In most cases, there will be special safety considerations to take into account since the study is conducted in a natural environment. If a field study is conducted near a pond, lake, or stream, then water safety needs to be taken into consideration. Since field studies often involve studying natural environments, it is important to consider how the study will affect the biotic and abiotic components of the environment. The investigator should take steps to minimize any potentially negative effects on organisms or their environment.

A2.6 Lab Reports

When carrying out inquiry investigations, it is important that scientists keep records of their plans and results, and share their findings. In order to have their investigations repeated (replicated) and accepted by the scientific community, scientists generally share their work by publishing reports in which details of their design, materials, procedure, evidence, analysis, and evaluation are provided.

Lab reports are prepared after an investigation is completed. To ensure that you can accurately describe the investigation, keep thorough and accurate records of your activities as you carry out the investigation. Your lab book or report should reflect the type of scientific inquiry that you used in the investigation (controlled experiment, correlational study, observational study, or field study) and should be based on the following headings, *as appropriate*:

TITLE

At the beginning of your report, write the number and title of your investigation. If you are designing your own investigation, create a title that suggests what the investigation is about. Include the date the investigation was conducted and the names of all lab partners (if you worked as a team).

PURPOSE

State the purpose of the investigation. Why are you doing this investigation?

TESTABLE QUESTION

State the question that you attempted to answer in the investigation. Sometimes the question is provided for you; other times you are expected to formulate your own. If it is appropriate to do so, state the question in terms of manipulated and responding variables.

HYPOTHESIS/PREDICTION

For a controlled experiment you will usually have to compose a hypothesis or prediction. This will be a proposed answer to your testable question. When writing a hypothesis, include both a prediction and a reason for the prediction, based on scientific theory, law, or other generalization. You may use the "If . . . , then . . . because . . ." form. A simple prediction may be written in the "If . . . , then . . ." form.

VARIABLES

Identify all major variables that you measured and/or controlled in the investigation. What is the manipulated variable? What is the responding variable? What are the major controlled variables?

EXPERIMENTAL DESIGN

Provide a brief general overview (one to three sentences) of what you did in your investigation. If your investigation involved manipulated, responding, and controlled variables, list them and indicate how they were changed, measured, or held constant. Identify any control or control group that was used in the investigation.

EQUIPMENT AND MATERIALS

Include a detailed list of all equipment and materials used, including sizes and quantities where appropriate. Be sure to include safety equipment, such as eye protection, lab apron, protective gloves, and tongs, where needed. Draw a diagram to show any complicated setup of apparatus.

PROCEDURE

Describe, in detailed, numbered steps, the procedure you followed to carry out your investigation. Your teacher may specify which style you should use. Examples of three common writing styles are

1. third person past tense ("The test tubes were heated . . .")
2. first person plural past tense ("We heated the test tubes . . .")
3. second person imperative ("Heat the test tubes . . .")

Include steps to clean up and dispose of waste.

OBSERVATIONS

Present your observations in a form that is easily understood. This includes all the qualitative and quantitative observations that you made. Be as precise as possible when describing quantitative observations. Include any unexpected observations and present your information clearly. If you have only a few observations, this could be a list; for controlled experiments and for many observations, a data table, labelled diagram, or written descriptions would be more appropriate.

ANALYSIS

Complete the questions found in the Analyze and Evaluate section of the investigation outline. These questions will prompt you to analyze and interpret your observations, answer a testable question, draw conclusions, and evaluate both your experiment and your conclusions. You will also be prompted to graph your data and analyze these graphs where applicable.

If you are writing up an investigation for which there are no questions, write your own analysis. Interpret your observations and present the evidence in the form of titled tables, graphs, or illustrations, as appropriate. Include any calculations, the results of which can be shown in a table. Make statements about any patterns or trends you observed. Conclude the analysis with a statement based only on the evidence you have gathered, answering the question that initiated the investigation.

EVALUATION

The evaluation is your judgment about the quality of evidence obtained and about the validity of the prediction and hypothesis (if present). This section can be divided into two parts:

(1) Evaluation of the procedure and/or experimental design: Did your procedure/design provide reliable and valid evidence to enable you to answer the question? Consider the experimental design, the procedure, and your laboratory skills. Were they all adequate? Are you confident enough in the evidence to use it to evaluate any prediction and/or hypothesis you made?

(2) Evaluation of the prediction/hypothesis: Was your prediction or hypothesis supported or not supported by the evidence? Answer the question that you posed at the beginning of the investigation on the basis of your evaluation of your prediction or hypothesis.

APPLY AND EXTEND

Answer any Apply and Extend questions in the investigation outline. Number your answers as they appear in the Apply and Extend section in the textbook.

A3 Laboratory Skills and Techniques

A3.1 Lighting and Using a Bunsen Burner

Practise and memorize the procedure outlined below.

Note the safety caution. You are responsible for your safety and the safety of others near you.

1. Turn the air and gas adjustments to the off position **(Figure 1)**.
2. Connect the burner hose to the gas outlet on the bench.
3. Turn the bench gas valve to the fully on position.
4. If you suspect that there may be any gas leaks, replace the burner. (Give the leaky burner to your teacher.)

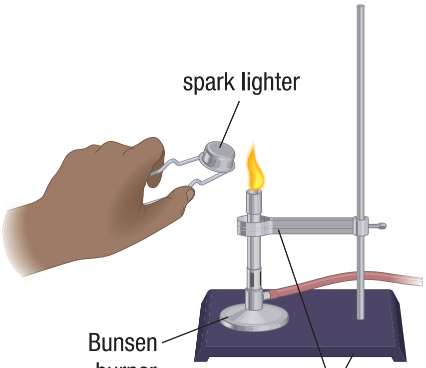


Figure 1 Lighting a Bunsen burner

5. Holding a spark lighter ready above and to one side of the barrel, open the burner gas valve. Immediately strike a few sparks until a small yellow flame results.
6. Adjust the airflow and obtain a pale blue flame with a dual cone **(Figure 2)**. In most common types of burners, rotating the barrel adjusts the air intake. Rotate the barrel slowly. If too much air is added, the flame may go out. If this happens, immediately turn the gas flow off. Relight the burner using the procedure outlined above.

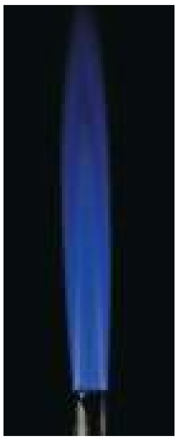


Figure 2 A blue flame indicates complete combustion by a Bunsen burner.

7. Adjust the gas valve on the burner to increase or decrease the height of the blue flame. The hottest part of the flame is the tip of the inner blue cone. You will usually use a 5 to 10 cm flame that just about touches the object being heated.

Never leave a lit burner unattended. If the burner is on but not being used, adjust the air and gas intakes to obtain a small yellow flame. This flame is more visible and, therefore, less likely to cause problems.

When lighting or using a laboratory burner, never position your head or fingers directly above the barrel. Tie back long hair and secure loose clothing.

A3.2 Using a Laboratory Balance

There are two types of balances: electronic and mechanical **(Figure 3)**. All balances must be handled carefully and kept clean. Always place chemicals on paper or into a container to avoid contamination and corrosion of the balance pan. Always record masses using the correct precision. On a centigram balance, mass is measured to the nearest hundredth of a gram (0.01 g). If you must move a balance, hold the instrument by the base and steady the beam. Never lift a balance by the beams or pans.

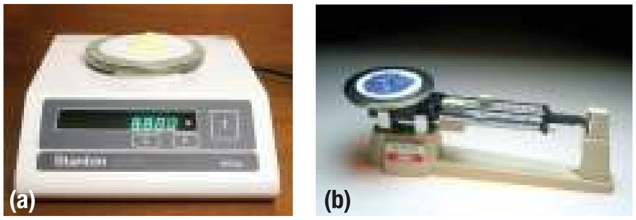


Figure 3 (a) An electronic balance (b) A mechanical balance

USING AN ELECTRONIC BALANCE

Electronic balances are sensitive to small movements and changes in level; do not lean on the counter when using the balance.

1. Place a container or massing paper on the balance.
2. Reset (tare) the balance so the mass of the container registers as zero.
3. Add the chemical until the desired mass of chemical is displayed. Air currents or the high sensitivity of the balance may cause the last digit to vary.
4. Remove the container and chemical.

USING A MECHANICAL BALANCE

There are several different kinds of mechanical balances. Although they work slightly differently, some general procedures apply to most of them.

1. Clean and zero the balance. (Turn the zero adjustment screw so that the beam is balanced when the instrument reads 0 g and no load is on the pan.)
2. Place the container or paper on the pan.

3. Move the largest beam mass one notch at a time until the beam drops, and then move the mass back one notch.
4. Repeat this process with the next smaller mass and continue until all masses have been moved and the beam is balanced. If you are using a dial type balance, the final step will be to turn the dial until the beam balances.
5. Record the mass of the container.
6. Set the masses on the beams to correspond to the total mass of the container (or paper) plus the desired sample.
7. Add the chemical until the beam is once again balanced.
8. Remove the chemical from the pan and return the balance to the zero position.

A3.3 Measuring Volume

When you are asked to measure a volume, the substance to be measured is usually a liquid. Choose the appropriate container to obtain the required precision of measurement. For example, if you need "approximately 100 mL of water," it is appropriate to use a 100 mL or 200 mL beaker. However, if you need to measure more precisely, choose the smallest available vessel that can measure that volume. For example, to measure 853 mL of a solution, use a 1 L graduated cylinder. Measuring smaller volumes precisely requires smaller, narrow measuring instruments. A 50 mL graduated cylinder is appropriate for measuring volumes between 10 mL and 50 mL. Any volume less than 10 mL should be measured in a pipette.

When taking a volume measurement in a graduated vessel, always look at the surface of the liquid at eye level, and read the bottom of the meniscus (curved surface). You might find it helpful to use a meniscus finder **(Figure 4)**.

Figure 4 A meniscus finder helps to make the meniscus more visible.

A3.4 Using a Pipette

A pipette is a specially designed glass or plastic tube used to measure precise volumes of liquids. There are two types of pipettes and a variety of sizes for each type. A volumetric pipette (**Figure 5(a)**) can be used to measure a fixed volume, such as 10.00 mL or 25.00 mL, accurate to within 0.04 mL. A graduated pipette (**Figure 5(b)**) measures a range of volumes, just as a graduated cylinder does. A 10 mL graduated pipette measures volumes accurate to within 0.1 mL.

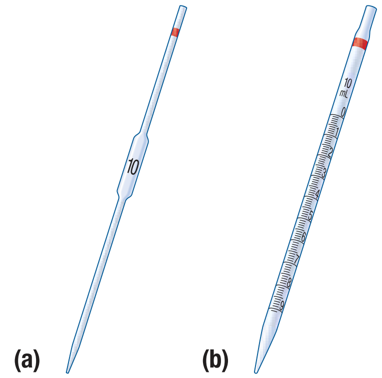


Figure 5 (a) A volumetric pipette (b) A graduated pipette

Never use your mouth to draw a liquid into a pipette. Always use a pipette bulb or pump.

1. Rinse the pipette with small volumes of distilled water using a wash bottle, and then with the sample solution. A clean pipette has no visible residue or liquid drops clinging to the inside wall.
2. Hold the pipette with your thumb and fingers near the top. Leave your index finger free.
3. Place the pipette in the sample solution, resting the tip on the bottom of the container if possible. Be careful that the tip does not hit the sides of the container.
4. Squeeze the bulb into the palm of your hand and then place the bulb firmly and squarely on the end of the pipette (**Figure 6**) with your thumb across the top of the bulb.

Figure 6 Release the bulb slowly. Pressing down with your thumb placed across the top of the bulb maintains a good seal. Setting the pipette tip on the bottom slows the rise or fall of the liquid.

5. Release your grip on the bulb until the liquid has risen above the calibration line. This may require bringing the level up in stages: remove the bulb, put your finger over the top of the pipette, squeeze the air out of the bulb, replace the bulb, and continue the procedure.
6. Remove the bulb, placing your index finger over the top. If you are using a dispensing bulb (**Figure 7**) or a pipette pump (**Figure 8**), leave it attached to the pipette.

Figure 7 A dispensing pipette bulb uses a small valve in the side stem to control the flow of liquid in a pipette.

Figure 8 Pipette pumps have a small roller that can be turned to draw liquid in or to let liquid out of the pipette.

7. Wipe all solution from the outside of the pipette using a paper towel.
8. While touching the tip of the pipette to the inside of a waste beaker, gently roll your index finger (or squeeze the valve of the dispensing bulb) to allow the liquid level to drop until the bottom of the meniscus reaches the desired calibration line (**Figure 9**).

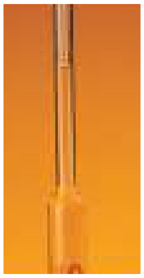


Figure 9 To allow the liquid to drop slowly to the calibration line, it is necessary for your finger and the pipette top to be dry. Also keep the tip on the bottom to slow down the flow.

To avoid errors, read the meniscus at eye level. Stop the flow when the bottom of the meniscus is on the desired calibration line. Use the bulb to raise the level of the liquid again if necessary.

9. While holding the pipette vertically, touch the pipette tip to the inside wall of a clean receiving container. Remove your finger or adjust the valve and allow the liquid to drain freely until the solution stops flowing.
10. Finish by touching the pipette tip to the inside of the container held at about a 45° angle (**Figure 10**). Do not shake the pipette. The delivery pipette is calibrated to leave a small volume in the tip.

Figure 10 A vertical volumetric pipette is drained by gravity and then the tip is placed against the inside wall of the container. A small volume is expected to remain in the tip.

A3.5 Using a Burette for Titration

1. Rinse the burette with small volumes of pure water using a wash bottle. Using a burette funnel, rinse with small volumes of the titrant (**Figure 11**). (If liquid droplets remain on the sides of the burette after rinsing, scrub the burette with a burette brush. If the tip of the burette is chipped or broken, replace the tip or the whole burette.)

Figure 11 A burette should be rinsed with water and then with the titrant before use.

2. Using a small funnel, pour the titrant solution into the burette until the level is near the top. Open the valve for maximum flow to clear any air bubbles from the tip and to bring the liquid level down to the scale.
3. Record the initial burette reading to the nearest 0.1 mL. Avoid errors by reading volumes at eye level with the aid of a meniscus finder.
4. Measure a known volume of the solution of unknown concentration into a clean Erlenmeyer flask. Place a white piece of paper beneath the flask to make it easier to detect colour changes.
5. Add an indicator if one is required. Add the smallest quantity necessary (usually 1 to 2 drops) to produce a noticeable colour change in your sample.
6. Add the titrant from the burette quickly at first, and then slowly, drop by drop, near the endpoint (**Figure 12**). Stop as soon as a drop of the titrant produces a permanent colour change in the sample solution. A permanent colour change is considered to be a noticeable change that lasts for at least 20 s after swirling.

Figure 12 Near the endpoint, continuous gentle swirling of the solution is particularly important.

7. Record the final burette reading to the nearest 0.1 mL.
8. The final burette reading for one trial becomes the initial burette reading for the next trial. Three trials with results within 0.2 mL are normally required for a reliable analysis of an unknown solution.
9. Drain and rinse the burette with pure water. Store the burette upside down with the valve open.

A3.6 Separating Mixtures

FILTRATION

If you have a mixture that includes a solid in a liquid (often a precipitate from a double displacement reaction in water), you can separate out the solid by filtration.

1. If your aim is to determine the mass of the solid, first measure the mass of a piece of filter paper.
2. Place the folded filter paper in a funnel. Rest the funnel in a clamp attached to a retort stand. Place a beaker below the funnel to catch the filtrate.
3. Slowly pour the mixture into the filter (**Figure 13**). Never fill the filter paper more than two-thirds full. Scrape any remaining solids out of the original container with a stirring rod or small spatula.

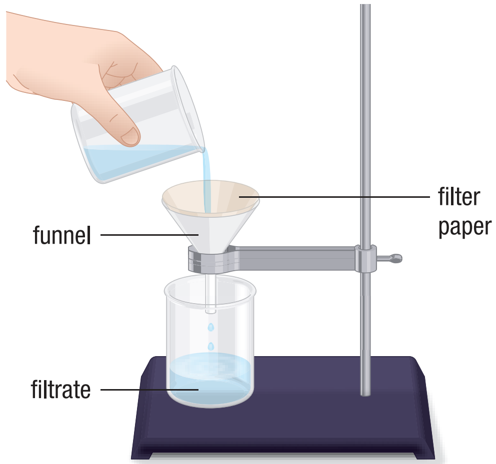


Figure 13 Do not overfill the filter paper as you pour the mixture in.

4. Use a wash bottle of distilled water to rinse the container that held the mixture. Pour the rinse water into the filter also. Repeat. Allow all the water to drain through into the lower beaker.
5. Carefully remove the filter paper and solid from the funnel and place it on several layers of paper towel. Open the filter paper so that it lies flat. Leave it somewhere warm overnight to dry.
6. Measure the mass of the dry filter paper and solid. Subtract the mass of the filter paper (your initial measurement) to determine the mass of the solid.

CRYSTALLIZATION

If you have an aqueous solution of an ionic compound, you can separate the solute by evaporating the water. Place the solution in a wide container and either set it somewhere warm or heat it gently over a heat source. If you use a heat source, the solution must be in a heat-resistant container such as a Pyrex beaker or an evaporating dish. If you are using a hot plate to warm a beaker, place the beaker on a wire gauze with a ceramic centre. If you are using an evaporating dish, place it on a clay pipe triangle and heat it gently with a Bunsen burner (**Figure 14**).

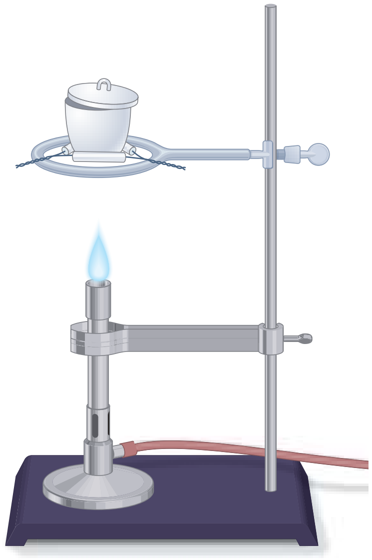


Figure 14 Apparatus for separation of a solute from solution by crystallization.

DECANTING

Decanting is a technique that removes the top layer of liquid from a heterogeneous mixture. It involves slowly tipping a beaker containing the mixture so that the top layer is carefully poured into another container and the denser portions stay behind. Place a glass rod in the receiving container and hold the spout of the tipped beaker against this rod as you pour (**Figure 15**). This prevents the liquid from running along the underside of the beaker.

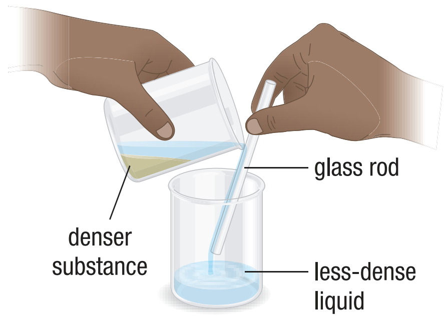


Figure 15 When decanting a less dense liquid from one vessel to another, hold a glass rod against the lip of the tilted vessel. This prevents drips from running down the outside of the vessel.

A3.7 Preparing a Standard Solution

A STANDARD SOLUTION FROM A SOLID

Sometimes you will have to prepare a solution of known concentration from a solid solute and water.

1. Calculate the required mass of solute from the volume and concentration of the solution.
2. Measure the required mass of solute on a piece of massing paper or in a clean, dry beaker (A3.2).
3. Pour distilled water into a beaker to a volume that is less than half of the final volume of the solution. Transfer the solute to the water. Stir to dissolve.
4. Transfer the solution into a clean volumetric flask. Use a funnel to avoid spills. Rinse the beaker and any other equipment two or three times with distilled water, adding the rinse water to the volumetric flask. Add more distilled water to the flask until the level of the liquid is a few millilitres less than the desired volume.
5. Use a dropper bottle to add the final few millilitres of distilled water. Use a meniscus finder to set the bottom of the meniscus on the calibration line (Figure 4).
6. Place a stopper in the mouth of the flask. Swirl the flask to mix the solution.

A STANDARD SOLUTION FROM A CONCENTRATED SOLUTION

Sometimes you will have to prepare a solution of known concentration using an existing stock (concentrated) solution.

1. Calculate the volume of stock solution required.
2. Add approximately one-half of the final volume of distilled water to the volumetric flask.
3. Measure the required volume of stock solution using a pipette. (Refer to A3.4, Using a Pipette.)
4. Transfer the stock solution slowly into the volumetric flask, using a funnel. Rinse the container two or three times with distilled water, adding the rinse water to the volumetric flask. Add more distilled water to the flask until the level of the liquid is a few millilitres less than the desired volume.

If water is added directly to some solids or concentrated liquids, there may be boiling or splattering. Always add a solid solute or concentrated liquids to water.

5. Use a dropper bottle to add the final few millilitres of distilled water. Use a meniscus finder to set the bottom of the meniscus on the calibration line.
6. Place a stopper in the mouth of the flask. Swirl the flask to mix the solution.

A3.8 Identifying Products (Diagnostic Tests)

The tests described in **Table 1** are commonly used to detect a specific substance. All diagnostic tests include a brief procedure, some expected evidence, and an interpretation of the evidence obtained. This is conveniently communicated using the format, "If [procedure] and [evidence], then [analysis]." For example, "If a drop of a solution is placed on blue litmus paper and the paper remains blue, then the solution is neutral or basic and its pH is 7 or higher."

Table 1 Some Standard Diagnostic Tests

Substance tested	Diagnostic test
water	If cobalt(II) chloride paper is exposed to a liquid or vapour and the paper turns from blue to pink, then water is likely present.
oxygen	If a glowing splint is inserted into the test tube and the splint glows brighter or relights, then oxygen gas is likely present.
hydrogen	If a flame is inserted into the test tube and a squeal or pop is heard, then hydrogen is likely present.
carbon dioxide	If the unknown gas is bubbled into a limewater solution and the limewater turns cloudy, then carbon dioxide is likely present.
halogens	If a few millilitres of a hydrocarbon solvent is added, with shaking, to a solution in a test tube and the colour of the solvent appears to be • light yellow-green, then chlorine is likely present • orange, then bromine is likely present • purple, then iodine is likely present
acid	If strips of blue and red litmus paper are dipped into the solution and the blue litmus turns red, then an acid is present.
base	If strips of blue and red litmus paper are dipped into the solution and the red litmus turns blue, then a base is present.
neutral solution	If strips of blue and red litmus paper are dipped into the solution and neither litmus changes colour, then only neutral substances are likely present.
neutral ionic solution	If a neutral solution is tested for conductivity with a multimeter and the solution conducts a current, then a neutral ionic substance is likely present.
neutral molecular solution	If a neutral solution is tested for conductivity with a multimeter and the solution does not conduct a current, then a neutral molecular substance is likely present.

Diagnostic tests can be constructed using any characteristic empirical property of a substance. For example, diagnostic tests for acids, bases, and neutral substances can be specified in terms of the pH of the solutions.

For specific chemical reactions, properties of the products that the reactants do not have, such as the insolubility of a precipitate, the failure to produce a gas, or the lack of colour of ions in aqueous solution, can be used to construct diagnostic tests.

If possible, you should use a control to illustrate that the test does not give the same results with other substances. For example, in the test for oxygen, you typically insert a glowing splint into a test tube that contains a gas suspected to be oxygen. You should also conduct the test on a test tube containing air, to compare the effects. This is the control test. If the splint reacts the same way in both test tubes, you could not conclude that the "oxygen" test tube contains any more oxygen than is present in air.

There are thousands of diagnostic tests. You can create some of these using data from the periodic table and from the solubility and ion colour tables (Appendix B4).

A3.9 Using Probes

Several companies produce electronic equipment to measure pH, electrical conductivity, temperature, gas pressure, and many other variables. Some of these probes are self-contained; others must be connected to a computer. Pay attention to the instructions provided when you use a probe. It is a delicate device and must be handled carefully. If you are not sure how to use the probe properly, ask your teacher.

A4 Scientific Publications

Communicating in Science

Advances in science and our understanding of the natural world are the result of scientists sharing ideas and information. Ernest Rutherford used alpha particles emitted by the radioactive element polonium to discover the nucleus. Rutherford's discovery was possible only because Marie Curie had recognized, a few years earlier, that polonium is radioactive.

It is important for scientists to share ideas and research with other scientists. New research findings are shared at conferences or in journal publications. Research scientists take pride in being published. Being published confirms that the research adds to the knowledge base of the scientific community. Sharing information helps to spread knowledge, solve problems, and inspire other scientists.

The Scientific Journal

Scientific journals are publications that are used to report the results of new research. There are thousands of different science journals published worldwide, and in many languages. A scientific journal may be specific to a subject (for example, *Journal of Atmospheric Chemistry*) or contain articles that cover a variety of subjects within a field (for example, *Nature*). These publications may be electronic (online) or in print and may be published weekly, monthly, bimonthly or quarterly.

Peer Review

When an article is submitted to a journal the research findings are critically reviewed by experts in that discipline. This ensures that the research presents ideas that are supported by practices of good science. High-quality evidence and appropriate conclusions are necessary for the article to be accepted for publication. An article that is submitted to a reputable journal may take months to be approved for publication. The article may be returned to the author(s) for revision if necessary.

Scientists aim to be published in the most respected journals. The peer-review process contributes to the reputation of a journal. It helps to maintain standards and provide credibility. A scientific journal becomes reputable by ensuring that only articles of high quality are published. Very prestigious journals (such as *Nature, Science*, and *Pure and Applied Chemistry*) are known for publishing research backed by only the best practices of science. Reputable journals are widely read and considered reliable by the scientific community. Publication in a reputable journal brings immediate recognition for the author(s). Research that is not published in a peer-reviewed journal is often overlooked.

Format of Research Articles

Research articles have specific sections. Articles often include an abstract, introduction, methods, results, discussion, conclusions, and references.

THE ABSTRACT

The abstract is a short summary of the article. It presents the purpose of the research, outlines the design of the methods used, and summarizes findings or conclusions. A well-written abstract is useful when looking for articles with specific information. Time may be saved by reading the abstract and then deciding if the article is going to be helpful in supporting research.

CITING SOURCES AND GIVING CREDIT

Once a scientist has found useful sources and included them in an article or paper, information must be provided about the article so that someone else who is interested in learning more will be able to find it. This also shows the reader that information supporting the research is current.

More importantly, citing another scientist's work provides a measure of value of what he or she has published. It gives credit to the scientist. More citations often mean that the work is worthwhile and influential. For example, in 1939 Linus Pauling wrote a very influential book entitled *The Nature of the Chemical Bond*. By 1969, this book had been cited over 16 000 times by other scientists working the field! Similarly, Ronald Gillespie of McMaster University developed a simple yet powerful theory to explain the shapes of molecules. Soon after his work was published in 1957, citations started appearing in scholarly journals, and have continued to appear ever since.

Further Reading

Day, R.A., & Gastel, B. (2006). *How to write and publish a scientific paper* (6th ed.). Cambridge: Cambridge University Press.

A5 Exploring Issues and Applications

Throughout this textbook you will have many opportunities to examine the connections between science, technology, society, and the environment (STSE) by exploring issues and applications.

An issue is a situation in which several points of view need to be considered in order to make a decision. There can be many positions, generally determined by the values that an individual or a society holds, on a single issue. Which solution is “best” is a matter of opinion; ideally, the solution that is implemented is the one that is most appropriate for society as a whole. Researching information about an issue will help you make an educated decision about it. All the skills listed in Section A5.1 may be useful in an activity that involves exploring an issue.

Scientific research produces knowledge or understanding of natural phenomena. Technologists and engineers look for ways to apply this knowledge in the development of practical products and processes. Technological inventions and innovations can have wide-ranging applications for, and impacts on, society and the environment. The purpose of exploring an application is to research a particular technological invention or innovation to determine how it works, how it is used, and how it may affect society and the environment. The skills of researching, communicating, and evaluating may be useful in an activity that involves exploring an application.

A5.1 Research Skills

The following skills are involved in many types of research. Some of these skills will help you research issues only, while some will help you research issues or applications. Refer to this section when you have questions about any of the following skills and processes.

DEFINING THE ISSUE

When exploring an issue, the first step in understanding the issue is to explain why it exists, the problems associated with it, and, if applicable, the individuals or groups, also known as stakeholders, that are involved in it. The issue includes information about the role a person takes when thinking about an issue as well as a description of who your audience will be. You could brainstorm questions involving Who? What? Where? When? Why? and How? Develop background information on the issue by clarifying facts and concepts, and identifying relevant attributes, features, or characteristics of the problem.

RESEARCHING

When beginning your research for both issues and applications, you need to formulate a research question that helps to limit, narrow, or define the scope of your research. You then need to develop a plan to find reliable and relevant sources of information. This includes outlining the stages of your research: gathering, sorting, evaluating, selecting, and integrating relevant information. You should gather information from a variety of sources if possible (for example, print, web, and personal interviews).

As you collect information, do your best to ensure that the information is reliable, accurate, and current. Avoid biased opinions: those that are not supported by or that ignore credible evidence. It is important to ensure that the information you have gathered addresses all aspects of the issue or application you are researching.

IDENTIFYING ALTERNATIVES

When exploring an issue, examine the situation and think of as many alternative solutions as you can. Be creative about combining the solutions. At this point, it does not matter if the solutions seem unrealistic. To analyze the alternatives, you should examine the issue from a variety of perspectives. Stakeholders may bring different viewpoints to an issue and these may influence their position on the issue. Consider these viewpoints as you brainstorm or hypothesize how different stakeholders would feel about your alternatives.

ANALYZING THE ISSUE

An important part of exploring an issue is analyzing the issue. First, you should establish criteria for evaluating your information to determine its relevance and significance. You can then evaluate your sources, determine what assumptions may have been made, and assess whether you have enough information to make your decision.

To effectively analyze an issue you should

- establish criteria for determining the relevance and significance of the data you have gathered
- evaluate the sources of information
- identify and determine what assumptions have been made
- challenge unsupported evidence
- evaluate the alternative solutions, possibly by conducting a risk-benefit analysis

Once the issue has been analyzed, you may begin to consider possible solutions. You may decide to carry out a risk-benefit analysis—a tool that enables you to look at each possible result of a proposed action and helps you make a decision. (See Section A5.2 for more information.)

DEFENDING A DECISION

After analyzing your information on your issue, you can answer your research question and take an informed position or draw a conclusion on the issue. If you are working as a group, this is the stage where everyone gets a chance to share ideas and information gathered about the issue. Then the group needs to evaluate all the possible alternatives and decide on their preferred solution based on the criteria.

Your position on the issue, or conclusion, must be justified using supporting information that you have researched. You should be able to defend your position to people with different perspectives. Ask yourself the following questions:

- Do I have supporting evidence from a variety of sources?
- Can I state my position clearly?
- Can I show why this issue is relevant and important to society?
- Do I have solid arguments (with solid evidence) supporting my position?
- Have I considered arguments against my position, and identified their faults?
- Have I analyzed the strong and weak points of each perspective?

COMMUNICATING

When exploring an issue, there are several things to consider when communicating your decision. You need to state your position clearly and take into consideration who your audience is. You should always support your decision with objective data and a persuasive argument if possible. Be prepared to defend your position against any opposition.

You should be able to defend your solution in an appropriate format—debate, class discussion, speech, position paper, multimedia presentation, brochure, poster, video, etc.

When exploring an application you should communicate the "need or want" for the application (why the application was developed in the first place), the "how" (how the application/technology actually works), and the risks and benefits to society, individuals, and the environment. You should conclude with your "assessment" of the application.

EVALUATING

The final phase of your decision making when exploring an issue includes evaluating the decision itself and the process used to reach the decision. After you have made a decision, carefully examine the thinking that led to your decision.

Some questions to guide your evaluation include:

- What was my initial perspective on the issue? How has my perspective changed since I first began to explore the issue?
- How did we make our decision? What process did we use? What steps did we follow?
- To what extent were my arguments factually accurate and persuasively made?
- In what ways does our decision resolve the issue?
- What are the likely short- and long-term effects of the decision?
- To what extent am I satisfied with the final decision?
- What reasons would I give to explain our decision?
- If we had to make this decision again, what would I do differently?

A5.2 Risk-Benefit Analysis Model

Risk-benefit analysis is a tool used to organize and analyze information gathered in research, especially when exploring a socio-scientific issue. A thorough analysis of the risks and benefits associated with each alternative solution can help you decide on the best alternative.

- Research as many aspects of the situation as possible. Look at it from different perspectives.
- Collect as much evidence as you can, including reasonable projections of likely outcomes if the proposal is adopted.
- Classify every individual potential result as being either a benefit or a risk.
- Quantify the size of the potential benefit or risk (perhaps as a dollar figure, or a number of lives affected, or on a scale of 1 to 5).
- Estimate the probability (percentage) of that event occurring.
- By multiplying the size of a benefit (or risk) by the probability of its happening, you can calculate a probability value for each potential result.
- Total the probability values of all the potential risks, and all the potential benefits.
- Compare the sums to help you decide whether to accept the proposed action.

A6 Math Skills

A6.1 Scientific Notation

It is difficult to work with very large or very small numbers when they are written in common decimal notation. Usually it is possible to accommodate such numbers by changing the SI prefix so that the number falls between 0.1 and 1000. For example, 237 000 000 mm can be expressed as 237 km, and 0.000 000 895 kg can be expressed as 0.895 mg. However, this prefix change is not always possible, either because an appropriate prefix does not exist or because it is essential to use a particular unit of measurement in a calculation. In these cases, the best method of dealing with very large and very small numbers is to write them using scientific notation. Scientific notation expresses a number by writing it in the form $a \times 10^n$, where $1 \leq |a| < 10$ and the digits in the coefficient a are all significant. **Table 1** shows situations where scientific notation would be used.

Table 1 Examples of Scientific Notation

Expression	Common decimal notation	Scientific notation
"124.5 million kilometres"	124 500 000 km	1.245×10^8 km
"154 thousand picometres"	154 000 pm	1.54×10^5 pm
"602 sextillion molecules"	602 000 000 000 000 000 000 000 molecules	6.02×10^{23} molecules

To multiply numbers in scientific notation, multiply the coefficients and add the exponents. To divide numbers in scientific notation, divide the coefficients and subtract the exponents. The answer is always expressed in scientific notation. Note that the coefficient should always be between 1 and 10. For example,

$$\begin{aligned}(4.73 \times 10^5 \text{ m})(5.82 \times 10^7 \text{ m}) &= 27.5 \times 10^{12} \text{ m}^2 \\ &= 2.75 \times 10^{13} \text{ m}^2\end{aligned}$$

$$\frac{(6.4 \times 10^6 \text{ m})}{(2.2 \times 10^3 \text{ s})} = 2.9 \times 10^3 \text{ m/s}$$

When evaluating exponents, the following rules apply:

$$x^a \cdot x^b = x^{a+b} \qquad (xy)^b = x^b y^b$$

$$\frac{x^a}{x^b} = x^{a-b} \qquad \left(\frac{x}{y}\right)^b = \frac{x^b}{y^b}$$

$$(x^a)^b = x^{ab}$$

SCIENTIFIC NOTATION WITH CALCULATORS

On many calculators, scientific notation is entered using a special key, labelled EXP or EE. This key includes "$\times$ 10" from the scientific notation; you need to enter only the exponent. For example, to enter

7.5×10^4 press 7.5 EXP 4

3.6×10^{-3} press 3.6 EXP +/− 3

Depending on the type of calculator you have, +/− may need to be entered after the relevant number.

A6.2 Uncertainty in Measurements

There are two types of quantities that are used in science: exact values and measurements. Exact values include defined quantities (1 m = 100 cm) and counted values (5 beakers or 10 trials). Measurements, however, are not exact because there is some uncertainty or error associated with every measurement.

PRECISION AND ACCURACY

"Precision" and "accuracy" are terms used to describe how close a measurement is to a true value. The precision of a measurement depends upon the gradations of the measuring device. Precision is the place value of the last measurable digit. For example, a measurement of 12.74 cm is more precise than a measurement of 127.4 cm because the first value was measured to hundredths of a centimetre, whereas the latter was measured only to tenths of a centimetre.

No matter how precise a measurement is, it still may not be accurate. Accuracy refers to how close a value is to its true, or accepted, value. An accurate measurement has a low uncertainty. **Figure 1** shows an analogy between precision and accuracy: the positions of darts on a dartboard.

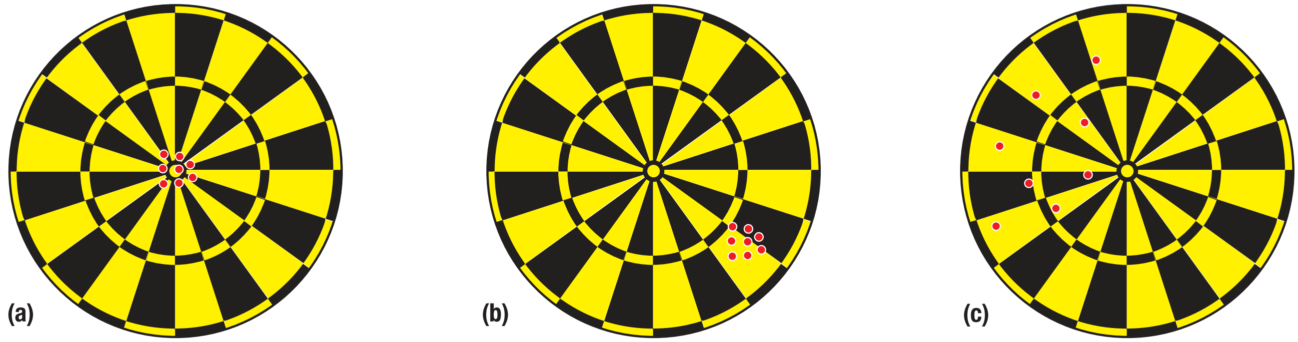


Figure 1 The positions of the darts in these diagrams represent measured or calculated results in a laboratory setting. In (a) the results are precise and accurate, in (b) they are precise but not accurate, and in (c) they are neither precise nor accurate.

How certain you are about a measurement depends on two factors: the precision of the instrument used and the size of the measured quantity. More precise instruments give more certain values. For example, a mass measurement of 13 g is less precise than a measurement of 13.12 g; you are more certain about the second measurement than the first. Certainty also depends on the size of the measurement. For example, consider the measurements 0.4 cm and 15.9 cm; both have the same precision. However, if the measuring instrument is precise to ± 0.1 cm, the first measurement is 0.4 ± 0.1 cm (0.3 cm or 0.5 cm) for an error of 25 %, whereas the second measurement could be 15.9 ± 0.1 cm (15.8 cm or 16.0 cm) for an error of 0.6 %. For both factors—the precision of the instrument used and the value of the measured quantity—the more digits there are in a measurement, the more certain you are about the measurement.

ROUNDING

When adding or subtracting measurements of different precisions, the answer is rounded to the same precision as the least precise measurement. For example, using a calculator,

$$11.7 \text{ cm} + 3.29 \text{ cm} + 0.542 \text{ cm} = 15.532 \text{ cm}$$

The answer must be rounded to 15.5 cm because the first measurement limits the precision to a tenth of a centimetre.

Follow these rules to round answers to calculations:

1. When the first digit to be dropped is 4 or less, the last digit retained should not be changed.
   3.141 326 rounded to 4 digits is 3.141
2. When the first digit to be dropped is greater than 5, or if it is a 5 followed by at least one digit other than zero, the last digit retained is increased by 1 unit.
   2.221 682 rounded to 5 digits is 2.2217
   4.168 501 rounded to 4 digits is 4.169
3. When the first digit discarded is 5 followed by only zeros, the last digit retained is increased by 1 if it is odd, but not changed if it is even.
   2.35 rounded to 2 digits is 2.4
   2.45 rounded to 2 digits is 2.4
   −6.35 rounded to 2 digits is −6.4

SIGNIFICANT DIGITS

The certainty of any measurement is communicated by the number of significant digits in the measurement. In a measured or calculated value, significant digits are the digits that are known reliably, or for certain, and include the last digit that is estimated or uncertain. Significant digits include all digits correctly reported from a measurement.

Follow these rules to decide if a digit is significant:

1. All non-zero digits are significant.
2. If a decimal point is present, zeros to the left of other digits (leading zeros) are not significant.
3. If a decimal point is not present, zeros to the right of the last non-zero digit (trailing zeros) are not significant.
4. Zeros placed between other digits are always significant.
5. Zeros placed after other digits to the right of a decimal point are significant.
6. When a measurement is written in scientific notation, all digits in the coefficient are significant.
7. Counted and defined values have infinite significant digits.

Table 2 shows examples of significant digits.

Table 2 Certainty in Significant Digits

Measurement	Number of significant digits
32.07 m	4
0.0041 g	2
5×10^5 kg	1
7002 N • m	4
6400 s	2
6.0000 A	5
204.0 cm	4
10.0 kJ	3
100 people (counted)	infinite

An answer obtained by multiplying and/or dividing measurements is rounded to the same number of significant digits as the measurement with the fewest significant digits. For example, using a calculator to solve the following equation,

$$77.8 \text{ km/h} \times 0.8967 \text{ h} = 69.76326 \text{ km}$$

However, the certainty of the answer is limited to three significant digits, so the answer is rounded up to 69.8 km. The same applies to scientific notation. For example,

$$(5.5 \times 10^4) + (4.236 \times 10^4) = 9.7 \times 10^4$$

MEASUREMENT ERROR

There are two types of measurement error: random error and systematic error. Random error results when an estimate is made to obtain the last significant digit for any measurement. The size of the random error is determined by the precision of the measuring instrument. For example, when measuring length with a measuring tape, it is necessary to estimate between the marks on the measuring tape. If these marks are 1 cm apart, the random error will be greater and the precision will be less than if the marks are 1 mm apart.

Such errors can be reduced by taking the average of several readings.

Systematic error is associated with an inherent problem with the measuring system, such as the presence of an interfering substance, incorrect calibration, or room conditions. For example, if a balance is not zeroed at the beginning, all measurements will have a systematic error; using a slightly worn metre stick will also introduce a systematic error.

REPORTING DATA INVOLVING MEASUREMENTS

A formal report of an experiment involving measurements should include an analysis of uncertainty, percentage uncertainty, and percentage error or percentage difference. Uncertainty is often assumed to be plus or minus half of the smallest division of the scale on the instrument; for example, the estimated uncertainty of 15.8 cm is $\pm$ 0.05 cm or $\pm$ 0.5 mm.

Whenever calculations involving addition or subtraction are performed, the uncertainties accumulate. Thus, to find the total uncertainty, the individual uncertainties must be added. For example,

$$(34.7 \text{ cm} \pm 0.05 \text{ cm}) - (18.4 \text{ cm} \pm 0.05 \text{ cm}) = 16.3 \text{ cm} \pm 0.10 \text{ cm}$$

Percentage uncertainty is calculated by dividing the uncertainty by the measured quantity and multiplying by 100. Use your calculator to prove that 28.0 cm $\pm$ 0.05 cm has a percentage uncertainty of $\pm$ 0.18 %.

Whenever calculations involving multiplication or division are performed, the percentage uncertainties must be added. If desired, the total percentage uncertainty can be converted back to uncertainty. For example, consider the area of a certain rectangle:

$$\begin{aligned} A &= lw \\ &= (28.0 \text{ cm} \pm 0.18\ \%)(21.5 \text{ cm} \pm 0.23\ \%) \\ &= 602 \text{ cm}^2 \pm 0.41\ \% \\ A &= 602 \text{ cm}^2 \pm 2.5 \text{ cm}^2 \end{aligned}$$

Percentage error can be determined only if it is possible to compare a measured value with that of the most commonly accepted value. The equation is

$$\%\ \text{error} = \frac{\text{measured value} - \text{accepted value}}{\text{accepted value}} \times 100$$

Percentage difference is useful for comparing two measurements when the true measurement is not known or for comparing a measured value to a predicted value. The percentage difference is calculated as

$$\%\ \text{difference} = \frac{\text{measured value} - \text{predicted value}}{\text{predicted value}} \times 100$$

A6.3 Use of Units

When solving problems in science it is important to denote the units that go with a numerical value. The measurement 170 is unacceptable since there are no units. The measurement 170 g/mL denotes a density, 170 °C denotes a temperature, 170 K denotes a temperature in kelvins, and 170 kPa denotes a pressure. Understanding and placing units with a value gives the proper context of the value.

You can also identify a formula by looking at the units. For instance, if a density of 9.01 g/cm^3 is given, you can note that the density units have grams (mass) divided by cm^3 (volume), so the formula for density is mass divided by volume or $d = \frac{m}{V}$.

A6.4 Graphing

There are many types of graphs that you can use to organize your data. You need to identify which type of graph is best for your data before you begin graphing. Three of the most useful kinds are bar graphs, circle (pie) graphs, and point-and-line graphs. When both variables are quantitative, use a point-and-line graph. The following guidelines show how to construct a point-and-line graph from the data in **Table 3**.

Table 3 Experimental Data for a Sample of Argon

Temperature (°C)	Volume (mL)
11	95.6
25	100.0
47	107.4
73	116.1
159	145.0
233	169.8
258	178.1

1. Use graph paper and construct your graph on a grid. The horizontal edge on the bottom of this grid is the x-axis and the vertical edge on the left is the y-axis. Do not be too thrifty with graph paper—a larger graph is easier to interpret.
2. Decide which variable goes on which axis and label each axis, including the units of measurement. The manipulated (independent) variable is generally plotted along the x-axis and the responding (dependent) variable along the y-axis.
3. Title your graph. The title should be a concise description of the data contained in the graph.
4. Determine the range of values for each variable. The range is the difference between the largest and smallest values. Graphs often include extra length on each axis, to make them appear less cramped.
5. Choose a scale for each axis. This will depend on how much space you have and the range of values for each axis. Each line on the grid usually increases steadily in value by a convenient number, such as 1, 2, 5, 10, 50, or 100.
6. Plot the points. Start with the first pair of values, which may or may not be at the origin of the graph.

7. After you have plotted and checked all the points, draw a line through them to show the relationship between the variables, if possible. Not all points may lie exactly on a line; small errors in each measurement may have occurred, causing the data points to move away from the perfect line. Draw a line that comes closest to most of the points. This is called the line of best fit—a smooth line that passes through or between the points so that there are about the same number of points on each side of the line. The line of best fit may be straight or curved (**Figure 2**). Graphs often use different colours or symbols to indicate the different sets, and include a legend. In some cases, it might be more appropriate to "join the dots" when graphing values that are counted rather than measured (for example, when plotting the counted number of bubbles produced rather than the measured volume of gas produced).

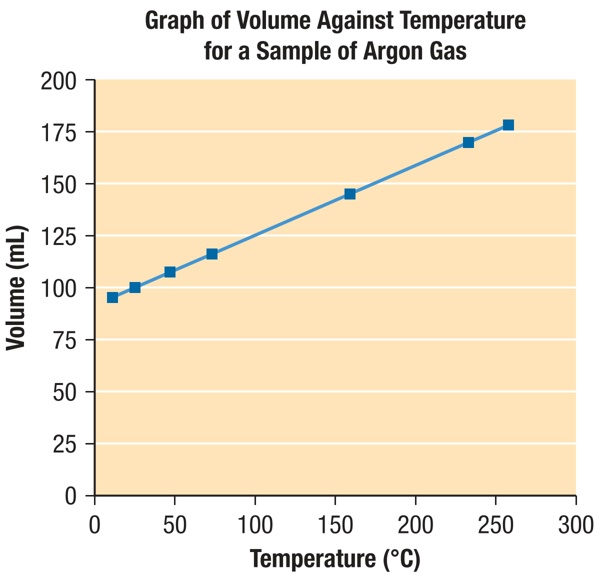


Figure 2 A point-and-line graph

A6.5 Problem Solving in Chemistry

Chemists are problem solvers who often need to determine the quantities of reactants and products in chemical reactions. These quantities frequently have to be determined indirectly, from other measurements. Two particularly useful methods for solving chemistry problems are the factor-label method (also known as dimensional analysis) and the formula method.

Tutorial 1 Problem-Solving Strategies

The factor-label method is used to solve all of the stoichiometry sample problems in this textbook. The formula method is used only in certain circumstances. The following sample problems explain and demonstrate these methods.

CASE 1: USING THE FACTOR-LABEL METHOD

Sample Problem 1: Finding the Amount Corresponding to a Given Mass

Copper metal reacts with silver nitrate solution to produce pure silver and copper(II) nitrate according to the following equation:

$$Cu(s) + 2\,AgNO_3(aq) \rightarrow 2\,Ag(s) + Cu(NO_3)_2(aq)$$

What amount of copper, n_{Cu}, is used in this reaction if 50.00 g of copper reacts with excess silver nitrate?

When solving problems using the factor-label method, you always begin by writing the required (unknown) value on the left of an equal symbol and the given value on the right:

required value = given value . . .

In this problem, the unknown value is n_{Cu} and the given value is 50.00 g Cu. Therefore, we begin the calculation by relating the two quantities, as follows:

$$n_{Cu} = 50.00\text{ g Cu} \ldots$$

Note that this is only a starting point in the calculation, not a true equation yet. Our goal here is to determine the value of the quantity on the left side of the equation (the value we are looking for) by performing calculations on the quantity on the right side of the equation (the quantity we are given). We do this by multiplying the quantity on the right side (50.00 g Cu) by a ratio that relates this quantity to the quantity on the left (n_{Cu}). We call this ratio a conversion factor. Thus,

$$n_{Cu} = 50.00\text{ g Cu} \times [\text{conversion factor}]$$

In this case, we need a conversion factor that relates mass of Cu (in grams) to amount of Cu (in moles). Ask yourself, "Is there a particular amount of copper that has a particular mass?" Yes, the molar mass of copper tells us that 1 mol of Cu has a mass of 63.55 g. We may represent this relationship in the form of a molar mass ratio, as follows:

$$\frac{1\text{ mol Cu}}{63.55\text{ g}} \quad \text{or} \quad \frac{63.55\text{ g}}{1\text{ mol Cu}}$$

Either of these forms of the ratio may be used as a conversion factor. However, in the factor-label method, we select the conversion factor that results in the quantity we are looking for, n_{Cu}, when it is multiplied by the "given" quantity (50.00 g Cu). In this case, we use the conversion factor $\frac{1\text{ mol Cu}}{63.55\text{ g}}$ since

$$n_{Cu} = 50.00\text{ }\cancel{\text{g Cu}} \times \frac{1\text{ mol Cu}}{63.55\text{ }\cancel{\text{g Cu}}}$$

$$n_{Cu} = 0.787\text{ mol Cu}$$

The amount of copper used in the reaction is 0.787 mol.

Notice that, while we could have tried to use the other conversion factor $\left(\frac{63.55\text{ g}}{1\text{ mol Cu}}\right)$, the calculation would not have resulted in a value for the desired quantity since common units do not cancel.

$$\cancel{n_{Cu} = 50.00\text{ g Cu} \times \frac{63.55\text{ g Cu}}{1\text{ mol Cu}}}$$

Sample Problem 2: Finding the Amount of a Different Substance, Using a Given Mass

Copper metal reacts with silver nitrate solution to produce pure silver and copper(II) nitrate according to the following equation:

$$Cu(s) + 2\,AgNO_3(aq) \rightarrow 2\,Ag(s) + Cu(NO_3)_2(aq)$$

What amount of silver, n_{Ag}, is produced in the reaction if 50.00 g of copper reacts with excess silver nitrate?

Unlike in Sample Problem 1, this is a stoichiometry problem in which we use a balanced chemical equation to determine the quantities of reactants and products. Therefore, we begin by writing the balanced equation, listing given (known) and required (unknown) values immediately below their respective entities in the equation:

$$Cu(s) + 2\,AgNO_3(aq) \rightarrow 2\,Ag(s) + Cu(NO_3)_2(aq)$$

$$\text{50.00 g Cu} \qquad\qquad n_{Ag}$$

Again, we begin by relating the unknown quantity, n_{Ag}, to the quantity given in the question, 50.00 g Cu, as follows:

$$n_{Ag} = 50.00\text{ g Cu} \ldots$$

However, unlike in Sample Problem 1, in which we calculated the amount of copper in 50.00 g of copper, here we are asked to determine the amount of silver formed when 50.00 g of copper reacts with silver nitrate, $AgNO_3(aq)$. To do this, we must use the mole ratio in the balanced equation between copper and silver. The ratio is

$$1\text{ mol Cu} : 2\text{ mol Ag}$$

As with all other ratios, we can represent this ratio in fraction form as

$$\frac{1\text{ mol Cu}}{2\text{ mol Ag}} \text{ or } \frac{2\text{ mol Ag}}{1\text{ mol Cu}}$$

We can use either of these ratios as conversion factors. However, you will notice that neither conversion factor results in the desired quantity, n_{Ag}, when multiplied to the given quantity, 50.0 g Cu:

$$\cancel{n_{Ag} = 50.00\text{ g Cu} \times \frac{1\text{ mol Cu}}{2\text{ mol Ag}}}$$

$$\cancel{n_{Ag} = 50.00\text{ g Cu} \times \frac{2\text{ mol Ag}}{1\text{ mol Cu}}}$$

However, if we convert the *mass* of copper to *amount* of copper first, we may then use one of the mole conversion factors to determine the amount of silver produced. To convert the mass of copper into amount of copper, multiply the mass of copper by the molar mass conversion factor (as in Sample Problem 1). Then multiply the result by the mole ratio:

$$n_{Ag} = 50.00\,\cancel{\text{g Cu}} \times \frac{1\,\cancel{\text{mol Cu}}}{63.55\,\cancel{\text{g Cu}}} \times \frac{2\text{ mol Ag}}{1\,\cancel{\text{mol Cu}}}$$

$$n_{Ag} = 1.574\text{ mol Ag}$$

Thus, 1.574 mol of silver metal is formed in the reaction.

Two conversion factors were used in this calculation. The first factor converts mass of Cu into amount of Cu and the second factor converts amount of Cu to amount of Ag.

Sample Problem 3: Finding the Mass of a Different Substance, Using a Given Mass

Copper metal reacts with silver nitrate solution to produce pure silver and copper(II) nitrate according to the following equation:

$$Cu(s) + 2\,AgNO_3(aq) \rightarrow 2\,Ag(s) + Cu(NO_3)_2(aq)$$

What mass of silver, m_{Ag}, is produced in the reaction if 50.00 g of copper reacts with excess silver nitrate?

In this problem we are asked to determine the mass of silver, m_{Ag}, produced when 50.00 g of copper reacts. As usual, we begin by relating the quantity we are asked to find, m_{Ag}, to the quantity given in the question, 50.00 g Cu:

$$m_{Ag} = 50.00\text{ g Cu} \ldots$$

We then multiply the quantity on the right (the given quantity) by as many conversion factors as are needed to determine the value of the required quantity, m_{Ag}. As in Sample Problem 2, we need to use a mole ratio from the balanced equation to convert amount of copper into amount of silver:

$$m_{Ag} = 50.00\,\cancel{\text{g Cu}} \times \frac{1\,\cancel{\text{mol Cu}}}{63.55\,\cancel{\text{g Cu}}} \times \frac{2\,\cancel{\text{mol Ag}}}{1\,\cancel{\text{mol Cu}}} \times \frac{107.87\text{ g Ag}}{1\,\cancel{\text{mol Ag}}}$$

$$m_{Ag} = 169.7\text{ g Ag}$$

Thus, 169.7 g of silver metal is formed in the reaction.

Notice that three different conversion factors are used to solve Sample Problem 3. The first conversion factor is the molar mass ratio for copper. This factor converts the given mass of copper into amount of copper. The second conversion factor is the mole ratio between copper and silver, obtained from the balanced equation. This factor converts the amount of copper into amount of silver. The third conversion factor is the molar mass ratio for silver. This factor converts amount of silver into mass of silver, the desired quantity.

Notice that in the first factor, the molar mass ratio of copper is written with amount in the numerator and mass in the denominator, but in the third factor, the molar mass ratio of silver is written with mass in the numerator and amount in the denominator. Remember that all conversion factors are ratios and thus may be written as reciprocals. For example, the molar mass ratio of silver may be used in the form

$$\frac{107.87 \text{ g Ag}}{1 \text{ mol Ag}} \quad \text{or} \quad \frac{1 \text{ mol Ag}}{07.87 \text{ g Ag}}$$

Once you have identified a necessary conversion factor, use it in the form that results in the desired quantity conversion.

The factor-label method provides an effective and efficient strategy for solving many problems in chemistry, especially stoichiometry problems. The strategy always involves multiplying a quantity given in the question by a series of ratios (conversion factors) that have the effect of determining the value of the desired quantity:

Desired quantity = given quantity × conversion factor × conversion factor × conversion factor . . .

CASE 2: USING THE FORMULA METHOD

You can use the formula method *only* if a known mathematical equation can be applied to the problem. Of the three sample problems solved above, using the factor-label method, only the first can be solved using the formula method because this is the only problem for which we have a mathematical equation. Sample Problem 4 shows how this problem-solving strategy is used.

Sample Problem 4: Finding the Amount Corresponding to a Given Mass

Copper metal reacts with silver nitrate solution to produce pure silver and copper(II) nitrate according to the following equation:

$$Cu(s) + 2\ AgNO_3(aq) \rightarrow 2\ Ag(s) + Cu(NO_3)_2(aq)$$

In a specific investigation, 50.00 g of copper reacts with excess silver nitrate to produce silver and copper(II) nitrate. What amount of copper, n_{Cu}, is used in the reaction?

Use the general formula for determining the amount of substance, n, from the mass of the substance, m, and the molar mass of the substance, M:

$$n = \frac{m}{M}$$

We add the subscript "Cu" to indicate that the variables refer to copper.

$$\begin{aligned} n_{Cu} &= \frac{m_{Cu}}{M_{Cu}} \\ &= \frac{50.0 \text{ g}}{63.55 \text{ g/mol}} \\ n_{Cu} &= 0.787 \text{ mol} \end{aligned}$$

Thus, 0.787 mol of copper metal, Cu(s), is used in the reaction.

Compare this calculation to the calculation performed in Sample Problem 1 (using the factor-label method). The approach is different, but the result is the same.

A7 Choosing Appropriate Career Pathways

Often, one of the most difficult tasks in high school is deciding what career path to follow after graduation. The science skills and concepts presented in this book will be of benefit to many careers, whether you are planning a career in scientific research (such as research geneticist or astrophysicist) or in areas related to science (such as environmental lawyer, pharmaceutical sales rep, or electrician). The strong critical-thinking and problem-solving skills that are emphasized in science programs are a valuable asset for any career.

Career Links and Pathways

Throughout this textbook you will have many opportunities to explore careers related to your studies in Chemistry. The Career Links icons found in the margins indicate that you can learn more about these careers on the Nelson Science website. At the end of each chapter you will also find a Career Pathways feature that illustrates sample educational pathways for some of the careers mentioned.

It is wise to begin researching academic requirements as early as possible. Understanding the options available to pursue a particular career will help you make decisions on whether to attend university or college, and which program of study you should take. In addition, understanding the terminology used by universities and colleges will play an integral role in planning your future.

University and College Programs

Undergraduate university programs generally lead to a three-year general bachelor degree or a four-year honours bachelor degree. These degree designations begin with a "B" followed by the area of specialization; for example, a B.Sc. (Hons.) indicates an Honours Bachelor of Science degree. These degrees can lead to employment or to further education in post-graduate programs at the Masters or Doctoral levels. The length of post-graduate degrees generally varies from one to four years.

College programs typically fall into three categories: one-year certificates, two-year diplomas, and three-year advanced diplomas. Certificates and diplomas can lead directly to employment opportunities or to graduate certificate programs. In some programs, there are "transfer agreements" with universities, which allow college graduates to enter university programs with advanced standing toward a university degree.

Pathways in Chemistry

The Career Pathways graphic organizer illustrates possible pathways to follow after high school. Certain pathways lead to careers via university while others may lead to careers via college. Look at **Figure 1** below. The pathways of three students are shown. Student A wishes to become a research scientist and must complete the Grade 11 and 12 University Chemistry courses (along with other prerequisites) and enter an undergraduate university program. Student A must obtain a Bachelor of Science degree, and then continue on to further education in Masters and Doctorate programs before becoming a research scientist.

Student B wishes to become a food scientist and must complete the Grade 11 University Chemistry course (along with other prerequisites) and enter an undergraduate university program. Student B must obtain a Bachelor of Science degree in Food Science or a related degree such as Biochemistry before becoming a food scientist.

Student C wishes to become a dental hygienist, and must complete Grade 12 University or College Preparation Chemistry followed by a Diploma course in Dental Hygiene.

Planning for Your Future

Planning ahead for your educational and career paths will provide a rewarding future. You should consult your guidance counsellors for specific advice on career planning and which courses you should take in high school. Take the time to research university and college websites for specific program information as these sites will provide the prerequisite information and, most often, career planning advice. While it may seem overwhelming at times, utilizing as many resources as possible will help alleviate some of the stress in planning your future.

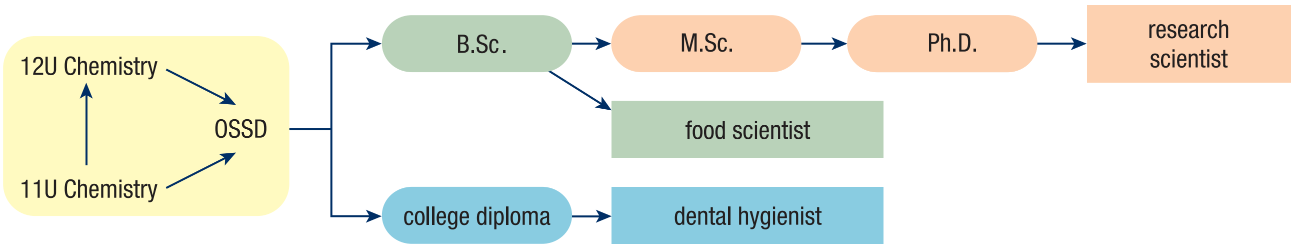


Figure 1 This graphic organizer shows pathways to several careers that involve chemistry. The blue boxes indicate careers that require a college diploma. Green boxes indicate careers following from undergraduate degrees. Orange boxes indicate a career pathway that involves one or more post-graduate university degrees.

B1 The Periodic Table

Nelson Chemistry 11

Key

atomic number	26
electronegativity	1.8
most common ion charge	3+
other ion charge	2+
symbol of element (solids in black, liquids in blue, gases in red)	Fe
name of element	iron
atomic mass (u)—based on C-12 atomic molar mass (g/mol)	55.85

Period	1	2	3	4	5	6	7	8	9
1	1 (1+, 1−) 2.2 **H** hydrogen 1.01								
2	3 (1+) 1.0 **Li** lithium 6.94	4 (2+) 1.6 **Be** beryllium 9.01							
3	11 (1+) 0.9 **Na** sodium 22.99	12 (2+) 1.3 **Mg** magnesium 24.31							
4	19 (1+) 0.8 **K** potassium 39.10	20 (2+) 1.0 **Ca** calcium 40.08	21 (3+) 1.4 **Sc** scandium 44.96	22 (4+, 3+) 1.5 **Ti** titanium 47.87	23 (5+, 4+) 1.6 **V** vanadium 50.94	24 (3+, 2+) 1.7 **Cr** chromium 52.00	25 (2+, 4+) 1.6 **Mn** manganese 54.94	26 (3+, 2+) 1.8 **Fe** iron 55.85	27 (2+, 3+) 1.9 **Co** cobalt 58.93
5	37 (1+) 0.8 **Rb** rubidium 85.47	38 (2+) 1.00 **Sr** strontium 87.62	39 (3+) 1.2 **Y** yttrium 88.91	40 (4+) 1.3 **Zr** zirconium 91.22	41 (5+, 3+) 1.6 **Nb** niobium 92.91	42 (6+) 2.2 **Mo** molybdenum 95.96	43 (7+) 1.9 **Tc** technetium (98)	44 (3+, 4+) 2.2 **Ru** ruthenium 101.07	45 (3+) 2.3 **Rh** rhodium 102.91
6	55 (1+) 0.8 **Cs** cesium 132.91	56 (2+) 0.9 **Ba** barium 137.33	57 (3+, 2+) 1.1 **La** lanthanum 138.91	72 (4+) 1.3 **Hf** hafnium 178.49	73 (5+) 1.5 **Ta** tantalum 180.95	74 (6+) 2.4 **W** tungsten 183.84	75 (7+) 1.9 **Re** rhenium 186.21	76 (4+) 2.2 **Os** osmium 190.23	77 (4+) 2.2 **Ir** iridium 192.22
7	87 (1+) 0.7 **Fr** francium (223)	88 (2+) 0.9 **Ra** radium (226)	89 (3+, 2+) 1.1 **Ac** actinium (227)	104 — **Rf** rutherfordium (261)	105 — **Db** dubnium (262)	106 — **Sg** seaborgium (266)	107 — **Bh** bohrium (264)	108 — **Hs** hassium (277)	109 — **Mt** meitnerium (268)
	Alkali metals	Alkaline earth metals							

Period					
6	58 (3+) 1.1 **Ce** cerium 140.12	59 (3+) 1.1 **Pr** praseodymium 140.91	60 (3+) 1.1 **Nd** neodymium 144.24	61 (3+) — **Pm** promethium (145)	62 (3+, 2+) 1.1 **Sm** samarium 150.36
7	90 (4+) 1.3 **Th** thorium 232.04	91 (5+, 4+) 1.5 **Pa** protactinium 231.04	92 (6+, 4+) 1.4 **U** uranium 238.03	93 (5+) 1.4 **Np** neptunium (237)	94 (4+, 6+) 1.3 **Pu** plutonium (244)

- Metals
- Metalloids
- Non-metals
- Hydrogen

Periodic Table of the Elements

Measured values are subject to change as experimental techniques improve. Atomic molar mass values in this table are based on IUPAC website values (2005 and 2007).

10	11	12	13	14	15	16	17	18	
								2 — — He helium 4.00	1
			5 — 2.0 B boron 10.81	6 — 2.6 C carbon 12.01	7 3− 3.0 N nitrogen 14.01	8 2− 3.4 O oxygen 16.00	9 1− 4.0 F fluorine 19.00	10 — — Ne neon 20.18	2
			13 3+ 1.6 Al aluminum 26.98	14 — 1.9 Si silicon 28.09	15 3− 2.2 P phosphorus 30.97	16 2− 2.6 S sulfur 32.07	17 1− 3.2 Cl chlorine 35.45	18 — — Ar argon 39.95	3
28 2+ 3+ 1.9 Ni nickel 58.69	29 2+ 1+ 1.9 Cu copper 63.55	30 2+ 1.7 Zn zinc 65.41	31 3+ 1.8 Ga gallium 69.72	32 4+ 2.0 Ge germanium 72.64	33 3− 2.2 As arsenic 74.92	34 2− 2.6 Se selenium 78.96	35 1− 3.0 Br bromine 79.90	36 — 3.0 Kr krypton 83.80	4
46 2+ 3+ 2.2 Pd palladium 106.42	47 1+ 1.9 Ag silver 107.87	48 2+ 1.7 Cd cadmium 112.41	49 3+ 1.8 In indium 114.82	50 4+ 2+ 2.0 Sn tin 118.71	51 3+ 5+ 2.1 Sb antimony 121.76	52 2− 2.1 Te tellurium 127.60	53 1− 2.7 I iodine 126.90	54 — 2.6 Xe xenon 131.29	5
78 4+ 2+ 2.3 Pt platinum 195.08	79 3+ 1+ 2.5 Au gold 196.97	80 2+ 1+ 2.0 Hg mercury 200.59	81 1+ 3+ 1.6 Tl thallium 204.38	82 2+ 4+ 2.3 Pb lead 207.2	83 3+ 5+ 2.0 Bi bismuth 208.98	84 2+ 4+ 2.0 Po polonium (209)	85 1− 2.2 At astatine (210)	86 — — Rn radon (222)	6
110 — Ds darmstadtium (281)	111 — Rg roentgenium (272)	112 — Cn copernicium (285)	113 — Uut ununtrium (284)	114 — Uuq ununquadium (289)	115 — Uup ununpentium (288)	116 — Uuh ununhexium (291)	117 — Uus ununseptium	118 — Uuo ununoctium (294)	7
							Halogens	Noble gases	

63 3+ 2+ — Eu europium 151.96	64 3+ 1.2 Gd gadolinium 157.25	65 3+ — Tb terbium 158.93	66 3+ 1.2 Dy dysprosium 162.50	67 3+ 1.2 Ho holmium 164.93	68 3+ 1.2 Er erbium 167.26	69 3+ 1.3 Tm thulium 168.93	70 3+ 2+ — Yb ytterbium 173.05	71 2+ 1.3 Lu lutetium 174.97	6
95 3+ 4+ 1.3 Am americium (243)	96 3+ 1.3 Cm curium (247)	97 3+ 4+ 1.3 Bk berkelium (247)	98 3+ 1.3 Cf californium (251)	99 3+ 1.3 Es einsteinium (252)	100 3+ 1.3 Fm fermium (257)	101 2+ 3+ 1.3 Md mendelevium (258)	102 2+ 3+ 1.3 No nobelium (259)	103 3+ — Lr lawrencium (262)	7

B2 Units, Symbols, Quantities, and Prefixes

Throughout *Nelson Chemistry 11* and in this reference section, we have attempted to be consistent in the presentation and usage of units. As far as possible, the text uses the International System of Units (SI). However, some other units have been included because of their practical importance, wide usage, or use in specialized fields. *Nelson Chemistry 11* has followed the most recent Canadian Metric Practice Guide (CAN/CSA-Z234.1-00), published in 2000 and updated in 2003 by the Canadian Standards Association.

Table 1 SI Base Units

Quantity	Symbol	Unit	Symbol
amount of substance	n	mole	mol
electric current	I	ampere	A
length	L, l, h, d, w	metre	m
luminous intensity	I_v	candela	cd
mass	m	kilogram	kg
temperature	T	kelvin	K
time	t	second	s

Table 2 Some SI-Derived Units

Quantity	Symbol	Unit	Symbol	SI base unit
pressure	P	newton per square metre	N/m^2	kg/m·s^2
amount concentration	c	mole per litre	mol/L	kmol/m^3
volume	V	cubic metre	m^3	m^3

Table 3 Defined and Measured Quantities

Defined (exact) quantities	Measured (uncertain) quantities
1 t = 1000 kg = 1Mg STP = 0 °C and 101.325 kPa (use 0 °C and 101 kPa) SATP = 25 °C and 100 kPa 0 °C = 273.15 K (use 273 K) 1 atm = 101.325 kPa (use 101 kPa) 1 atm = 760 mm Hg 1 bar = 100 kPa	R = 8.314 kPa·L/(mol·K) V_{STP} = 22.4 L/mol V_{SATP} = 24.8 L/mol N_A = 6.02 × 10^{23} entities/mol

Table 4 Stoichiometry Symbols and Units

Quantity	Quantity symbol	Unit(s)
amount of substance	n	mol
mass of substance	m	g or kg
amount concentration of solution	c	mol/L
volume of solution or gas	V	mL or L
molar mass of substance	M	—
molar volume of gas	—	L

Table 5 Numerical Prefixes

Powers and subpowers of ten		
Prefix	**Power**	**Symbol**
deca	10^1	da
hecto	10^2	h
kilo	10^3	k
mega	10^6	M
giga	10^9	G
tera	10^{12}	T
peta	10^{15}	P
deci	10^{-1}	d
centi	10^{-2}	c
milli	10^{-3}	m
micro	10^{-6}	μ
nano	10^{-9}	n
pico	10^{-12}	p
femto	10^{-15}	f
atto	10^{-18}	a

Some Examples of Prefix Use

0.00350 L = 3.50 × 10^{-3} L = 3.50 mL
0.27 m = 27 × 10^{-2} m = 27 cm
3 000 000 000 Hz = 3 × 10^9 Hz = 3 GHz

Table 6 Greek Letters Used in Chemistry

A	α	alpha
B	β	beta
Γ	γ	gamma
Δ	δ	delta

Table 7 Physical Constants

Quantity	Symbol	Approximate value
atomic mass unit	u	1.661×10^{-27} kg
Avogadro's constant	N_A	$6.022\ 141\ 99 \times 10^{23}$ (use 6.02×10^{23})

LEARNING **TIP**

SI Prefixes

Sometimes it is difficult to remember the metric prefixes. A mnemonic is a saying that helps you remember something. "King Henry Doesn't Mind Drinking Chocolate Milk" is a mnemonic for kilo, hecto, deca, metre, deci, centi, and milli. Another helpful hint is that mega (M) represents a million ($\times 10^6$) and tera (T) represents a trillion ($\times 10^{12}$). The first letters of the prefix and of what it represents are the same.

B3 Elements and Compounds

Table 1 The Elements

Element	Symbol	Atomic number	Ionization energy (kJ/mol)	Electro-negativity*	Electron affinity (kJ/mol)	Ionic radius (pm)	Common ion charge
actinium	Ac	89	509	1.1		111	3+
aluminum	Al	13	578	1.6	42.5	50	3+
americium	Am	95	578	1.3		97.5	3+
antimony	Sb	51	834	2.1	100.9	76	3+
argon	Ar	18	1521	—			
arsenic	As	33	947	2.2	78	222	
astatine	At	85		2.2	[270]	227	1−
barium	Ba	56	503	0.9	[14]	135	2+
berkelium	Bk	97	601	1.3		98	3+
beryllium	Be	4	899	1.6		31	2+
bismuth	Bi	83	703	2.0	91.3	95	3+
boron	B	5	801	2.0	26.7		
bromine	Br	35	1140	3.0	324.54	196	1−
cadmium	Cd	48	868	1.7		97	2+
calcium	Ca	20	590	1.0	1.78	99	2+
californium	Cf	98	608	1.3		95	3+
carbon	C	6	1086	2.6	121.85		
cerium	Ce	58	528	1.1		102	3+
cesium	Cs	55	376	0.8	45.50	169	1+
chlorine	Cl	17	1251	3.2	348.57	181	1−
chromium	Cr	24	653	1.7	64.3	64	3+

Element	Symbol	Atomic number	Ionization energy (kJ/mol)	Electro-negativity	Electron affinity (kJ/mol)	Ionic radius (pm)	Common ion charge
cobalt	Co	27	758	1.9	63.9	74.5	2+
copper	Cu	29	745	1.9	119.2	72	2+
curium	Cm	96	581	1.3		97	3+
dysprosium	Dy	66	572	1.2		91.2	3+
einsteinium	Es	99	619	1.3		98	3+
erbium	Er	68	589	1.2		89.0	3+
europium	Eu	63	547	—		94.7	3+
fermium	Fm	100	627	1.3		97	3+
fluorine	F	9	1681	4.0	328.16	136	1−
francium	Fr	87		0.7	[44]	180	1+
gadolinium	Gd	64	592	1.2		93.8	3+
gallium	Ga	31	579	1.8	29	62.0	3+
germanium	Ge	32	762	2.0	119.0	53.0	4+
gold	Au	79	890	2.5	222.75	91	3+
hafnium	Hf	72	680	1.3	[≈0]	78	4+
helium	He	2	2372	—			
holmium	Ho	67	581	1.2		90.1	3+
hydrogen	H	1	1312	2.2	72.55	10^{-3}/154	1+/1−
indium	In	49	558	1.8	29	81	3+
iodine	I	53	1008	2.7	295.15	216	1−
iridium	Ir	77	880	2.2	151.0	64	4+
iron	Fe	26	759	1.8	14.6	64.5	3+
krypton	Kr	36	1351	3.0			
lanthanum	La	57	538	1.1	[48]	106	3+
lawrencium	Lr	103		—		94	3+
lead	Pb	82	716	2.3	35.1	120	2+
lithium	Li	3	520	1.0	59.63	68	1+
lutetium	Lu	71	524	1.3		86.1	3+
magnesium	Mg	12	738	1.3		65	2+
manganese	Mn	25	717	1.6		80	2+
mendelevium	Md	101	635	1.3		114	2+
mercury	Hg	80	1007	2.0		110	2+
molybdenum	Mo	42	685	2.2	72.2	62	6+

Element	Symbol	Atomic number	Ionization energy (kJ/mol)	Electro-negativity	Electron affinity (kJ/mol)	Ionic radius (pm)	Common ion charge
neodymium	Nd	60	530	1.1		98.3	3+
neon	Ne	10	2081	—			
neptunium	Np	93	605	1.4		75	5+
nickel	Ni	28	737	1.9	111.5	72	2+
niobium	Nb	41	664	1.6	86.2	72	5+
nitrogen	N	7	1402	3.0			
nobelium	No	102	642	1.3		110	2+
osmium	Os	76	840	2.2	[19]	65	4+
oxygen	O	8	1314	3.4	140.98	140	
palladium	Pd	46	805	2.2	54.2	86	2+
phosphorus	P	15	1012	2.2	72.03	212	
platinum	Pt	78	870	2.3	205.3	70	4+
plutonium	Pu	94	585	1.3		86	4+
polonium	Po	84	812	2.0	[183]	65	4+
potassium	K	19	419	0.8	48.38	138	1+
praseodymium	Pr	59	523	1.1		99	3+
promethium	Pm	61	535	—		97	3+
protactinium	Pa	91	568	1.5		78	5+
radium	Ra	88	509	0.9		148	2+
radon	Rn	86	1037	—			
rhenium	Re	75	760	1.9	[14]	60	7+
rhodium	Rh	45	720	2.3		75	3+
rubidium	Rb	37	403	0.8	46.88	148	1+
ruthenium	Ru	44	711	2.2	[101]	77	3+
samarium	Sm	62	543	1.1		95.8	3+
scandium	Sc	21	631	1.4	18.1	81	3+
selenium	Se	34	941	2.6	194.96	198	
silicon	Si	14	786	1.9			
silver	Ag	47	731	1.9	125.6	126	1+
sodium	Na	11	496	0.9	52.87	95	1+
strontium	Sr	38	549	1.0	4.6	113	2+
sulfur	S	16	1000	2.6	200.41	184	
tantalum	Ta	73	761	1.5	31.1	68	5+

Element	Symbol	Atomic number	Ionization energy (kJ/mol)	Electro-negativity	Electron affinity (kJ/mol)	Ionic radius (pm)	Common ion charge
technetium	Tc	43	702	1.9	[53]	58	
tellurium	Te	52	869	2.1	190.15	221	2−
terbium	Tb	65	564	—	(−48)	92.3	3+
thallium	Tl	81	589	1.6	−9	144	1+
thorium	Th	90	587	1.3		94	4+
thulium	Tm	69	596	1.3		88.0	3+
tin	Sn	50	709	2.0	107.3	71	4+
titanium	Ti	22	658	1.5	7.6	68	4+
tungsten	W	74	770	2.4	78.6	65	6+
uranium	U	92	598	1.4		73	6+
vanadium	V	23	650	1.6	50.7	59	5+
xenon	Xe	54	1170	2.6			
ytterbium	Yb	70	603	—		86.8	3+
yttrium	Y	39	616	1.2	29.6	93	3+
zinc	Zn	30	906	1.7		74.0	2+
zirconium	Zr	40	660	1.3	41.1	79	4+

Values in this table are taken from Lange's *Handbook of Chemistry* unless otherwise stated.
*Electronegativity values are taken from WebElements.com.

Table 2 Common Chemicals

Common name	Recommended name	Formula	Common use/source
acetic acid	ethanoic acid	$CH_3COOH(aq)$	vinegar
acetone	propanone	$(CH_3)_2CO(l)$	nail polish remover
acetylene	ethyne	$C_2H_2(g)$	cutting/welding torch
ASA (Aspirin)	acetylsalicylic acid	$C_6H_4COOCH_3COOH(s)$	for pain relief medication
baking soda	sodium hydrogen carbonate	$NaHCO_3(s)$	leavening agent
battery acid	sulfuric acid	$H_2SO_4(aq)$	car batteries
bleach	sodium hypochlorite	$NaClO(s)$	bleach for clothing
bluestone	copper(II) sulfate pentahydrate	$CuSO_4 \cdot 5\ H_2O(s)$	algicide/fungicide
brine	aqueous sodium chloride	$NaCl(aq)$	water-softening agent
carbon monoxide	carbon monoxide	$CO(g)$	toxic product of incomplete combustion
citric acid	2-hydroxy-1,2,3-propanetricarboxylic acid	$C_3H_4OH(COOH)_3$	in fruit and beverages
CFC	chlorofluorocarbon	$C_xCl_yF_z(l)$; e.g., $C_2Cl_2F_4(l)$	refrigerant
charcoal/graphite	carbon	$C(s)$	fuel/lead pencils
dry ice	carbon dioxide	$CO_2(g)$	"fizz" in carbonated beverages

Common name	Recommended name	Formula	Common use/source
ethylene	ethene	$C_2H_4(g)$	for polymerization
ethylene glycol	1,2-ethandiol	$C_2H_4(OH)_2(l)$	radiator antifreeze
formaldehyde	methanal	$CH_2O(g)$	preservative, solvent
freon-12	dichlorodifluoromethane	$CCl_2F_2(l)$	refrigerant
Glauber's salt	sodium sulfate decahydrate	$Na_2SO_4 \cdot 10H_2O(s)$	solar heat storage
glucose	D-glucose; dextrose	$C_6H_{12}O_6(s)$	in plants and blood
grain alcohol	ethanol (ethyl alcohol)	$C_2H_5OH(l)$	beverage alcohol
gypsum	calcium sulfate dihydrate	$CaSO_4 \cdot 2H_2O(s)$	wallboard
lime (quicklime)	calcium oxide	$CaO(s)$	masonry
limestone	calcium carbonate	$CaCO_3(s)$	chalk and building materials
lye (caustic soda)	sodium hydroxide	$NaOH(s)$	oven/drain cleaner
malachite	copper(II) hydroxide carbonate	$Cu(OH)_2 \cdot CuCO_3(s)$	copper mineral
methyl hydrate	methanol (methyl alcohol)	$CH_3OH(l)$	gas-line antifreeze
milk of magnesia	magnesium hydroxide	$Mg(OH)_2(s)$	antacid (for indigestion)
MSG	monosodium glutamate	$NaC_5H_8NO_4(s)$	flavour enhancer
muriatic acid	hydrochloric acid	$HCl(aq)$	in concrete etching
natural gas	methane	$CH_4(g)$	fuel
nitrogen dioxide	nitrogen dioxide	$NO_2(g)$	air pollutant
ozone	ozone	$O_3(g)$	atmospheric gas; ground-level pollutant
PCBs	polychlorinated biphenyls	$(C_6H_xCl_y)_2$; e.g., $(C_6H_4Cl_2)_2(l)$	in transformers
potash	potassium chloride	$KCl(s)$	fertilizer
radon	radon	$Rn(g)$	radioactive indoor air pollutant
road salt	calcium chloride or sodium chloride	$CaCl_2(s)$ or $NaCl_2(s)$	melts ice
rotten-egg gas	hydrogen sulfide	$H_2S(g)$	in natural gas
rubbing alcohol	2-propanol	$CH_3CHOHCH_3(l)$	for massage
sand (silica)	silicon dioxide	$SiO_2(s)$	in glass making
slaked lime	calcium hydroxide	$Ca(OH)_2(s)$	limewater
soda ash	sodium carbonate	$Na_2CO_3(s)$	in laundry detergents
sugar	sucrose	$C_{12}H_{22}O_{11}(s)$	sweetener
sulfur dioxide	sulfur dioxide	$SO_2(g)$	industrial air pollutant; major cause of acid precipitation
table salt	sodium chloride	$NaCl(s)$	seasoning
washing soda	sodium carbonate decahydrate	$Na_2CO_3 \cdot 10\ H_2O(s)$	water softener
vitamin C	ascorbic acid	$H_2C_6H_6O_6(s)$	vitamin
VOCs	mixture of volatile organic compounds	—	air pollutant

B4 Cations and Anions

Table 1 Common Cations

Ion	Name
H^+	hydrogen
Li^+	lithium
Na^+	sodium
K^+	potassium
Cs^+	cesium
Be^{2+}	beryllium
Mg^{2+}	magnesium
Ca^{2+}	calcium
Ba^{2+}	barium
Al^{3+}	aluminum
Ag^+	silver

Table 2 Common Anions

Ion	Name
H^-	hydride
F^-	fluoride
Cl^-	chloride
Br^-	bromide
I^-	iodide
O^{2-}	oxide
S^{2-}	sulfide
N^{3-}	nitride
P^{3-}	phosphide

Table 3 Selected Multivalent Cations

Metal	Ions	Classical names	IUPAC names
copper, Cu	Cu^+ Cu^{2+}	cuprous cupric	copper(I) copper(II)
iron, Fe	Fe^{2+} Fe^{3+}	ferrous ferric	iron(II) iron(III)
tin, Sn	Sn^{2+} Sn^{4+}	stannous stannic	tin(II) tin(IV)
lead, Pb	Pb^{2+} Pb^{4+}	plumbous plumbic	lead(II) lead(IV)
manganese, Mn	Mn^{2+} Mn^{3+} Mn^{4+} Mn^{6+} Mn^{7+}	—	manganese(II) manganese(III) manganese(IV) manganese(VI) manganese(VII)
chromium, Cr	Cr^{2+} Cr^{3+}	chromous chromic	chromium(II) chromium(III)
gold, Au	Au^+ Au^{3+}	—	gold(I) gold(III)
nickel, Ni	Ni^{2+} Ni^{3+}	—	nickel(II) nickel(III)

Table 4 Common Polyatomic Ions

Ion	Name
$C_2H_3O_2^-$	acetate
NH_4^+	ammonium
BO_3^{3-}	borate
BrO_3^-	bromate
CO_3^{2-}	carbonate
ClO_3^-	chlorate
ClO_2^-	chlorite
CrO_4^{2-}	chromate
CN^-	cyanide
$Cr_2O_7^{2-}$	dichromate
$H_2PO_4^-$	dihydrogen phosphate
$H_2PO_3^-$	dihydrogen phosphite
HCO_3^-	hydrogen carbonate (bicarbonate)
HPO_4^{2-}	hydrogen phosphate
HPO_3^{2-}	hydrogen phosphite
HSO_4^-	hydrogen sulfate (bisulfate)
HS^-	hydrogen sulfide (bisulfide)
HSO_3^-	hydrogen sulfite (bisulfite)
ClO^-, OCl^-	hypochlorite
H_3O^+	hydronium
OH^-	hydroxide
IO_3^-	iodate
NO_2^-	nitrite
NO_3^-	nitrate
$C_2O_4^{2-}$	oxalate
ClO_4^-	perchlorate
MnO_4^-	permanganate
SCN^-	thiocyanate
O_2^{2-}	peroxide
PO_4^{3-}	phosphate
SO_4^{2-}	sulfate
SO_3^{2-}	sulfite
$S_2O_3^{2-}$	thiosulfate

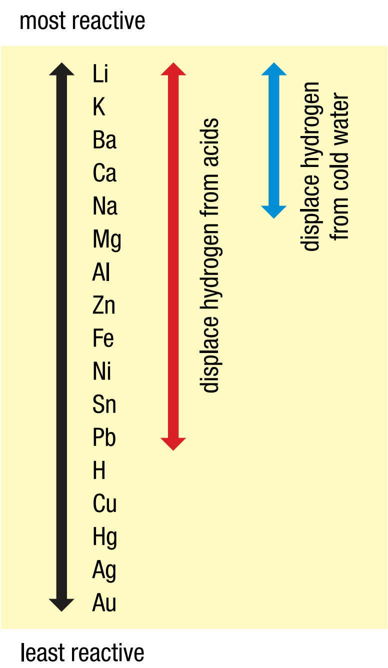
Table 5 Ion Colours

Ion in solution	Solution colour
Groups 1, 2, 17	colourless
Cr^{2+}	blue
Cr^{3+}	green
Co^{2+}	pink
Cu^{+}	green
Cu^{2+}	blue
Fe^{2+}	pale green
Fe^{3+}	yellow-brown
Mn^{2+}	pale pink
Ni^{2+}	green
CrO_4^{2-}	yellow
$Cr_2O_7^{2-}$	orange
MnO_4^{-}	purple
Ion	**Flame colour**
Li^{+}	bright red
Na^{+}	yellow
K^{+}	violet
Ca^{2+}	yellow-red
Sr^{2+}	bright red
Ba^{2+}	yellow-green
Cu^{2+}	green
Pb^{2+}	light blue-grey
Zn^{2+}	whitish green

Table 6 Solubility of Ionic Compounds at Room Temperature

Solubility	Ion	Exceptions
very soluble (aq) ≥ 0.1 mol/L	NO_3^{-}	none
	halides	except with Cu^{+}, Ag^{+}, Hg_2^{2+}, Pb^{2+}
	SO_4^{2-}	except with Ca^{2+}, Ba^{2+}, Sr^{2+}, Hg^{2+}, Pb^{2+}, Ag^{+}
	$C_2H_3O_2^{-}$	Ag^{+}
	Na^{+}, K^{+}	none
	NH_4^{+}	none
slightly soluble (s) < 0.1 mol/L	CO_3^{2-}	except with Group 1 ions and NH_4^{+}
	PO_4^{3-}	except with Group 1 ions and NH_4^{+}
	OH^{-}	except with Group 1 ions, Ca^{2+}, Ba^{2+}, Sr^{2+}
	S^{2-}	except with Groups 1 and 2 ions and NH_4^{+}

THE ACTIVITY SERIES OF METALS



B5 Naming Conventions

Table 1 Common Hydrates

	Name of hydrate	Examples
Traditional name	name of the ionic compound + Greek prefix + "water"	• copper(II) sulfate pentahydrate • magnesium sulfate heptahydrate
Alternative names	name of the ionic compound + number of water molecules + "water"	• copper(II) sulfate-5-water • magnesium sulfate-7-water
	name of the ionic compound + "water" + ratio of formula units to water molecules	• copper(II) sulfate—water (1/5) • magnesium sulfate—water (1/7)

Table 2 Prefixes Used in the Names of Hydrates and Molecular Compounds

Number of atoms or water molecules in the chemical formula	Prefix
1	mono or mon
2	di
3	tri
4	tetra
5	penta
6	hexa
7	hepta
8	octa
9	nona
10	deca

Table 3 Binary Molecular Compounds

Name of compound	Prefix
prefix + name of first element + prefix + name of second element [exception: omit "mono" for first element]	• dinitrogen monoxide • carbon dioxide

Table 4 Oxyanions in Ionic Compounds

Name of parent oxyanion		Examples
stem of the non-metal name + -*ate*		• chlorate, ClO_3^- • nitrate, NO_3^- • sulfate, SO_4^{2-} • phosphate, PO_4^{3-} • carbonate, CO_3^{2-}
Names of related oxyanions		**Examples**
If oxyanion has one more oxygen atom than the parent oxyanion	*per* + stem of non-metal name + -*ate*	• perchlorate, ClO_4^-
If oxyanion has one fewer oxygen atom than the parent oxyanion	stem of non-metal name+ -*ite*	• nitrite, NO_2^- • sulfate, SO_3^{2-}
If oxyanion has two fewer oxygen atoms than the parent oxyanion	*hypo* + stem of non-metal name + -*ite*	• hypochlorite, ClO^-

Table 5 Binary Acids and Oxyacids

Name of binary acid	Examples
hydro + stem of anion name + -*ic acid*	• hydrochloric acid, HCl • hydrocyanic acid, HCN
Name of parent oxyacid	**Examples**
If the anion name ends in -*ate*, then the acid name ends in -*ic acid*.	• sulfate ion → sulfuric acid, H_2SO_4 • acetate ion → acetic acid, $C_2H_4O_2$
Names of related oxyacids	
If the anion name starts with *per*- and ends in -*ate*, then the acid name is *per-____-ic acid*	
If the anion name ends in -*ite*, then the acid name ends in -*ous acid*	
If the anion name starts with *hypo_* and ends in -*ite*, then the acid name is *hypo____-ous acid*	

Table 6 Oxyacids

Acid name	Chemical formula	Parent oxyanion
acetic acid	$HC_2H_3BrO_2(aq)$	bromate, BrO_3^-
bromic acid	$HBrO_3(aq)$	bromate, BrO_3^-
carbonic acid	$H_2CO_3(aq)$	carbonate, CO_3^{2-}
chloric acid	$HClO_3(aq)$	chlorate, ClO_3^-
iodic acid	$HIO_3(aq)$	iodate, IO_3^-
nitric acid	$HNO_3(aq)$	nitrate, NO_3^-
phosphoric acid	$H_3PO_4(aq)$	phosphate, PO_4^{3-}
sulfuric acid	$H_2SO_4(aq)$	sulfate, SO_4^{2-}
perbromic acid	$HBrO_4(aq)$	perbromate, BrO_4^-
hypobromous	$HBrO(aq)$	hypobromite, BrO^-
perchloric acid	$HClO_4(aq)$	perchlorate, ClO_4^-
chlorous	$HClO_2(aq)$	chlorite, ClO_2^-
hypochlorous	$HClO(aq)$	hypochlorite, ClO^-
hypofluorous	$HFO(aq)$	hypofluorite, FO^-
periodic acid	$HIO_4(aq)$	periodate, ClO_4^-
hypoiodous	$HIO(aq)$	hypoiodite, IO^-
nitrous acid	$HNO_2(aq)$	nitrite, NO_2^-

B6 Summary of Reaction Types

Synthesis reaction: $A + B \rightarrow AB$
Decomposition reaction: $AB \rightarrow A + B$
Single displacement reaction: $A + BC \rightarrow AC + B$
Double displacement reaction: $AB + CD \rightarrow AD + CB$ [In an aqueous reaction at least one of the products is a precipitate.]
Combustion of a hydrocarbon: $C_xH_y + O_2(g) \rightarrow$
$CO_2(g) + H_2O(g)$ + energy [complete combustion]
Combustion of an element: $A + n\ O_2(g) \rightarrow A_xO_y$

B7 Summary of Bond Characteristics

Electronegativity difference, ΔEN	Bond type
$\Delta EN = 0$	non-polar covalent
$0 < \Delta EN < 1.7$	polar covalent
$\Delta EN \geq 1.7$	ionic

APPENDIX C ANSWERS

These pages include numerical and short answers to chapter section questions, tutorial practice questions, and Chapter Self-Quiz, Chapter Review, Unit Self-Quiz, and Unit Review questions.

Unit 1

Are You Ready? pp. 4–5

2. (a) compound (d) mixture
   (b) element (e) compound
   (c) mixture
3. (c) F: halogens; Mg: alkaline earth metals; Xe: noble gases; Na: alkali metals
4. (a) one (c) eight (except He)
   (b) seven (d) two
9. (a) v (d) ii
   (b) vi (e) i
   (c) iii (f) iv

1.1 Questions, p. 10

6. (a) International Union of Pure and Applied Chemistry

1.2 Questions, p. 16

2. Ca − 2, C − 4, Si − 4
10. 119 protons, 119 electrons, 183 neutrons

1.3 Questions, p. 22

2. (a) argon (c) neon
   (b) neon (d) neon
4. (a) oxide (d) sulfate
   (b) copper(I) (e) hydroxide
   (c) tin(IV) (f) ammonium
6. (a) NO_3^-
   (b) CO_3^{2-}
   (c) $C_2H_3O_2^-$
   (d) MnO_4^-
7. Ca^{2+}, cation; CO_3^{2-} , anion
8. Ne, Na^+, Al^{3+} and O^{2-}
9. conductivity test, flame test

1.4 Questions, p. 29

3. Ag-107
4. 28.11 u
7. Cl-35: 17 protons, 17 electrons, and 18 neutrons
   Cl-37: 17 protons, 17 electrons, and 20 neutrons
9. 39.1 u

1.5 Questions, p. 33

2. (a) chlorine (c) silicon
   (b) krypton (d) rubidium
6. (a) seven (e) eight
   (b) six (f) four
   (c) two (g) three
   (d) one
7. (a) Cs more reactive than Ba
   (b) F more reactive than C
   (c) Na more reactive than Ar
   (d) Mg more reactive than Si

1.7 Questions, p. 41

1. (a) actinium (c) radon
   (b) cesium
2. F<C<Li<K<Rb<Cs
3. magnesium atom

Chapter 1 Self-Quiz, p. 47

1. (b)	7. (a)	13. T
2. (a)	8. (d)	14. F
3. (c)	9. (c)	15. F
4. (b)	10. T	16. F
5. (d)	11. F	17. T
6. (b)	12. F	18. T

Chapter 1 Review, p. 48

1. (c)	7. (b)	13. F
2. (a)	8. (d)	14. F
3. (d)	9. (b)	15. T
4. (a)	10. T	16. F
5. (c)	11. F	17. T
6. (b)	12. T	18. F

19. (a) vi (d) ii
    (b) iii (e) i
    (c) iv (f) v
24. (a) 13 (b) 27
27. sharing, losing, or gaining electrons
36. (a) Groups 1, 2, and 13 to 18
37. (a) left side
    (b) right side
    (c) next to the staircase line
39. picometre
43. (a) the atomic mass unit, u
48. (a) ion source, analyzer, and detector
51. (a) fluorine
    (b) oxygen
63. (a) Fe^{2+}; ferrous, or iron(II), ion
    (b) Fe^{3+}; ferric, or iron(III), ion
64. (c) IO_4^-
67. 52.1 u
68. (a) K-39
69. 10.8 u
70. (a) fluorine

2.3 Questions, p. 73

2. Cs<K<Ca<Fe<Br<Cl<F
5. (a) 1.6, polar covalent
   (b) 1.8, ionic
   (c) 0, non-polar covalent
   (d) 0.6, polar covalent
   (e) 0.8, polar covalent
   (f) 2.2, ionic
6. (a) H-F (c) C-N
   (b) O-H
7. (a) non-polar (d) polar
   (b) non-polar (e) ionic
   (c) polar
8. (a) Fr-F, $\Delta EN = 3.3$

2.4 Tutorial 1 Practice, p. 75

1. (a) MgO (c) K_2O
   (b) AlF_3
2. (a) $Mg(OH)_2$ (c) $AlPO_4$
   (b) $NaHCO_3$

2.4 Tutorial 2 Practice, p. 77

1. (a) copper(II) sulfate (c) tin(IV) chloride
   (b) copper(I) chloride (d) tin(II) oxide

2.4 Tutorial 2 Practice, p. 78

2. (a) lead(IV) sulfite
   (b) lead(II) nitrate
   (c) copper(I) phosphate
   (d) iron(III) hydroxide
   (e) sodium hypochlorite
   (f) ammonium carbonate

2.4 Tutorial 3 Practice, p. 79

1. (a) calcium chloride dihydrate
   (b) $Na_2SO_4 \bullet 10H_2O$, sodium sulfate decahydrate

2.4 Tutorial 4 Practice, p. 80

1. (a) carbon tetrachloride
   (b) nitrogen dioxide
   (c) diphosphorus pentoxide
   (d) carbon tetrafluoride
2. (a) CO
   (b) SO_2
   (c) PF_5

2.4 Questions, p. 81

1. (a) LiCl (d) Al_2O_3
   (b) K_2S (e) Na_2SO_4
   (c) $FeCl_2$ (f) SnO_2
3. (a) magnesium chloride
   (b) cesium oxide
   (c) iron(II) sulfide
   (d) sodium phosphate
   (e) ammonium nitrate
   (f) aluminum sulfate
   (g) magnesium chlorate
   (h) lead(II) bromate
   (i) zinc hydrogen phosphate
   (j) sodium cyanide
4. (a) phosphorous pentachloride
   (b) dinitrogen pentoxide
   (c) carbon tetrafluoride
   (d) sulfur dioxide
5. $CoCl_2 \bullet 6H_2O$, $CoCl_2$
6. (a) PCl_3 (c) CO
   (b) CCl_4 (d) NO
7. (a) potassium hydroxide
   (b) sodium nitrite
   (c) copper(I) chloride
   (d) sodium hydroxide
   (e) calcium carbonate

Chapter 2 Self-Quiz, p. 87

1. (b)	8. (b)	15. T
2. (c)	9. (c)	16. T
3. (d)	10. F	17. F
4. (a)	11. F	18. F
5. (a)	12. T	19. F
6. (c)	13. F	20. T
7. (b)	14. F	

Chapter 2 Review, p. 88

1. (c)	8. (b)	15. T
2. (c)	9. (d)	16. F
3. (d)	10. (b)	17. F
4. (d)	11. (d)	18. T
5. (a)	12. F	19. F
6. (b)	13. F	20. T
7. (b)	14. T	21. F

22. (a) iii (e) ii
(b) iv (f) v
(c) i (g) vi
(d) vii

32. (b) X_2Y_5

37. double bonds: CS_2 and $COCl_2$; triple bonds: HCN and NO^+

38. all except bromine, Br_2

40. (a) polar covalent
(b) ionic
(c) polar covalent
(d) non-polar covalent
(e) ionic

41. (a) CuO (d) BaF_2
(b) $Al(NO_3)_3$ (e) PbO_2
(c) $MnCl_2$ (f) $Fe_2(SO_4)_3$

42. (a) CS_2 (c) Cl_2O
(b) As_2O_3 (d) Sb_2O_5

43. (a) strontium sulfide
(b) ammonium sulfate
(c) tin(II) fluoride
(d) iron(III) phosphate
(e) calcium hydroxide
(f) magnesium carbonate

44. (a) nitrogen trifluoride
(b) diboron trioxide
(c) diodine pentoxide
(d) bromine monoxide

45. $Fe(NO_3)_3 \bullet 9H_2O$

46. zinc chloride hexahydrate

55. (c) C_2D (d) formula unit

64. (a) $C_{27}H_{46}O$

3.3 Tutorial 1 Practice, p. 107

1. (a) polar (d) non-polar
(b) non-polar (e) non-polar
(c) polar (f) non-polar

3.3 Questions, p. 108

1. (a) K-Br, H-Br, O-F, C-H
(b) C-F, O-H, C-O, H-H

4. (a) symmetrical (d) symmetrical
(b) asymmetrical (e) asymmetrical
(c) symmetrical

5. (a) non-polar (d) non-polar
(b) polar (e) polar
(c) polar

6. (a) non-polar (c) non-polar
(b) polar (d) polar

3.4 Questions, p. 115

1. (a) hydrogen bonding
(b) dipole–dipole forces

2. London dispersion forces, dipole–dipole forces, hydrogen bonds

4. (a) London dispersion forces
(b) London dispersion forces, dipole–dipole forces
(c) London dispersion forces, hydrogen bonds
(d) London dispersion forces
(e) London dispersion forces, dipole–dipole forces

5. H_2, CH_4, C_3H_8, CH_3Cl, NH_3, CH_3OH, H_2O

Chapter 3 Self-Quiz, p. 127

1. (c)	8. (d)	15. F
2. (b)	9. (a)	16. T
3. (b)	10. F	17. T
4. (a)	11. F	18. T
5. (a)	12. T	19. F
6. (d)	13. F	20. F
7. (b)	14. T	

Chapter 3 Review, p. 128

1. (c)	9. (a)	17. F
2. (d)	10. (c)	18. T
3. (a)	11. (b)	19. F
4. (c)	12. T	20. F
5. (d)	13. T	21. F
6. (a)	14. F	22. T
7. (d)	15. T	
8. (a)	16. F	

23. (a) i (b) iii (c) ii

24. food, fuel, and medicines

25. glass, metal, plastic, and paper

28. (a) covalent bonds
(b) ionic bonds
(c) ionic bonds
(d) covalent bonds
(e) covalent bonds

38. (a) N–F, H–Br, Cl–Cl
(b) As–O, C–N, C–S

39. (a) δ^+ N–F δ^-
(b) δ^+ As–O δ^-

42. (a) force 1: polar covalent bond; force 2: dipole–dipole force

Unit 1 Self-Quiz, p. 136

1. (b)	15. (d)	29. F
2. (c)	16. (b)	30. T
3. (c)	17. (a)	31. F
4. (a)	18. (d)	32. T
5. (a)	19. (b)	33. T
6. (d)	20. (c)	34. T
7. (a)	21. (d)	35. F
8. (b)	22. F	36. T
9. (c)	23. T	37. F
10. (d)	24. T	38. F
11. (d)	25. F	39. T
12. (c)	26. F	40. T
13. (a)	27. T	41. F
14. (a)	28. F	

Unit 1 Review, p. 138

1. (b)	9. (a)	17. F
2. (c)	10. (b)	18. T
3. (b)	11. (c)	19. T
4. (b)	12. (b)	20. F
5. (a)	13. T	21. T
6. (d)	14. F	22. T
7. (a)	15. F	23. T
8. (d)	16. F	24. F

25. (a) iii
(b) ii
(c) i
(d) iv

28. (a) Br^- (d) Mg^{2+}
(b) N^{3-} (e) S^{2-}
(c) K^+

29. (a) sulfate (e) perchlorate
(b) nitrate (f) cyanide
(c) carbonate (g) sulfite
(d) bromate

30. (a) ClO_3^- (e) $Cr_2O_7^{2-}$
(b) NH_4^+ (f) MnO_4^-
(c) HCO_3^- (g) HSO_4^-
(d) ClO^-

39. (a) sodium bromide
(b) magnesium sulfide
(c) copper(I) hydroxide
(d) tin(II) chloride
(e) potassium sulfite

40. (a) sulfur trioxide
(b) arsenic trichloride
(c) nitrogen monoxide
(d) carbon tetrachloride
(e) diphosphorus pentoxide

41. (a) NaF (f) $Mg(NO_3)_2$
(b) $CaCl_2$ (g) $Ba_3(PO_4)_2$
(c) Hg_2O (h) $MgSO_3$
(d) KCN (i) $Ni(ClO_4)_2$
(e) $(NH_4)_2SO_4$ (j) Cu_2S

42. (a) SO_2 (d) PCl_3
(b) CI_4 (e) N_2O_4
(c) SiO_2

48. (a) dipole–dipole forces, London dispersion forces, and hydrogen bonds.
(b) dipole–dipole forces and London dispersion forces

56. (a) mercury(I) chloride
(b) mercury(II) oxide
(c) mercury(I) nitrate

58. 2.4×10^2 u

65. (a) polar covalent bond
(b) polar covalent bond
(c) ionic bond
(d) non-polar covalent bond
(e) polar covalent bond

Unit 2

Are You Ready? pp. 148–149

1. (a) colour change
(b) $CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$
(c) bubbles of gas
(d) oxygen
(e) $2\ H_2O_2(aq) \rightarrow 2\ H_2O(l) + O_2(g)$
(f) changes in energy, precipitate formation

2. (a) $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$
(e) 36 g

3. (a) metal (b) XCl_2
(c) group 2

4. (a) K_2S (f) CS_2
(b) NH_4ClO_3 (g) KH_2PO_4
(c) Na_2SO_3 (h) $Ba(ClO)_2$
(d) $Fe(NO_2)_3$ (i) $MnSO_4$
(e) $Ca_3(PO_4)_2$ (j) NH_4ClO_4

5. (a) aluminum chloride
(b) zinc chlorate
(c) lead(IV) oxide
(d) ammonium carbonate
(e) sodium phosphite
(f) calcium hydrogen carbonate
(g) zinc chlorite
(h) sulfur dioxide
(i) dinitrogen monoxide
(j) iron(III) hydroxide

6. (a) $4\ K + O_2 \rightarrow 2\ K_2O$
(b) $P_4 + 6\ Cl_2 \rightarrow 4\ PCl_3$
(c) $C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$
(d) $2\ Fe + 3\ H_2SO_4 \rightarrow Fe_2(SO_4)_3 + 3\ H_2$
(e) $2\ HBr + Mg(OH)_2 \rightarrow MgBr_2 + 2\ H_2O$

(f) $2\ Na_3PO_4 + 3\ CaF_2 \rightarrow 6\ NaF + Ca_3(PO_4)_2$
(g) $4\ (NH_4)_3PO_4 + 3\ Sn(NO_3)_4 \rightarrow Sn_3(PO_4)_4 + 12\ NH_4NO_3$

8. (a) pink sample on the right
(b) $NaHCO_3$
(c) raise the pH

9. (a) able to neutralize acids, be safe to handle
(b) $NaHCO_3(aq)$

4.1 Tutorial 1 Practice, p. 155

1. (a) $4\ P + 5\ O_2 \rightarrow 2\ P_2O_5$
(b) $K_2O + H_2O \rightarrow 2\ KOH$
(c) $2\ AlBr_3 + 3\ K_2SO_4 \rightarrow 6\ KBr + Al_2(SO_4)_3$
(d) $FeCl_3 + 3\ NaOH \rightarrow Fe(OH)_3 + 3\ NaCl$
(e) $2\ AgNO_3 + H_2S \rightarrow Ag_2S + 2\ HNO_3$
(f) $(NH_4)_2CO_3 \rightarrow 2\ NH_3 + H_2O + CO_2$

4.1 Questions, p. 155

3. (a) copper metal + nitric acid → nitrogen dioxide + water + copper(II) nitrate
(b) $2Cu(s) + 4\ HNO_3(aq) \rightarrow 2\ NO_2(g) + 2\ H_2O(l) + 2Cu(NO_3)_2(aq)$
(c) 3
(d) 4
(e) 12
(f) (l) is liquid; (aq) means aqueous or "in water"

5. (a) $S_8 + 8\ O_2 \rightarrow 8\ SO_2$
(b) $N_2 + 3\ H_2 \rightarrow 2\ NH_3$
(c) $2\ Na + 2\ H_2O \rightarrow 2\ NaOH + H_2$
(d) $3\ Li + AlCl_3 \rightarrow 3\ LiCl + Al$
(e) $2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$
(f) $2\ N_2 + 5\ O_2 \rightarrow 2\ N_2O_5$
(g) $6\ Li + B_2O_3 \rightarrow 3\ Li_2O + 2\ B$
(h) $Fe_2O_3 + 3\ H_2SO_4 \rightarrow Fe_2(SO_4)_3 + 3\ H_2O$
(i) $2\ H_3PO_4 + 3\ Ca(OH)_2 \rightarrow Ca_3(PO_4)_2 + 6\ H_2O$
(j) $4\ NH_3 + 3\ O_2 \rightarrow 2\ N_2 + 6\ H_2O$
(k) $Ca_3(PO_4)_2 + SiO_2 + C \rightarrow 3\ CaSiO_3 + CO + 2\ P$
(l) $2\ C_6H_6 + 15\ O_2 \rightarrow 12\ CO_2 + 6\ H_2O$

4.2 Tutorial 1 Practice, p. 157

(a) $Ca(s) + Br_2(l) \rightarrow CaBr_2(s)$
(b) $4\ Al(s) + 3\ O_2(g) \rightarrow 2\ Al_2O_3(s)$

4.2 Questions, p. 161

2. (a) $Zn + S \rightarrow ZnS$
(b) $CaCl_2 \rightarrow Ca + Cl_2$
(c) $NH_3 + HCl \rightarrow NH_4Cl$
(d) $2\ K_2O \rightarrow 4\ K + O_2$
(e) $2\ AlCl_3 \rightarrow 2\ Al + 3\ Cl_2$
(f) $Mg(OH)_2 \rightarrow H_2O + MgO$

3. (a) oxygen
(b) decomposition
(c) $2\ H_2O_2(aq) \rightarrow 2\ H_2O(l) + O_2(g)$

4. (a) $4\ Al(s) + 3\ O_2(g) \rightarrow 2\ Al_2O_3(s)$
(b) $CuCO_3(s) \rightarrow CuO(s) + CO_2(g)$
(c) $2\ NI_3(s) \rightarrow N_2(g) + 3\ I_2(g)$

6. (a) $S(s) + O_2(g) \rightarrow SO_2(g)$
(b) $2\ SO_2(g) + O_2(g) \rightarrow 2\ SO_3(g)$
(c) $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

7. (a) carbon dioxide
(b) $2\ NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(g) + CO_2(g)$

8. (a) decomposition
(b) oxygen
(c) mercury
(d) mercury(II) oxide, HgO
(e) $2\ HgO(s) \rightarrow 2\ Hg(l) + O_2(g)$

9. (a) $2\ NaN_3(s) \rightarrow 3\ N_2(g) + 2\ Na(g)$

4.4 Tutorial 1 Practice, p. 166

1. (a) $Mg(s) + 2\ AgNO_3(aq) \rightarrow 2\ Ag(s) + Mg(NO_3)_2(aq)$
(b) $Zn(s) + FeCl_2(aq) \rightarrow Fe(s) + ZnCl_2(aq)$
(c) $Ni(s) + Al(NO_3)_3 \rightarrow$ no reaction

4.4 Tutorial 1 Practice, p. 167

2. (a) $Ca(s) + 2\ HBr(aq) \rightarrow H_2(g) + CaBr_2(aq)$
(b) $Cu(s) + H_2O(l) \rightarrow$ no reaction
(c) $Ba(s) + 2\ H_2O(l) \rightarrow H_2(g) + Ba(OH)_2(aq)$
(d) $Hg(l) + H_2SO_4(aq) \rightarrow$ no reaction

4.4 Questions, p. 169

2. (a) $Al(s) + 3\ AgNO_3(aq) \rightarrow 3\ Ag(s) + Al(NO_3)_3(aq)$
(b) $Zn(s) + Pb(NO_3)_2(aq) \rightarrow Pb(s) + Zn(NO_3)_2(aq)$
(c) $Au(s) + H_2O(l) \rightarrow$ no reaction
(d) $Mg(s) + H_2SO_4(aq) \rightarrow H_2(g) + MgSO_4(aq)$
(e) $Ca(s) + 2\ H_2O(l) \rightarrow H_2(g) + Ca(OH)_2(aq)$
(f) $2\ Al(s) + 6\ HCl(aq) \rightarrow 3\ H_2(g) + 2\ AlCl_3(aq)$

3. (a) $Br_2(l) + 2\ NaI(aq) \rightarrow I_2(aq) + 2\ NaBr(aq)$
(b) $Cl_2(g) + KF(aq) \rightarrow$ no reaction
(c) $F_2(g) + CaBr_2(aq) \rightarrow Br_2(aq) + CaF_2(aq)$

6. (a) $2\ Mg(s) + CO_2(s) \rightarrow 2\ MgO(s) + C(s)$

4.6 Tutorial 1 Practice, p. 175

1. (a) $Na_2S(aq) + Pb(NO_3)_2(aq) \rightarrow PbS(s) + 2\ NaNO_3(aq)$
(b) $NH_4Cl(aq) + K_2SO_4(aq) \rightarrow$ no reaction
(c) $2\ FeCl_3(aq) + 3\ Na_2CO_3(aq) \rightarrow Fe_2(CO_3)_3(s) + 6\ NaCl(aq)$

4.6 Questions, p. 177

2. (a) double displacement
(b) single displacement
(c) double displacement
(d) single displacement

3. (a) $PbSO_4$, slightly soluble
(b) $(NH_4)_3PO_4$, very soluble
(c) $CaSO_4$, slightly soluble
(d) $Al_2(SO_4)_3$, very soluble
(e) $Ca_3(PO_4)_2$, slightly soluble
(f) $BaSO_4$, slightly soluble
(g) $(NH_4)_2CO_3$, very soluble
(h) $CaCO_3$, slightly soluble

4. (a) $ZnCl_2(aq) + 2\ KOH(aq) \rightarrow Zn(OH)_2(s) + 2\ KCl(aq)$
(b) $Ni(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow NiCO_3(s) + 2\ NaNO_3(aq)$
(c) $Ba(OH)_2(aq) + K_2SO_4(aq) \rightarrow BaSO_4(s) + 2\ KOH(aq)$
(d) $3\ FeSO_4(aq) + 2\ K_3PO_4(aq) \rightarrow Fe_3(PO_4)_2(s) + 3\ K_2SO_4(aq)$
(e) $ZnS(s) + 2\ HCl(aq) \rightarrow H_2S(g) + ZnCl_2(aq)$
(f) $CaCO_3(s) + 2\ HNO_3(aq) \rightarrow CO_2(g) + H_2O(l) + Ca(NO_3)_2(aq)$
(g) $MgSO_3(aq) + 2\ HCl(aq) \rightarrow SO_2(g) + H_2O(l) + MgCl_2(aq)$

5. (a) $2\ AgNO_3(aq) + K_2SO_4(aq) \rightarrow Ag_2SO_4(s) + 2\ KNO_3(aq)$
(b) $NH_4Cl(aq) + Na_2S(aq) \rightarrow$ no reaction
(c) $3\ Pb(NO_3)_2(aq) + 2\ Na_3PO_4(aq) \rightarrow Pb_3(PO_4)_2(s) + 6\ NaNO_3(aq)$
(d) $BaCl_2(aq) + Ca(OH)_2(aq) \rightarrow$ no reaction
(e) $CuSO_4(aq) + K_2CO_3(aq) \rightarrow CuCO_3(s) + K_2SO_4(aq)$

6. (a) $BaCO_3(s) + H_2SO_4(aq) \rightarrow CO_2(g) + H_2O(l) + BaSO_4(s)$
(b) $4\ HF(aq) + SiO_2(s) \rightarrow 2\ H_2O(l) + SiF_4(aq)$
(c) $Ni(NO_3)_2(aq) + K_2CO_3(aq) \rightarrow NiCO_3(s) + 2\ KNO_3(aq)$
(d) $H_2SO_4(aq) + Na_2CO_3(aq) \rightarrow CO_2(g) + H_2O(l) + Na_2SO_4(aq)$
(e) $2\ HCl(aq) + ZnS(s) \rightarrow H_2S(g) + ZnCl_2(aq)$

Chapter 4 Self-Quiz, p. 183

1. (b)	7. (d)	13. F
2. (d)	8. (a)	14. T
3. (a)	9. (b)	15. F
4. (c)	10. (d)	16. F
5. (c)	11. F	17. F
6. (a)	12. T	

Chapter 4 Review, p. 184

1. (a)	8. (d)	15. T
2. (b)	9. (b)	16. F
3. (d)	10. (d)	17. F
4. (a)	11. F	18. F
5. (d)	12. F	19. T
6. (c)	13. F	
7. (b)	14. F	

20. (a) iv (c) ii
(b) i (d) iii

39. (a) $2\ NaClO_3(s) \rightarrow 2\ NaCl(s) + 3\ O_2(g)$
(b) $2\ C_2H_2(g) + 5\ O_2(g) \rightarrow 2\ H_2O(g) + 4\ CO_2(g)$
(c) $2\ Na(s) + 2\ H_2O(l) \rightarrow 2\ NaOH(aq) + H_2(g)$

40. (a) $Ca(s) + 2\ HNO_3(aq) \rightarrow H_2(g) + Ca(NO_3)_2(aq)$
(b) $2\ Fe(NO_3)_3(aq) + 3\ (NH_4)_2CO_3(aq) \rightarrow Fe_2(CO_3)_3(s) + 6\ NH_4NO_3(aq)$
(c) $2\ AgNO_3(aq) + K_2CrO_4(aq) \rightarrow Ag_2CrO_4(s) + 2\ KNO_3(aq)$

41. $C_6H_{12}O_6(aq) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(l) +$ energy

43. (a) decomposition (c) synthesis
(b) synthesis (d) decomposition

44. (a) $2\ HgO \rightarrow 2\ Hg + O_2$
(b) $4\ Na + O_2 \rightarrow 2\ Na_2O$
(c) $2\ NH_3 \rightarrow N_2 + 3\ H_2$

45. (a) carbon dioxide
(b) $Ag_2CO_3 \rightarrow Ag_2O + CO_2$

46. $2\ H_2O \rightarrow O_2 + 2\ H_2$

50. (a) $2\ Al(s) + 3\ H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3\ H_2(g)$
(b) $Zn(s) + Cu(NO_3)_2(aq) \rightarrow Cu(s) + Zn(NO_3)_2(aq)$
(c) no reaction

51. (a) single displacement
(b) $Cl_2(g) + 2\ KI(aq) \rightarrow I_2(aq) + 2\ KCl(aq)$

52. (a) single displacement
(b) hydrogen gas
(c) $2\ Fe(s) + 3\ H_2SO_4(aq) \rightarrow Fe_2(SO_4)_3(aq) + 3\ H_2(g)$

54. (a) very soluble
(b) very soluble
(c) slightly soluble

55. (a) $CdSO_4(aq) + H_2S(aq) \rightarrow CdS(s) + H_2SO_4(aq)$
(b) cadmium sulfide

56. (a) $Ni(OH)_2$

57. (a) $CaCO_3(s) + 2\,HCl(aq) \rightarrow CO_2(g) + H_2O(l) + CaCl_2(aq)$
    (b) release of $CO_2(g)$
58. (a) $NH_4NO_3(s) \rightarrow N_2O(g) + 2\,H_2O(g)$
    (b) decomposition
    (c) $2\,NH_3(g) + 2\,O_2(g) \rightarrow N_2O(g) + 3\,H_2O(l)$
    (d) double displacement
59. (a) $H_2(g) + O_2(g) \rightarrow H_2O_2(l)$
    (b) synthesis
60. (a) synthesis reaction
    (b) neutralization reaction
61. (a) double displacement reaction
    (b) single displacement reaction
62. (a) $AlCl_3(aq) + 3\,AgNO_3(aq) \rightarrow 3\,AgCl(s) + Al(NO_3)_3(aq)$; double displacement reaction
    (b) $2\,Li(s) + H_2SO_4(aq) \rightarrow Li_2SO_4(aq) + H_2(g)$; single displacement reaction
    (c) $H_2SO_4(aq) + 2\,NaOH(aq) \rightarrow 2\,H_2O(l) + Na_2SO_4(aq)$; neutralization reaction
63. (c) $Pb(NO_3)_2(aq) + 2\,NaI(aq) \rightarrow PbI_2(s) + 2\,NaNO_3(aq)$
64. (a) $Zn(s) + 2\,HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
66. (a) reactants: $Pb(NO_3)_2$, KI
    products: PbI_2, KNO_3
    (b) $Pb(NO_3)_2(aq) + 2\,KI(aq) \rightarrow PbI_2(s) + 2\,KNO_3(aq)$
67. (a) HgO
    (b) Hg and O_2
    (c) $2\,HgO(s) \rightarrow 2\,Hg(l) + O_2(g)$
68. (a) H_2 and Br_2
    (b) HBr
    (c) $Br_2(l) + H_2(g) \rightarrow 2\,HBr(g)$
69. (a) no products
    (b) $KCl(aq) + Na(s) \rightarrow$ no reaction
72. (a) $LiCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + LiNO_3(aq)$
73. (a) $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2\,NaCl(aq)$
    (b) $Ba(NO_3)_2(aq) + K_2CO_3(aq) \rightarrow BaCO_3(s) + 2\,KNO_3(aq)$
    (c) $MgCl_2(aq) + 2\,NaOH(aq) \rightarrow Mg(OH)_2(s) + 2\,NaCl(aq)$
    (d) no reaction
74. (a) $Mg(OH)_2(s) + 2\,HCl(aq) \rightarrow 2\,H_2O(l) + MgCl_2(aq)$
    (b) neutralization
75. (a) $SiO_2(s) + 3\,C(s) \rightarrow SiC(s) + 2\,CO(g)$
    (b) double displacement
76. $HNO_3(aq) + KOH(aq) \rightarrow H_2O(l) + KNO_3(aq)$
77. (a) sodium hydrogen carbonate
    (b) $NaHCO_3(s) + HF(aq) \rightarrow NaF(s) + H_2O(g) + CO_2(g)$

5.1 Questions, p. 197

2. (a) $2\,C_4H_{10}(g) + 13\,O_2(g) \rightarrow 8\,CO_2(g) + 10\,H_2O(g)$
3. (a) C_5H_{12} (c) C_6H_{14}
   (b) C_9H_{20}
4. (a) $2\,C_6H_{14}(g) + 19\,O_2(g) \rightarrow 12\,CO_2(g) + 14\,H_2O(g)$
   (b) $C_6H_{14}(g) + 8\,O_2(g) \rightarrow 3\,CO_2(g) + 3\,CO(g) + 7\,H_2O(g)$
11. (a) myth (d) myth
    (b) myth (e) true
    (c) true (f) true

5.3 Questions, p. 204

3. (a) basic (c) acidic
   (b) basic (d) acidic
4. (a) $4\,K(s) + O_2(g) \rightarrow 2\,K_2O(s)$; $2\,K_2O(s) + H_2O(l) \rightarrow 2\,KOH(aq)$
   (b) $2\,Cl_2(g) + O_2(g) \rightarrow 2\,Cl_2O(g)$; $Cl_2O(g) + H_2O(l) \rightarrow 2\,HClO(aq)$
   (c) $P_4(s) + 5\,O_2(g) \rightarrow 2\,P_2O_5(s)$; $P_2O_5(s) + 3\,H_2O(l) \rightarrow 2\,H_3PO_4(aq)$
5. $K_2O(s) + H_2O(l) \rightarrow 2\,KOH(aq)$
9. (b) $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(aq)$

5.4 Questions, p. 211

1. (a) $2\,HNO_3(aq) + Ca(OH)_2(aq) \rightarrow 2\,H_2O(l) + Ca(NO_3)_2(aq)$
   (b) $2\,HNO_3(aq) + K_2CO_3(s)(aq) \rightarrow H_2O(l) + CO_2(g) + 2\,KNO_3(aq)$
   (c) $3\,HC_2H_3O_2(aq) + Al(OH)_3(s) \rightarrow 3\,H_2O(l) + Al(C_2H_3O_2)_3(aq)$
   (d) $H_3PO_4(aq) + 3\,NaHCO_3(aq) \rightarrow 3\,H_2O(l) + 3\,CO_2(g) + Na_3PO_4(aq)$
   (e) $2\,H_3PO_4(aq) + 3\,Ca(OH)_2(aq) \rightarrow 6\,H_2O(l) + Ca_3(PO_4)_2(s)$
   (f) $H_2SO_4(aq) + Ca(HCO_3)_2(aq) \rightarrow 2\,H_2O(l) + 2\,CO_2(g) + CaSO_4(s)$
   (g) $2\,HClO_3(aq) + CaCO_3(s) \rightarrow H_2O(l) + CO_2(g) + Ca(ClO_3)_2(aq)$
2. $Ba(OH)_2(aq) + H_2SO_4(aq) \rightarrow 2\,H_2O(l) + BaSO_4(s)$;
   $Ba(HCO_3)_2(aq) + H_2SO_4(aq) \rightarrow 2\,H_2O(l) + 2\,CO_2(g) + BaSO_4(s)$
3. (a) carbonic acid
   (b) $2\,H_2CO_3(aq) + 2\,NaOH(aq) \rightarrow 3\,H_2O(l) + CO_2(g) + Na_2CO_3(aq)$

Chapter 5 Self-Quiz, p. 235

1. (d)	7. (a)	13. F
2. (b)	8. (b)	14. T
3. (c)	9. F	15. F
4. (c)	10. F	16. F
5. (b)	11. T	17. T
6. (d)	12. F	18. T

Chapter 5 Review, p. 236

1. (c)	8. (b)	15. F
2. (b)	9. (c)	16. T
3. (a)	10. F	17. F
4. (b)	11. T	18. F
5. (c)	12. F	19. F
6. (a)	13. F	
7. (d)	14. F	

20. (a) iii (c) iv
    (b) i (d) ii
21. (a) i (d) ii
    (b) vi (e) iii
    (c) v (f) iv
22. (a) iii (d) vi
    (b) v (e) iv
    (c) ii (f) i
23. carbon monoxide
26. Super Glue
27. metal
35. $C_5H_{12}(l) + 8\,O_2(g) \rightarrow 5\,CO_2(g) + 6\,H_2O(g)$
39. (a) $2\,Sr(s) + O_2(g) \rightarrow 2\,SrO(s)$; $SrO(s) + H_2O(l) \rightarrow Sr(OH)_2(aq)$
    (b) $P_4(s) + 3\,O_2(g) \rightarrow P_4O_6(s)$; $P_4O_6(s) + 6\,H_2O(l) \rightarrow 4\,H_3PO_3(aq)$
40. HCl(aq)
41. (a) $3\,HNO_3(aq) + Al(OH)_3(aq) \rightarrow 3\,H_2O(l) + Al(NO_3)_3(aq)$
    (b) aluminum nitrate
59. (b) $Li_2O(s) + H_2O(l) \rightarrow 2\,LiOH(aq)$
60. (a) $S_8(s) + 8\,O_2(g) \rightarrow 8\,SO_2(g)$
61. (a) acid + hydrogen carbonate $\rightarrow$ water + carbon dioxide + an ionic compound
68. (a) chemical

Unit 2 Self-Quiz, p. 244

1. (a)	13. (b)	25. T
2. (d)	14. (d)	26. F
3. (c)	15. (a)	27. T
4. (d)	16. (b)	28. T
5. (a)	17. (a)	29. F
6. (c)	18. (c)	30. T
7. (b)	19. (b)	31. T
8. (b)	20. F	32. T
9. (a)	21. F	33. F
10. (c)	22. T	34. T
11. (a)	23. F	35. F
12. (c)	24. F	

Unit 2 Review, p. 246

1. (b)	10. (d)	19. F
2. (a)	11. (c)	20. F
3. (c)	12. (d)	21. T
4. (c)	13. (b)	22. T
5. (d)	14. (a)	23. T
6. (b)	15. F	24. F
7. (a)	16. F	25. F
8. (d)	17. T	26. T
9. (b)	18. T	

27. (a) iii (c) i
    (b) iv (d) ii
28. (a) i (c) ii
    (b) iv (d) iii
40. hydrogen sulfide, H_2S
46. (a) sodium hydroxide, NaOH
50. (a) hydrogen peroxide, H_2O_2
53. (a) silicon, Si, and carbon dioxide, CO_2
    (b) silicon carbide, SiC, and silicon dioxide, SiO_2
    (c) $2\,Si(s) + CO_2(g) \rightarrow SiC(s) + SiO_2(s)$
54. (a) $2\,Al_2O_3 + 3\,C \rightarrow 4\,Al + 3\,CO_2$
    (b) $2\,Fe(OH)_2 + H_2O_2 \rightarrow 2\,Fe(OH)_3$
    (c) $2\,Ag_2O \rightarrow 4\,Ag + O_2$
56. (a) $H_2O + K_2O \rightarrow 2\,KOH$
    (b) $2\,Ca + O_2 \rightarrow 2\,CaO$
    (c) $CaCO_3 \rightarrow CaO + CO_2$
58. (a) synthesis
    (b) $Hg(l) + Cl_2(g) \rightarrow HgCl_2(s)$
59. (a) decomposition
    (b) $2\,AgNO_3(s) \rightarrow 2\,AgNO_2(s) + O_2(g)$
61. (a) single displacement
    (b) potassium + water $\rightarrow$ hydrogen + potassium hydroxide
    (c) $2\,K(s) + 2\,H_2O(l) \rightarrow 2\,KOH(aq) + H_2(g)$
63. (a) no reaction
    (b) $Ca(s) + 2\,H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$
    (c) $2\,Li(s) + 2\,H_2O(l) \rightarrow 2\,LiOH(aq) + H_2(g)$
    (d) no reaction
64. (a) double displacement
    (b) silver chloride and nitric acid
    (c) $AgNO_3 + HCl \rightarrow AgCl + HNO_3$

65. (a) lead(II) carbonate
 (b) $Pb(NO_3)_2(s) + Na_2CO_3(aq) \rightarrow PbCO_3(s) + 2\ NaNO_3(aq)$
71. (a) basic oxide
 (b) acidic oxide
 (c) acidic oxide
 (d) basic oxide
 (e) acidic oxide
 (f) basic oxide
80. $2\ H_2 + O_2 \rightarrow 2\ H_2O$
81. (a) $2\ LiOH + CO_2 \rightarrow Li_2CO_3 + H_2O$
82. (a) $SO_3(g) + H_2O(g) \rightarrow H_2SO_4(aq)$
 (b) synthesis reaction
83. $4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$
 $2\ NO(g) + O_2(g) \rightarrow 2\ NO_2(g)$
 $3\ NO_2(g) + H_2O(l) \rightarrow 2\ HNO_3(aq) + NO(g)$
84. (a) $2\ NaN_3(s) \rightarrow 2\ Na(g) + 3\ N_2(g)$
86. (a) single displacement reaction
 (b) hydrogen gas, H_2
 (c) $2\ K(s) + 2\ H_2O(l) \rightarrow 2\ KOH(aq) + H_2(g)$
87. (a) double displacement and decomposition
90. (a) double displacement
 (b) $H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$
91. (a) double displacement
 (b) $MnSO_4(aq) + 2\ NH_4OH(aq) \rightarrow (NH_4)_2SO_4(aq) + Mn(OH)_2(s)$
103. (a) $4\ P + 3\ O_2 \rightarrow 2\ P_2O_3$

Unit 3

Are You Ready? pp. 256–257

1. (a) (i) 2 carbon atoms, 2 hydrogen atoms
 (ii) 2 carbon atoms, 4 hydrogen atoms, 2 oxygen atoms
 (b) (i) ethyne: C_2H_2
 (ii) ethanoic acid: $C_2H_4O_2$ or $HC_2H_3O_2$
2. (a) both are $C_3H_6O_3$
 (b) no
4. (a) ionic
 (b) molecular
 (c) molecular
 (d) ionic
 (e) ionic
 (f) molecular
5. (a) $2\ KClO_3 \rightarrow 2\ KCl + 3\ O_2$
 (b) $P_4 + 5\ O_2 \rightarrow 2\ P_2O_5$
 (c) $Al_2(SO_4)_3 + 3\ Ca(OH)_2 \rightarrow 2\ Al(OH)_3 + 3\ CaSO_4$
 (d) $2\ C_2H_2 + 5\ O_2 \rightarrow 4\ CO_2 + 2\ H_2O$
 (e) $3\ Fe + Al_2(SO_4)_3 \rightarrow 3\ FeSO_4 + 2\ Al$
 (f) $2\ H_2O_2 \rightarrow 2\ H_2O + O_2$
6. (a) NR
 (b) $Ag(s) + Mg(NO_3)_2(aq)$
 (c) NR
 (d) NR
 (e) $H_2(g) + PbCl_2(aq)$
7. (a) $2\ H_2O_2 \rightarrow 2\ H_2O + O_2$
 (b) decomposition
 (c) 3.6 g
 (d) 7.2 g of water, 6.4 g of oxygen
8. (a) barium chlorate
 (b) tin(II) sulfite
 (c) iron(III) chloride
 (d) manganese(II) phosphate
 (e) lead(II) nitrate
 (f) tetraphosphorus decoxide
9. (a) $Cu(NO_3)_2$
 (b) CCl4
 (c) CO
 (d) Fe(ClO3)3
 (e) $Ca(HCO_3)_2$
10. (a) $CuO + H_2 \rightarrow Cu + H_2O$
 (b) $Pb(NO_3)_2 + 2\ KI \rightarrow PbI_2 + 2\ KNO_3$
 (c) $BaCl_2 + Na_2SO_4 \rightarrow BaSO_4 + 2\ NaCl$
 (d) $2\ Na_3PO_4 + 3\ CaCl_2 \rightarrow Ca_3(PO_4)_2 + 6\ NaCl$
11. (a) 52.6 %
 (b) 47.4 %
12. $B = AC,\ C = \dfrac{B}{A}$
13. (a) exact
 (b) measurement
 (c) measurement
 (d) exact
 (e) exact
 (f) measurement
14. (a) 12.37 g
 (c) 12.37 g
 (d) −0.05 g
15. (a) 90.4 g
 (b) 2.0 g/mL
 (c) 450 atoms
 (d) 4.7×10^{-3} L
 (e) 1.59×10^{-25} g/u
 (f) 1.3 mL

6.1 Practice Questions, p. 263

1. (a) quantitative
 (b) quantitative
 (c) qualitative
 (d) qualitative
 (e) qualitative
 (f) quantitative
2. (a) mass and age

6.3 Questions, p. 270

5. (a) 48
 (b) 2.4×10^{22} molecules CO_2
6. (a) 2.00×10^{-2}
 (b) 2×10^{22}
 (c) 6.7×10^{-4}
7. 7.9×10^{-8} mol

6.4 Questions, p. 277

4. (a) 159.70 g/mol
 (b) 100.09 g/mol
 (c) 114.26 g/mol
 (d) 206.98 g/mol
 (e) 96.11 g/mol
5. 0.0901 mol
6. 4.12×10^{-4} mol
7. 2.20×10^{-2} g
8. 2.1×10^{-4} mol
9. (a) 63 000 kg
 (b) 6.4×10^{5} mol
10. 9.00 g
11. 58 g/mol
12. 23.0 g

6.5 Questions, p. 283

1. (a) 1.2×10^{24}
 5.3×10^{19}
2. (a) 0.500 mol
 (b) 1.00×10^{2} mol
 (c) 0.420 mol
3. (a) 4.2×10^{24} atoms C, 8.4×10^{24} atoms O
 (b) 310 g
4. 9.704×10^{23} formula units
5. (a) 1.388×10^{-3} mol
 (b) 6.68×10^{21} atoms
6. (a) 4.92×10^{22} atoms
 (b) 1.41×10^{29} atoms
7. (a) 4.3×10^{19} formula units
 (b) 1.3×10^{20} oxide ions
8. (a) 6.3×10^{-3} mol
 (b) 3.8×10^{21} FPO_3^{2-} ions; 7.6×10^{21} Na^{+} ions
9. (a) 2.84×10^{24} atoms Ag; 1.8×10^{22} atoms Au

6.6 Questions, p. 288

1. 40.0 % Ca, 12 % C, 48.0 % O
2. 52.2 % C; 34.6 % O; 13 % H
4. (a) 0.80 g H
 (b) 92.0 % C, 8.0 % H
5. (a) C_3H_4
 (b) same percentage
 (c) C_2H_4
7. (a) 3.09 % H, 65.31 % O, 31.60 % P
 (b) 79.85 % Cu, 20.15 % S
 (c) 69.94 % Fe, 30.06 % O
 (d) 17.48 % B, 77.62 % O, 4.90 % H

6.7 Questions, p. 293

2. H_2CO_3, $K_2Cr_2O_7$, $C_3H_6O_3N$
3. (a) same, NO_2
 (b) not the same
 (c) same, CH
 (d) same, C_6H_5O
4. (a) 2, C_2H_7
 (b) 3, $C_2H_2O_3$
 (c) 2, $Na_2S_2O_3$
 (d) 3, $C_4H_{11}O_3$
5. $C_3H_4O_3$
6. compound 1: NO; compound 2: NO_2
7. $C_6H_{11}ON$
8. Al_4C_3
9. HgO
10. (a) 74.2 % X
 (b) Na, sodium

6.9 Questions, p. 300

2. (a) N_2O_4
 (b) C_6H_{12}
 (c) $C_6H_9O_9$
 (d) $C_2F_2Br_2O_2$
6. empirical formula: $C_4N_2H_5O$; molecular formula: $C_8N_4H_{10}O_2$
7. (a) $C_{20}H_{28}O_2$
8. empirical formula: CH_2O; molecular formula: $C_6H_{12}O_6$

Chapter 6 Self-Quiz, p. 307

1. (b)	8. (c)	15. T
2. (a)	9. (d)	16. F
3. (b)	10. T	17. F
4. (c)	11. T	18. F
5. (c)	12. F	19. T
6. (d)	13. F	20. T
7. (d)	14. T	

Chapter 6 Review, p. 308

1. (a)	4. (b)	7. (a)
2. (b)	5. (c)	8. (b)
3. (d)	6. (c)	9. (a)

10. (a) ii, iv (c) i
 (b) i, iii, iv

11. F	14. T	17. F
12. T	15. F	18. T
13. T	16. F	19. F

20. (a) qualitative analysis
 (b) quantitative analysis
21. (a) quantitative analysis
 (b) qualitative analysis
22. (b) 9.28×10^{-23} g
23. (b) 1.1×10^{19}
24. 1.50×10^{-2}
27. 26.98 g/mol
28. (a) 16.00 u
 (b) 32.00 g/mol
 (c) 48.00 g/mol

29. (b) 237.95 g/mol
31. (a) 158.04 g/mol
    (b) 1.63 mol
32. 150 g
33. 0.84 kg
34. 4.1×10^{23} Si atoms
35. (a) 8.4×10^{23} molecules
    (b) 6.7×10^{24} atoms
36. 0.253 mol
38. 75.0 %
39. 25.53 % Mg, 74.47 % Cl
41. 38.7 % K, 13.8 % N, 47.5 % O
42. (a) 22.4 g Br
    (b) 10.1 % Al, 89.9 % Br
43. 5.93 % H, 94.1 % S
46. KI
47. (a) 25.8 % O
    (b) Na_2O
48. $POCl_3$
49. N_2O_4
50. PCl_5
53. (a) quantitative (d) qualitative
    (b) quantitative (e) quantitative
    (c) qualitative
54. 4.98×10^{-11} mol
55. 1.5×10^{18} km
59. $MgSO_4 \cdot 7\ H_2O$ and $MgSO_4 \cdot 5\ H_2O$
60. (a) 275 g
    (b) 0.662 mol
61. (a) 20.2 mol
    (b) 669 g
62. 3.76×10^{24} atoms
63. (a) 7.22×10^{22} atoms
    (b) 6.54×10^{22} atoms
64. (a) 3.9×10^{22} SO_3 molecules
    (b) 1.09×10^{25} atoms
65. 46.5 % Fe
66. 52.1 % C, 13.2 % H, 34.7 % O
68. 69.9 % Fe, 30.1 % O
69. 40.0 % Ca, 12.0 % C, 48.0 % O
70. 56.3 % P, 43.7 % S
71. (a) CO_2 and CH_4
    (b) $C_6H_{12}O_6$ and H_2O_2
73. (a) 56.9 % S
    (b) MgS
74. $CaCl_2$
75. (a) 53.3 % O
    (b) CH_2O
76. (a) 71.1 % O
    (b) HCO_2
    (c) $H_2C_2O_4$
77. P_4O_{10}
79. (b) 90 mol N (c) 182 kg H

7.1 Questions, p. 320

3. 2:3
6. (a) 2 mol NH_3, 2.5 mol O_2
   (b) 1.2 mol NH_3, 1.5 mol O_2
   (c) 5.2×10^{-3} mol NH_3, 6.5×10^{-3} mol O_2
8. (b) 2 molecules N_2, 6 molecules H_2
9. (a) 0.75 mol Mg_3N_2
   (b) 4.5 mol H_2O
   (c) 0.042 mol Mg_3N_2; 0.13 mol $Mg(OH)_2$; 0.083 mol NH_3

7.2 Questions, p. 325

3. 2.0 mol and 1.5 mol; 0.60 mol and 0.45 mol; 108 g and 96 g
4. (a) 20 %
5. (a) 0.757 g CaO
   (b) 0.243 g H_2O
6. (a) single displacement
   (b) 5400 g I_2
7. (a) $4\ Fe(s) + 3\ O_2(g) \rightarrow 2\ Fe_2O_3(s)$
   (b) 20 mol Fe; 10 mol Fe_2O_3
   (c) 3.22 g Fe
   (d) 4.60 g
8. (a) 2.20 g CO_2
9. (a) $NH_4NO_3(s) \rightarrow N_2O(g) + 2\ H_2O(g)$
   (b) decomposition
   (c) 0.550 g N_2O
10. 1050 kg HNO_3

7.3 Questions, p. 330

2. (a) limiting: H_2; excess: N_2
   (b) 6 NH_3 molecules
   (c) 1 molecule N_2 excess
3. (a) $Mg(s) + 2\ HCl(aq) \rightarrow H_2(g) + MgCl_2(aq)$
   (b) single displacment
7. (a) limiting: H_2; excess: Au_2S_3

7.4 Questions, p. 335

3. (a) limiting: Mg; excess: N_2
   (b) limiting: Ca; excess: $AlCl_3$
   (c) limiting: FeS_2; excess: O_2
4. (a) 0.48 mol Ag
   (b) 0.040 mol $AgNO_3$ excess
5. (a) 0.35 mol $AlCl_3$
   (b) 0.15 mol HCl excess
6. 450 g SO_3
7. (a) $H_2(g) + Cl_2(g) \rightarrow 2\ HCl(g)$
   (b) 329 g HCl
8. 0.73 g $AlCl_3$
9. 25.4 g CO_2
10. (a) Cl_2
    (b) 40.1 g $TiCl_4$
11. 10 g HCN

7.5 Questions, p. 339

6. (a) $Zn(s) + 2\ AgNO_3(aq) \rightarrow 2\ Ag(s) + Zn(NO_3)_2(aq)$
   (b) 9.9 g Ag
   (c) 73 %
7. (a) 130 g Cl_2
   (b) 220 g NaCl
8. 87.2 %
9. (a) 13 g $(NH_4)_2SO_4$
   (b) 79 %
10. 59.2 %
11. (a) $N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$
    (b) 298 g NH_3
    (c) 26.9 mol N_2
    (d) 755 g N_2
12. 1.4 kg CO
13. 125 g S
14. (a) 51.6 g P_4O_{10}
    (b) 67.8 g H_3PO_4

Chapter 7 Self-Quiz, p. 345

1. (c)	8. T	15. F
2. (b)	9. F	16. T
3. (b)	10. T	17. F
4. (a)	11. F	18. T
5. (c)	12. T	19. F
6. (d)	13. F	20. T
7. (b)	14. F	

Chapter 7 Review, p. 346

1. (d)	11. F
2. (a)	12. F
3. (c)	13. T
4. (a)	14. F
5. (b)	15. T
6. (a)	16. F
7. (b)	17. T
8. (d)	18. F
9. (d)	19. F
10. (c)	

20. (a) ii (c) iii
    (b) iv (d) i
21. (a) 3:1
    (b) 1:1
    (c) 3:2
22. (a) 0.6 mol Cu
    (b) 0.2 mol CuO
23. nitrogen, $N_2(g)$
24. antacid tablet
25. yellow-orange
38. (a) 0.25; 1.0
    (b) 0.800; 0.400
39. (a) 85.3 mol O_2
    (b) 13 mol H_2O
41. (a) $2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$
    (b) 22 kg CO_2
42. (a) 1:2
    (b) 1:4
44. (a) $2\ Cu(s) + O_2(g) \rightarrow 2\ CuO(s)$
    (b) 7.0 mol of copper(II) oxide; 0.5 mol of excess
45. (a) 14.4 g C; 21.6 g H_2O
    (b) 54.0 g $C_6H_{12}O_6$
46. 1.00×10^{2} g Hg; 8.00 g O_2
50. (a) 0.903 g Ag
    (b) 9.03 %
51. (a) $CH_4(g) + 2\ O_2(g) \rightarrow CO_2(g) + 2\ H_2O(g)$
    (b) 1.5×10^{8} tonnes
55. (a) $Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$
57. (a) 13.0 g AsH_3
    (b) 3.32 g As_2O_3
58. (a) $Fe(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2(g)$
    (b) 1.45 g H_2
    (c) 29.8 g H_2SO_4
59. (a) 5600 kg NO
    (b) 1300 kg NH_3
60. 31.4 g PH_3
61. 3.94 kg $NaAu(CN)_2$
62. (a) 29.5 g $C_6H_5CH_3$
    (b) 74.6 %
63. (a) 17 g
    (b) 13 g $C_7H_6O_3$

Unit 3 Self-Quiz, p. 354

1. (a)	13. (c)	25. F
2. (d)	14. (d)	26. T
3. (a)	15. (d)	27. T
4. (b)	16. (d)	28. F
5. (b)	17. T	29. T
6. (b)	18. F	30. F
7. (d)	19. F	31. T
8. (d)	20. T	32. T
9. (b)	21. T	33. F
10. (a)	22. F	34. F
11. (d)	23. F	35. F
12. (c)	24. T	

Unit 3 Review, p. 356

1. (c)
2. (b)
3. (c)
4. (a)
5. (b)
6. (a)
7. (a)
8. (b)
9. (b)
10. (d)
11. (b)
12. (b)
13. (a)
14. (d)
15. (b)
16. (a)
17. F
18. F
19. F
20. F
21. F
22. F
23. F
24. F
25. T
26. F
27. T
28. T
29. F
30. F
31. F
34. 7 mol; 4×10^{24} molecules
35. 1.3×10^{-1} mol
36. 0.250 mol and 39.9 g Fe_2O_3
37. (a) 2.0×10^6 g
    (b) 4.3 %
    (c) $\$4.3 \times 10^9$
38. 2.33×10^{-23} g/atom
39. 74.10 g/mol
40. 22.6 g Ti
41. 0.0642 mol Zn
44. 2.23×10^{23} atoms Al
45. (a) 6.02×10^{23} molecules
    (b) 4.82×10^{24} atoms
46. 1.5×10^{21} atoms
47. 5.7×10^{19} molecules HNO_3
48. 37 g H; 63 g O_2
49. 77.5 % KNO_3; 22.5 % KCl
50. 69.94 % Fe; 30.06 % O
51. 17.27 % N
52. 6.73 % H
53. 50 g tin; 80 g antimony; 370 g lead
54. Al_2S_3
55. empirical formula CH_3; molecular formula C_2H_6
57. 581 g
58. 2 mol H_2O: 1 mol O_2
59. 10 mol NH_3
60. 3.50 mol NO
61. 23.8 g F_2
63. Na
64. Fe
65. ethylene oxide
66. 20 molecules CO_2
76. (a) CH
    (b) (i) C = 92.3 %, H = 7.76 %
        (ii) C = 92.3 %, H = 7.76 %
        (iii) C = 92.3 %, H = 7.76 %
    (c) (i) 13.0 g/mol
        (ii) 26 g/mol
        (iii) 78 g/mol
    (d) all correspond to diagram A
77. $C_8H_{16}O_4$
78. 8.05×10^{21} atoms

Unit 4

Are You Ready? pp. 366–367

3. relative density
5. (a) ionic
   (b) greater than the melting point of ice
7. (a) between molecules: London forces, dipole–dipole attractions and hydrogen bonding
   between ions: ionic bonding
   (b) London forces, dipole–dipole attractions, hydrogen bonding, ionic bonds
   (c) H_2, H_2S, H_2O
8. (b) same result
9. (a) $Pb(NO_3)_2(aq) + 2\ NaI(aq) \rightarrow PbI_2(s) + 2\ NaNO_3(aq)$
   (b) PbI_2
   (c) 29.97 g NaI
10. (a) potassium sulfide
    (b) aluminum hydroxide
    (c) nitrogen dioxide
    (d) diphosphorus pentoxide
    (e) lead(II) sulfite
11. (a) KNO_2
    (b) $Fe_2(SO_4)_3$
    (c) $Ca(HCO_3)_2$
    (d) $MnSO_4{\bullet}7H_2O$
    (e) SO_3
12. 74.10 g/mol
13. (a) 10
    (b) basic
    (c) $Ca(OH)_2$
14. (a) $Ba(OH)_2(aq) + H_2SO_4(aq) \rightarrow BaSO_4(s) + 2\ H_2O(l)$

8.1 Questions, p. 375

2. surface tension

8.2 Questions, p. 381

1. (a) solution
   (b) pure substance
   (c) heterogeneous mixture
   (d) heterogeneous mixture
   (e) pure substance
   (f) solution
   (g) solution
   (h) solution
   (i) solution
   (j) solution

8.3 Tutorial 1 Practice, p. 384

1. (a) $CaCl_2(s) \rightarrow Ca^{2+}(aq) + 2\ Cl^-(aq)$
   (b) $NH_4NO_2(s) \rightarrow NH_4^+(aq) + NO_2^-(aq)$
   (c) $Fe(OH)_3(s) \rightarrow Fe^{3+}(aq) + 3\ OH^-(aq)$
   (d) $Al_2(SO4)_3(s) \rightarrow 2\ Al^{3+}(aq) + 3\ SO_4^{2-}(aq)$

8.3 Questions, p. 389

1. (a) non-polar
6. (a) $Ca(NO_3)_2(s) \rightarrow Ca^{2+}(aq) + 2\ NO_3^-(aq)$
   (b) $KClO_4(s) \rightarrow K^+(aq) + ClO_4^-(aq)$
   (c) $(NH_4)_2CO_3(s) \rightarrow 2\ NH_4^+(aq) + CO_3^{2-}(aq)$
   (d) $Fe_2(SO_3)_3(s) \rightarrow 2\ Fe^{3+}(aq) + 3\ SO_3^{2-}(aq)$

8.5 Questions, p. 397

6. (a) 30 g
   (b) least soluble: $KClO_3$; most soluble: NH_4Cl
   (c) KNO_3
   (d) 50 g
   (e) 24 °C

8.6 Questions, p. 402

3. 0.400 mol/L
4. 170 g
5. 730 mL
6. 2.50×10^3 mL
7. 22 g
10. 11 g

8.7 Questions, p. 405

2. (a) 25 mL, 50 mL, and 75 mL
   (b) same amount of solute

8.8 Questions, p. 411

1. 1.0×10 g Cl_2
2. 430 g H_2O
3. 1.0 L 2-propanol
4. (a) 18% W/V
   (b) 1.0 mol/L
5. (a) 30.08 ppm
   (b) 33.2 L
6. 2.5×10^{-2} ppb
9. (a) 4.2 g $NaHCO_3$
   (b) 0.050 mol $NaHCO_3$
   (c) 8.4 % W/V $NaHCO_3$
   (d) 84 000 ppm
   (e) 1.0 mol/L
10. (a) 3.671×10^{20} kg
    (b) 4.4×10^9 kg Au

Self-Quiz, p. 415

1. (c)
2. (a)
3. (b)
4. (c)
5. (a)
6. (b)
7. (a)
8. (c)
9. (b)
10. (b)
11. T
12. F
13. F
14. F
15. T
16. T
17. F
18. F
19. T
20. F
21. T

Chapter 8 Review, pp. 416–421

1. (a)
2. (b)
3. (d)
4. (d)
5. (c)
6. (a)
7. (c)
8. (b)
9. (d)
10. F
11. F
12. F
13. T
14. F
15. F
16. T
17. F
18. T
19. (a) v
    (b) ii
    (c) i
    (d) iv
    (e) iii
20. (a) iii
    (b) vi
    (c) v
    (d) i
    (e) iv
    (f) ii
24. concentrated solution
33. (a) $c = \dfrac{n}{V}$
    (b) $V = \dfrac{n}{c}$
    (c) $n = cV$
35. (a) $c_c = \dfrac{c_dV_d}{V_c}$
    (b) $V_c = \dfrac{c_dV_d}{c_c}$
    (c) $c_d = \dfrac{c_cV_c}{V_d}$
    (d) $V_d = \dfrac{c_cV_c}{c_d}$
37. (a) 1.00×10^9 g
    (b) 1.00×10^6 L
43. (a) heterogeneous
    (b) homogeneous
    (c) homogeneous
    (d) homogeneous
44. (a) concentrated
    (b) dilute
    (c) dilute
47. (a) $KCl(s) \rightarrow K^+(aq) + Cl^-(aq)$
    (b) $MgBr_2(s) \rightarrow Mg^{2+}(aq) + 2\ Br^-(aq)$
    (c) $Li_2SO_4(s) \rightarrow 2\ Li^+(aq) + SO_4^{2-}(aq)$
54. 0.16 mol/L
55. 1.7 mol/L
57. 1.0 mol/L
58. 1340 g
59. (a) $c = \dfrac{n}{V}$
    (b) $V = \dfrac{n}{c}$
    (c) $c_d = \dfrac{c_cV_c}{V_d}$

63. 0.647 mol/L

66. (a) 2.9×10^5 ppb
 (b) 0.9×10^8 ppt

73. 3 L

74. (a) 20 %

9.1 Tutorial 1 Practice, p. 427

2. $Ca^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CaCO_3(s)$

9.1 Questions, p. 428

2. (a) $2\ AgNO_3(aq) + Sn(s) \rightarrow 2\ Ag(s) + Sn(NO_3)_2(aq)$
 (b) $2\ Ag^+(aq) + Sn(s) \rightarrow 2\ Ag(s) + Sn^{2+}(aq)$
 (c) potassium carbonate
 (d) $Sn^{2+}(aq) + CO_3^{2-}(aq) \rightarrow SnCO_3(s)$
3. (a) Product with lowest solubility: FeS
 Spectator ions: $Na^+(aq)$, $Br^-(aq)$
 Net ionic equation:
 $Fe^{2+}(aq) + S^{2-}(aq) \rightarrow FeS(s)$
 (b) Product with lowest solubility: $BaCO_3$
 Spectator ions: $NH_4^+(aq)$, $OH^-(aq)$
 Net ionic equation:
 $Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)$
5. $Cl_2(g) + 2\ I^-(aq) \rightarrow I_2(s) + 2\ Cl^-(aq)$
6. (a) NaOH, KOH
 (b) $Fe^{3+}(aq) + 3\ OH^-(aq) \rightarrow Fe(OH)_3(s)$
7. $2\ Cl^-(aq) + 2\ H_2O\ (l) \rightarrow Cl_2(g) + H_2(g) + 2\ OH^-(aq)$
8. (b) $3\ Ca^{2+}(aq) + 2\ PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$
9. (a) $2\ Al(s) + 3\ CuCl_2(aq) \rightarrow 3\ Cu(s) + 2\ AlCl_3(aq)$
10. (a) $Ca^{2+}(aq) + C_2O_4^{2-}(aq) \rightarrow CaC_2O_4(s)$

9.2 Questions, p. 436

5. 3×10^{-3} g
8. (a) $Ca^{2+}(aq) + CO_3^{2-} \rightarrow CaCO_3(s)$

9.3 Questions, p. 441

1. (a) qualitative (d) quantitative
 (b) quantitative (e) quantitative
 (c) qualitative
5. (a) yellow-green (c) green
 (b) whitish green (d) yellow

9.5 Questions, p. 449

1. (a) 0.13 mol/L
 (b) 7.5 g/L
2. (a) $NiSO_4(aq) + 2\ NaOH(aq) \rightarrow Ni(OH)_2(s) + Na_2SO_4(aq)$
 (b) 1.16 g
3. (a) 0.759 mol/L
 (b) 0.304 L
4. 0.701 L
5. (a) $2\ Al(s) + 3\ CuSO_4(aq) \rightarrow 3\ Cu(s) + Al_2(SO_4)_3(aq)$
 (b) 0.270 g
6. 1.9×10^{12} L
7. $CaCl_2$
8. (a) Na^+: 1.0 mol/L
 (b) NH_4^+: 0.4 mol/L
 (c) Fe^{3+}: 3.0 mol/L
9. 9.0 g

Chapter 9 Self-Quiz, p. 455

1. (b)	7. (a)	13. F
2. (d)	8. (b)	14. T
3. (c)	9. (c)	15. F
4. (c)	10. F	16. F
5. (b)	11. F	17. T
6. (d)	12. F	18. T

Chapter 9 Review, p. 456

1. (b)	7. (b)	13. T
2. (b)	8. (a)	14. F
3. (c)	9. (c)	15. T
4. (b)	10. F	16. F
5. (a)	11. T	17. F
6. (a)	12. F	18. T

19. (a) vi (d) i
 (b) iii (e) iv
 (c) ii (f) v
22. single-displacement and double-displacement
23. double-displacement
24. physical contaminants
30. (a) qualitative analysis
 (b) quantitative analysis
36. 3:1
38. (a) $2\ Na^+(aq) + 2\ I^-(aq) + Pb^{2+}(aq) + 2\ ClO_3^-(aq) \rightarrow 2\ Na^+(aq) + PbI_2(s) + 2\ ClO_3^-(aq)$
40. (a) very soluble (e) very soluble
 (b) very soluble (f) slightly soluble
 (c) very soluble (g) very soluble
 (d) slightly soluble (h) very soluble
42. cations NH_4^+, K^+, and Na^+, anion: NO_3^-
44. (a) heterogeneous mixture
 (b) heterogeneous mixture
 (c) homogeneous mixture
45. (a) collection
 (b) coagulation, flocculation, and sedimentation
 (c) filtration
 (d) softening
49. chlorination and filtration
50. (a) $2\ NaOH(aq) + H_2SO_4(aq) \rightarrow 2\ H_2O\ (l) + Na_2SO_4(aq)$
52. (a) $(NH_4)_2SO_4(s) \rightarrow 2\ NH^{4+}(aq) + SO_4^{2-}(aq)$
 (b) 2:1
53. (a) $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
54. (a) $Fe(s) + Cu(NO_3)_2(aq) \rightarrow Cu(s) + Fe(NO_3)_2(aq)$
 (b) $Fe(s) + Cu^{2+}(aq) + 2\ NO_3^-(aq) \rightarrow Cu(s) + Fe^{2+}(aq) + 2\ NO_3^-(aq)$
 (c) $Fe(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Fe^{2+}(aq)$

10.1 Tutorial 1 Practice, p. 468

1. (a) hydrobromic acid
 (b) phosphorus acid
 (c) hydrosulfurous acid
 (d) sulfurous acid
 (e) periodic acid
 (f) hypofluorous acid
2. (a) HF(aq)
 (b) $H_2SO_3(aq)$
 (c) HClO(aq)
 (d) $HBrO_4(aq)$

10.1 Questions, p. 469

2. (a) $HC_2H_3O_2(aq) + NaHCO_3(s) \rightarrow CO_2(g) + H_2O(l) + NaC_2H_3O_2(aq)$
3. (a) Ag: no; Zn: yes; Al: yes
 (b) $H_2SO_4(aq) + Zn(s) \rightarrow H_2(g) + ZnSO_4(aq)$
 $3\ H_2SO_4(aq) + 2\ Al(s) \rightarrow 3\ H_2(g) + Al_2(SO_4)_3(aq)$
4. (a) carbonic acid
 (b) hydroiodic acid
 (c) hydrosulfuric acid
 (d) phosphoric acid
 (e) nitric acid
 (f) hydrofluoric acid
 (g) nitrous acid
 (h) sulfurous acid
 (i) phosphorus acid
 (j) periodic acid
 (k) hypochlorous acid
6. (a) HBr(aq) (c) $HClO_2(aq)$
 (b) $HClO_4(aq)$ (d) HI(aq)
7. (a) magnesium hydroxide
 (b) potassium hydroxide

10.2 Questions, p. 475

8. (a) 100 times
10. (a) 10^3 times (c) 10^7 times
 (b) 10^4 times (d) 10^7 times
11. (a) 10^4 times (c) 10^2 times
 (b) 10^3 times (d) 10^5 times
12. (b) $H_2SO_4(aq) + Ba(OH)_2(aq) \rightarrow 2\ H_2O(l) + BaSO_4(s)$

10.3 Questions, p. 485

3. (a) 1.00 mol/L
8. (b) 50 mL
 (c) methyl red
9. (a) $CaCO_3(s) + 2\ HCl(aq) \rightarrow CO_2(g) + 2\ H_2O(l) + CaCl_2(aq)$
 (b) 0.63 g
10. (a) $2\ NH_4OH(aq) + H_2SO_4(aq) \rightarrow 2\ H_2O(l) + (NH_4)_2SO_4(aq)$
 (b) 8.54 mL
 (c) 0.43 mol/L
11. 0.0160 mol/L
12. (a) $KH(IO_3)_2(aq) + KOH(aq) \rightarrow H_2O(l) + 2\ KIO_3(aq)$
 (b) 0.151 mol/L
13. 0.598 mol/L

Chapter 10 Self-Quiz, p. 491

1. (b)	7. (c)	13. F
2. (c)	8. (a)	14. T
3. (d)	9. (b)	15. F
4. (a)	10. T	16. F
5. (b)	11. F	17. T
6. (c)	12. F	18. T

Chapter 10 Review, pp. 492–496

1. (d)	7. (b)	13. F
2. (c)	8. (a)	14. T
3. (b)	9. (a)	15. F
4. (b)	10. F	16. T
5. (d)	11. F	17. F
6. (c)	12. T	18. T

19. (a) i (d) iv
 (b) iii (e) ii
 (c) v
24. (a) carbon dioxide gas
 (b) hydrogen gas
39. (a) hydrobromic acid (c) nitric acid
 (b) phosphorous acid (d) perchloric acid
40. (a) barium hydroxide
 (b) lithium hydroxide
 (c) strontium hydroxide
 (d) beryllium hydroxide
41. (a) $H_3PO_4(aq)$ (c) $H_2S(aq)$
 (b) $HNO_2(aq)$ (d) $H_2Te(aq)$
42. (a) $Ca(OH)_2(aq)$ (c) NaOH(aq)
 (b) $NH_4OH(aq)$ (d) $Al(OH)_3(aq)$
43. (a) NR
 (b) $Mg(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
 (c) $2\ K(s) + 2\ HCl(aq) \rightarrow 2\ KCl(aq) + H_2(g)$
 (d) NR

(e) $Ca(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$
(f) $2\ Al(s) + 6\ HCl(aq) \rightarrow 2\ AlCl_3(aq) + 3\ H_2(g)$

45. carbonic acid is $H_2CO_3(aq)$; phosphoric acid is $H_3PO_4(aq)$
46. (b) $NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$; $Mg(OH)_2(s) + 2\ HCl(aq) \rightarrow 2\ MgCl_2(aq) + 2\ H_2O(l)$; $CaCO_3(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$
50. 0.093 mol/L
51. (b) 16.31 mL
(c) $H_2SO_4(aq) + 2\ KOH(aq) \rightarrow K_2SO_4(aq) + 2\ H_2O(l)$
(d) 0.054 mol/L
52. (a) 26.02 mL
(b) 1.3×10^{-3} mol
(c) 0.14 mol/L HCl
53. (a) base
(b) potassium hydroxide, KOH
(c) provide an electric current
55. (a) $2\ NO_2(g) + H_2O(l) \rightarrow HNO_2(aq) + HNO_3(aq)$
(b) $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$
(c) $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$
61. (a) basic

Unit 4 Self-Quiz, pp. 500–501

1. (b)	13. (a)	25. T
2. (d)	14. (b)	26. F
3. (a)	15. (d)	27. F
4. (b)	16. (c)	28. F
5. (c)	17. (b)	29. T
6. (a)	18. (a)	30. T
7. (d)	19. T	31. T
8. (b)	20. F	32. F
9. (b)	21. F	33. F
10. (c)	22. F	34. T
11. (d)	23. F	35. T
12. (c)	24. T	36. F

Unit 4 Review, pp. 502–509

1. (c)	9. (c)	17. T
2. (b)	10. (d)	18. F
3. (a)	11. (b)	19. T
4. (c)	12. (a)	20. T
5. (a)	13. F	21. F
6. (c)	14. T	22. F
7. (a)	15. F	23. F
8. (d)	16. F	24. T

25. (a) i (e) ii
(b) ii (f) iii
(c) i (g) i
(d) ii
26. (a) iv (d) vi
(b) i (e) ii
(c) v (f) iii
27. hydrogen bonding
28. ethanol
30. a gas
32. nitrate, NO_3^-
33. silver ions
34. total ionic equation
35. a net ionic equation
36. $SO_4^{2-}(aq)$
38. "potable" means "safe to drink"
40. reagent B
42. acids are colourless and bases are pink
43. (a) carbonic acid
(b) hydrofluoric acid
(c) phosphoric acid
(d) hydrosulfuric acid
(e) perchloric acid
44. (a) $H_3PO_3(aq)$ (d) $HClO_2(aq)$
(b) $HCl(aq)$ (e) $HClO_3(aq)$
(c) $HI(aq)$
47. $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
54. heterogeneous
60. (a) $2\ Al(s) + 6\ HCl(aq) \rightarrow 2\ AlCl_3(aq) + 3\ H_2(g)$
(b) Al
(c) $2\ Al(s) + 6\ H^+(aq) + 6\ Cl^-(aq) \rightarrow 2\ Al^{3+}(aq) + 6\ Cl^-(aq) + 3\ H_2(g)$
(d) $Cl^-(aq)$
(e) $2\ Al(s) + 6\ H^+(aq) \rightarrow 2\ Al^{3+}(aq) + 3\ H_2(g)$
65. (a) 2.00 mol/L
(b) 4.00 mol/L
(c) 6.00 mol/L
67. $CaO(s) + 2\ HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l)$
68. (a) iron(II) chloride, $FeCl_3(aq)$, and hydrogen sulfide, $H_2S(g)$
(b) $FeS(s) + 2\ HCl(aq) \rightarrow FeCl_2(aq) + H_2S(g)$
69. (a) basic solution
70. (a) HA(aq)
(b) $HA \rightarrow H^+(aq) + A^-(aq)$
(c) $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$
71. (a) ionic compound and water
(b) $H_2SO_4(aq) + 2\ KOH(aq) \rightarrow 2\ H_2O(l) + K_2SO_4(aq)$
(c) $2\ H^+(aq) + SO_4^{2-}(aq) + 2\ K^+(aq) + 2\ OH^-(aq) \rightarrow 2\ H_2O(l) + SO_4^{2-}(aq) + 2\ K^+(aq)$
(d) $2\ H^+(aq) + 2\ OH^-(aq) \rightarrow 2\ H_2O(l)$
75. (a) $Li_2S(aq) \rightarrow 2\ Li(aq)^+ + NO_3^-(aq)$
(b) $MgCl_2(aq) \rightarrow Mg^{2+}(aq) + 2\ Cl^-(aq)$
(c) $Al_2(SO_4)_3(aq) \rightarrow 2\ Al^{3+}(aq) + 3\ SO_4^{2-}(aq)$
78. 3.42 L
79. 0.294 mol/L
86. (a) $Na_2CO_3(aq) + CaCl_2(aq) \rightarrow CaCO_3(s) + 2\ NaCl(aq)$
(b) 0.04 mol
(c) 1:1
(d) 0.04 mol
(e) 0.16 L
88. (a) $2\ FeCl_3(aq) + 3\ H_2S(aq) \rightarrow 6\ HCl(aq) + Fe_2S_3(s)$
(b) 0.55 mol of $FeCl_3$, 1.10 mol of H_2S
(c) 3:2
(d) 2:1
(e) $FeCl_3$
(f) 57 g
90. (a) the pH 2 solution
(b) 1000 times
(c) pH 2
91. (a) $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$
(b) 1.0 mol/L
(c) 0.001 mol/L
93. (a) 0.01540 L
(b) $H_2SO_4(aq) + 2\ KOH(aq) \rightarrow K_2SO_4(aq) + 2\ H_2O(l)$
(c) 0.00308 mol
(d) 1.54×10^{-3} mol
(e) 1.54×10^{-1} mol/L
94. 0.178 mol/L
101. (b) $H_2SO_3(aq)$

Unit 5

Are You Ready? pp. 512–513

1. solid
2. gas
3. gas
6. cm^3, L, mL
7. (b) −70 °C
8. (a) direct relationship
(b) inverse relationship
9. (a) most: gas A; least: gas C
(b) (i) gas A, 87 mL; gas B, 66 mL; gas C, 36 mL
(ii) gas A, 58 mL; gas B, 49 mL; gas C, 30 mL
12. (a) Alberta and Ontario; 50 mt
(c) Alberta
14. balloons 1 and 2 rise; balloon 3 falls
15. (a) $2\ SO_2(g) + O_2(g) \rightarrow 2\ SO_3(g)$
(b) $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$
(c) 64.07 g/mol
(d) 1.6×10^4 mol $SO_2(g)$
(e) 1.6×10^4 mol $SO_3(g)$
(f) 1.6×10^4 mol $H_2SO_4(aq)$
(g) 1.5 t H_2SO_4
16. (a) $B = \dfrac{DEC}{FA}$ (c) $F = \dfrac{DEC}{AB}$
(b) $D = \dfrac{ABF}{CE}$

11.4 Questions, p. 533

5. (a) volatile organic compounds
7. (a) air quality health index
8. (a) O_3: max noon, min midnight; NO_2: max midnight, min noon

11.7 Questions, p. 546

1. (a) 440 mm Hg (c) 980 mm Hg
(b) 16.7 kPa (d) 130 kPa

11.8 Questions, p. 553

2. (a) 1427 °C (c) 5 K
(b) 2600 K (d) −219 °C
3. 0.15 L
4. 0.277 L
5. 0.016 L
6. 0.12 L

11.9 Questions, p. 562

2. 9.0×10 L
3. 3.0×10^2 kPa
4. (a) 42.2 kPa (c) 610 mL
(b) 240 mL
5. 4.4×10^2 kPa
6. 8.1×10^2 C
7. 89.2 kPa
8. 241 kPa
9. (a) decrease
(b) 2.2×10^2 kPa
10. 1.6×10^3 cm^3
11. 2.07×10^4 kPa
13. 5.8 mL
14. 15.4 L
15. 49 °C

Chapter 11 Self-Quiz, p. 567

1. (a)	8. (c)	15. T
2. (b)	9. (b)	16. F
3. (d)	10. F	17. T
4. (c)	11. F	18. F
5. (a)	12. F	
6. (b)	13. F	
7. (a)	14. T	

Chapter 11 Review, pp. 568–573

1. (d)	8. (c)	15. F
2. (b)	9. (d)	16. F
3. (d)	10. (a)	17. T
4. (d)	11. (c)	18. F
5. (a)	12. F	19. T
6. (a)	13. T	
7. (b)	14. F	

20. (a) ii (c) iv
    (b) i (d) iii
21. (a) solid (d) gas
    (b) liquid (e) gas
    (c) solid (f) liquid, gas
25. dust mites
27. iodine, bromine, chlorine
31. 18.7 mL Ar
41. 3.60×10^3 L
42. −8 °C
44. 8×10 mL
46. (a) $C_2H_4O_2(l) + NaHCO_3(s) \rightarrow NaC_2H_3O_2(aq) + CO_2(g) + H_2O(l)$
48. 3×10^9 t
58. (a) 6080 kPa
    (b) about 60 times greater
69. (a) 3260 kg (c) 2630 kg
    (b) 2140 m^3

12.1 Questions, p. 581

1. volume triples
2. 23 L
3. (a) 1.81 mol (d) 1.93×10^6 mL
   (b) 120 L (e) 1.47 mol
   (c) 61.4 g
4. (a) 49.9 L
   (b) 0.52 mol H_2, 1.1 g

12.2 Questions, p. 589

4. 1.10×10^2 L
5. 95.5 g
6. −270 °C
7. 203 kPa
8. (a) 69.5 g/mol
9. (a) 21.9 L/mol
10. 1.8 g/L

12.4 Questions, p. 597

3. 2.30×10^2 kPa
4. (a) 97.5 kPa
5. (a) 3.17 kPa
   (b) 96.8 kPa
   (c) 0.020 mol
6. 103.41 kPa

12.5 Questions, p. 603

1. (a) $2\,C_4H_{10}(g) + 13\,O_2(g) \rightarrow 8\,CO_2(g) + 10\,H_2O(g)$
   (b) 22.4 L
2. (a) 33.6 L
   (b) 37.16 L
3. 9 L
4. (a) 2.0×10^2 kg, 1.3×10^5 L
   (b) 9 trees
5. (a) 250 L
   (b) 6.31 kg
6. 150 L
7. 45.1 L

Chapter 12 Self-Quiz, p. 609

1. (d)	4. (a)	7. (d)
2. (c)	5. (b)	8. (d)
3. (c)	6. (c)	9. (a)
10. T	13. T	16. F
11. F	14. T	17. T
12. F	15. F	18. F

Chapter 12 Review, pp. 610–615

1. (b)	7. (a)	13. F
2. (c)	8. (b)	14. F
3. (d)	9. (c)	15. T
4. (b)	10. F	16. F
5. (d)	11. T	17. T
6. (b)	12. T	18. T

19. (a) iv (c) ii
    (b) i (e) iii
28. 101 kPa
37. (a) $3\,H_2(g) + N_2(g) \rightarrow 2\,NH_3(g)$
    (b) 1.5 L N_2
    (c) 3.0 L NH_3
38. 38 L
39. 140 L
40. 19.8 g
43. 1.1×10^{-3} mol
44. 1.3 g/L
45. 2.7×10^{-3} mol
49. 103 kPa
50. (a) 101.3 kPa
51. 154 kPa
52. (a) $2\,C_4H_{10}(g) + 13\,O_2(g) \rightarrow 8\,CO_2(g) + 10\,H_2O(g)$
    (b) 26 L
    (c) 2.0×10 L
    (d) 16 L
53. (a) $2\,C_2H_6(g) + 7\,O_2(g) \rightarrow 4\,CO_2(g) + 6\,H_2O(g)$
    (c) 125 L
    (d) 35 L
55. 8.5×10^3 mol
57. 21.2 L
58. 1.50 g
60. (a) 32.0 g/mol
    (b) oxygen
62. (a) 12.7 L
63. (a) $C_6H_{12}O_6(aq) \rightarrow 2\,CH_3CH_2OH(l) + 2\,CO_2(g) + \text{energy}$
    (b) 89.6 L
    (c) 78 g
64. (c) 117 mol
67. (a) 31.6 kPa contributed by argon; 126 kPa contributed by helium
    (c) 15.8 kPa is contributed by argon; 63.2 kPa is contributed by helium

Unit 5 Self-Quiz, pp. 618–619

1. (b)	13. (d)	25. F
2. (a)	14. (b)	26. F
3. (d)	15. (b)	27. F
4. (b)	16. (a)	28. T
5. (b)	17. T	29. F
6. (d)	18. F	30. F
7. (a)	19. F	31. T
8. (c)	20. F	32. T
9. (b)	21. F	33. F
10. (c)	22. T	34. F
11. (a)	23. T	35. F
12. (c)	24. F	36. T

Unit 5 Review, pp. 620–627

1. (c)	5. (b)	9. (b)
2. (d)	6. (d)	10. (c)
3. (a)	7. (d)	11. (a)
4. (a)	8. (b)	12. (c)
13. (b)	19. F	25. F
14. (d)	20. T	26. T
15. F	21. F	27. F
16. T	22. F	28. F
17. T	23. T	29. F
18. F	24. F	

30. (a) iii (d) i
    (b) v (e) iv
    (c) ii
31. (a) gas (f) gas
    (b) solid (g) gas and liquid
    (c) liquid and gas (h) solid
    (d) solid (i) liquid
    (e) solid
32. the troposphere
33. sulfur dioxide, $SO_2(g)$
34. (a) ozone, other gases, and particulate matter
    (b) sunlight
48. 40 cm^3
54. 470 cm^3
55. −37 °C
56. 2.39×10^3 m^3
57. 29.3 cm^3
59. (a) $2\,NH_3(aq) \rightarrow N_2(g) + 3\,H_2(g)$
60. 11.1 L
61. 2.91 L
62. 0.6 g/L
64. 30 kPa
65. (a) 150 kPa
    (b) 130 kPa for methane, 40 kPa for ethane, 30 kPa for propane
66. 82 kPa
68. 30 L
69. 40 L
77. (a) highest constant pressure: neon; lowest constant pressure: krypton
81. 56.6 g
82. (a) flasks A, C, G, I
    flasks D, F
    (c) 2:1
    (d) 2:1
83. 17 g/mol
84. (a) 108 g/mol
85. (a) 18.8 g/mol
    (b) $M_{methane} = 16.05$ g/mol
    $M_{ethane} = 30.08$ g/mol
    (c) X = 0.200
    (d) ethane: 20.0 %, methane: 80.0 %
86. yes
88. (a) 95.0 kPa
    (b) helium: 15 kPa, neon: 6.7 kPa, and argon: 10 kPa
    (c) 32 kPa
89. (a) nitrogen: 2.5×10^2 kPa, argon: 49 kPa, oxygen: 2.9×10^2 kPa
    (b) 5.9×10^2 kPa
90. (a) 0.0041 mol
    (b) 0.0231 mol
    (c) 0.0272 mol
    (d) 2000 atm
91. (a) $Zn(s) + 2\,HCl(aq) \rightarrow ZnCl_2 + H_2(g)$;
    $2\,Al(s) + 6\,HCl(aq) \rightarrow 2\,AlCl_3(aq) + 3\,H_2(g)$
    (b) reaction with zinc: 1.0×10^2 mL
    reaction with aluminum: 3.7×10^2 mL
    (c) zinc
97. helium

Glossary

A

absolute temperature a measurement of the average kinetic energy of the entities in a substance; unit symbol K

absolute zero the theoretical temperature at which the entities of a material contain no kinetic energy and therefore transmit no thermal energy; equal to −273.15 °C

acid a compound that produces hydrogen ions when mixed with water, forming a solution that conducts electricity, tastes sour, turns blue litmus red, and neutralizes bases

acidic oxide an oxide that forms an acidic solution when dissolved in water; a non-metallic oxide

activity series a list of elements arranged in order of their observed reactivity in single displacement reactions

actual yield the amount or mass of product actually collected during an experiment or industrial process

air pollution chemicals and particles in the atmosphere that harm living organisms or damage the environment

Air Quality Health Index (AQHI) a numerical scale used to indicate overall air quality based on concentrations of air pollutants including ozone, particulate matter, and nitrogen dioxide

alloy a solution of two or more metals

alpha particle a product of nuclear decay emitted by certain radioisotopes; a positively charged particle with the same structure as the nucleus of a helium atom

amalgam an alloy (solution) of mercury with other metals

amount (*n*) the quantity of a substance, measured in moles

amount concentration (*c*) the amount (in moles) of solute dissolved per litre of solution; unit symbol mol/L

anion a negatively charged ion formed by the addition of one or more electrons to a neutral atom

aqueous solution a solution in which water is the solvent

aquifer a layer of underground rock that holds a considerable quantity of water; an important source of fresh water

atmospheric pressure the force per unit area exerted by air on all objects

atom the smallest particle of an element

atomic mass the weighted average of the masses of all the naturally occurring isotopes of an element

atomic mass unit a very small unit of mass defined as $\frac{1}{12}$ the mass of a carbon-12 atom; unit symbol u

atomic number (*Z*) the unique number of protons in one atom of an element

atomic radius a measurement of the size of an atom, usually expressed in picometres (pm); the distance from the centre of an atom to the outermost electrons

Avogadro's constant (N_A) the number of entities in 1 mol of a substance

Avogadro's law the statement that the volume of a gas is directly related to the amount of gas, when the temperature and pressure of the gas remain constant; equal volumes of gases, under identical conditions, contain the same number of entities

B

base a compound that produces hydroxide ions when mixed with water, forming a solution that conducts electricity, tastes bitter, turns red litmus blue, and neutralizes acids

basic oxide an oxide that forms a basic solution when dissolved in water; a metallic oxide

beta particle a product of nuclear decay emitted by certain radioisotopes; a negatively charged particle identical to an electron

binary ionic compound a compound that consists of ions of only two elements

biodegradable capable of being broken down (decomposed) rapidly by the action of moisture, heat, and micro-organisms

bioremediation the process of using organisms to treat contaminated land or water

bonding capacity the number of covalent bonds that an atom can form

bonding electron an electron, in the valence shell of an atom, that is available to form a covalent bond with another atom

Boyle's law the statement that as the volume of a gas is decreased, the pressure of the gas increases proportionally, provided that the temperature and amount of gas remain constant; the volume and pressure of a gas are inversely proportional

Brownian motion the random movement of microscopic particles suspended in a liquid or a gas

burette a calibrated tube used to deliver variable known volumes of a liquid during a titration

C

carbon sequestration the process of removing carbon dioxide and other forms of carbon from the atmosphere, and then storing it

catalyst a substance that makes a chemical reaction occur faster without itself being consumed in the reaction

cation a positively charged ion formed by the removal of one or more electrons from the valence shell of a neutral atom

Charles' law the statement that as the temperature of a gas is increased, the volume of the gas increases proportionally, provided that the pressure and amount of gas remain constant; the volume and temperature of a gas are directly proportional

chemical bond the force of attraction between two atoms or ions

chemical reaction a process in which one or more substances change into one or more new substances

combined gas law the statement that the product of the pressure and volume of a gas sample is proportional to its absolute temperature in kelvins

combustion a chemical reaction in which a fuel burns in oxygen to produce combustion products and, often, a flame

complete combustion of a hydrocarbon the combustion (burning) of a hydrocarbon in a plentiful supply of oxygen to produce carbon dioxide, water, and energy

compostable the ability of a material to decompose naturally, resulting in a product that is able to sustain plant life

concentrated solution a solution with a relatively large quantity of solute dissolved per unit volume of solution

concentration the ratio of the quantity of solute to the quantity of solution or solvent; usually quantity of solute per unit volume of solution

covalent bond the bond that results from the sharing of a pair of electrons by two atoms

D

Dalton's law of partial pressures the statement that the total pressure of a mixture of non-reacting gases is equal to the sum of the partial pressures of the individual gases

decomposition reaction a reaction in which a large or more complex compound breaks down to form two (or more) simpler products; general pattern: $AB \rightarrow A + B$

diatomic made up of two atoms

dilute solution a solution with a relatively small quantity of solute dissolved per unit volume of solution

dilution the process of reducing the concentration of a solution; usually done by adding more solvent

dipole–dipole force an intermolecular force of attraction that forms between the slightly positive end of one polar molecule and the slightly negative end of an adjacent polar molecule

dissociation the separation of individual ions from an ionic compound as it dissolves in water

double displacement reaction a reaction in which elements in two compounds displace each other or trade places, producing two new compounds; general pattern: $AB + CD \rightarrow AD + CB$

E

effective nuclear charge the net force experienced by an electron in an atom due to the positively charged nucleus

electrolyte a compound that dissolves in water, producing a solution that conducts electricity

electron a negatively charged particle in an atom or ion

electron affinity the energy change that occurs when an electron is added to a neutral atom in the gaseous state

electronegativity the ability of an atom to attract bonding electrons to itself

electronegativity difference (ΔEN) the difference in electronegativities of two bonded atoms or ions

empirical formula a formula that shows the simplest whole-number ratio of elements in a compound

empirical knowledge knowledge that comes from investigation and observation

endpoint the point during a titration when a sudden change in an observable property of the solution occurs, usually a change in colour of an acid–base indicator or a significant change in pH

energy level a theoretical sphere around an atom where electrons exist; electron orbit

equivalence point the point in a titration when neutralization is complete

excess reagent a reactant that is still present after the reaction is complete

F

filtrate the clear liquid (solvent and any dissolved substances) collected after a mixture is filtered to remove any solid components

flame test a diagnostic test used to identify a specific element

flash smelting a fairly new technology for separating a metal from its ore by heating ore in an atmosphere of almost pure oxygen

formula equation a chemical equation in which all compounds are represented by their chemical formulas

formula unit the smallest repeating unit in an ionic crystal

full or stable octet an electron arrangement where the valence shell is filled with 8 valence electrons (2 for hydrogen and helium)

G

gamma ray a form of high-energy electromagnetic radiation emitted by certain radioisotopes

Gay-Lussac's law the statement that as the temperature of a gas increases, the pressure of the gas increases proportionally, provided that the volume and amount of gas remain constant; the temperature and pressure of a gas are directly proportional

greenhouse effect a natural process whereby gases and clouds absorb infrared radiation emitted from Earth's surface and radiate it, heating the atmosphere and Earth's surface

greenhouse gas any of several atmospheric gases (such as water vapour, carbon dioxide, and methane) that allow solar radiation to pass through the atmosphere but absorb infrared radiation emitted by Earth, thereby trapping thermal energy and making Earth warmer

group a column of elements in the periodic table; sometimes referred to as a family

H

hard water water that has a relatively high concentration of Ca^{2+} and Mg^{2+}

heterogeneous mixture a mixture that contains two or more phases

hydrate an ionic compound that contains water as part of its crystal structure

hydration the process in which ions are surrounded by water molecules

hydrogen bond an unusually strong dipole–dipole force between a hydrogen atom attached to a highly electronegative atom (N, O, or F) and a highly electronegative atom in another molecule

I

ideal gas a hypothetical gas composed of particles that have no size, travel in straight lines, and have no attraction to each other (no intermolecular forces)

ideal gas law the statement that the product of the pressure and volume of a gas is directly proportional to the amount and the absolute temperature of the gas; $pV = nRT$

immiscible unable to mix to form a solution; usually describing liquids that do not readily mix

incomplete combustion of a hydrocarbon the combustion (burning) of a hydrocarbon in a limited supply of oxygen; products may include carbon dioxide, carbon monoxide, soot, water, and energy

intermolecular force the attractive force between molecules

ion a charged entity formed when an atom gains or loses one or more electrons

ionic bond the electrostatic force of attraction between a positive ion and a negative ion; a type of chemical bond

ionic compound a pure substance composed of positively charged ions and negatively charged ions in a fixed ratio

ionic radius a measurement of the size of an ion, usually expressed in picometres (pm); the distance from the centre of an ion to the outermost electrons

ionization the formation of ions from uncharged molecules

ionization energy the quantity of energy required to remove an electron from an atom or ion in the gaseous state

isotope a form of an element in which the atoms have the same number of protons as all other forms of that element, but a different number of neutrons

isotopic abundance the percentage of a given isotope in a sample of an element

K

Kelvin temperature scale a temperature scale that includes absolute zero and the same-sized unit divisions as the Celsius temperature scale

kinetic energy energy possessed by moving objects

kinetic molecular theory the idea that all substances are composed of entities that are in constant, random motion

L

law of combining volumes the statement that, when measured at the same temperature and pressure, volumes of gaseous reactants and products of chemical reactions are always in simple ratios of whole numbers

law of conservation of mass the statement that, during a chemical reaction, the total mass of reactants equals the total mass of products

law of definite proportions the statement that a compound always contains the same proportion of elements by mass

leachate a liquid contaminated with dissolved or suspended substances picked up as the liquid passed through old mines or dump sites

Lewis structure a representation of covalent bonding based on Lewis symbols; a model in which shared electron pairs are shown as lines, and unshared electrons are shown as dots

Lewis symbol a representation of an element consisting of the chemical symbol and dots to represent the valence electrons; electron dot diagram

limiting reagent the reactant that is completely consumed in a chemical reaction; the reactant that determines how much product will be formed

London dispersion force a weak attractive force acting between all entities, including non-polar molecules and unbonded atoms, caused by the temporary imbalance of electrons within entities

lone pair a pair of electrons that is not involved in covalent bonding

M

mass number (*A*) the sum of the protons and neutrons in the nucleus of an atom

mass spectrometer a measuring instrument used to determine the mass and abundance of isotopes

matter anything that has mass and takes up space

metalloid an element that has properties of both metals and non-metals

metallurgy the science and technology of separating and refining metals from their ores and subsequent processing

mineral a naturally occurring solid that has a definite crystal structure and chemical composition

miscible able to mix to form a solution; usually describing liquids that mix with each other in all proportions to form a solution

molar mass the mass of 1 mol of a substance; unit symbol g/mol

molar volume the volume that one mole of gas occupies at a specified temperature and pressure

mole a unit of amount; the amount of substance containing 6.02×10^{23} entities; unit symbol mol

mole ratio the ratio of the amounts of the entities in a chemical reaction

molecular compound a pure substance composed of molecules made up of two or more non-metallic elements

molecular element a pure substance composed of molecules made up of two or more atoms of the same element

molecular formula a formula that shows the element symbols and exact number of each type of atom in a molecular compound

multivalent the property of having more than one possible valence

N

net ionic equation a chemical equation that includes only the entities that react during the reaction

neutralization reaction the reaction of an acid and a base in which the resulting solution has a pH closer to 7 than either of the reactants; a type of double displacement reaction that produces water and an ionic compound

neutron a neutral particle in an atom's nucleus

non-polar covalent bond a covalent bond formed between atoms with identical (or very similar) electronegativities

non-polar molecule a molecule in which the electrons are equally distributed among the atoms, resulting in no localized charges

non-renewable resource a natural resource that cannot be replaced as quickly as it is consumed

nuclear radiation energy or very small particles emitted from the nucleus of a radioisotope as it decays

O

octet rule a generalization stating that when atoms combine, they tend to achieve 8 valence electrons

off-gassing the release of one or more gases from a substance or product at normal temperatures and pressures

ore rock containing a relatively high proportion of a desirable mineral

organic compound a molecular compound that contains one or more carbon–carbon bonds and, often, one or more carbon–hydrogen bonds

oxide a compound composed of oxygen and one other element

oxyacid an acid that includes oxygen in its formula

oxyanion a negatively charged polyatomic ion that contains oxygen

P

partial pressure the pressure that a gas in a mixture would exert if it were the only gas present in the same volume and at the same temperature

particulate matter the mixture of very small solid and liquid particles found in the atmosphere

percentage composition the percentage, by mass, of each element in a compound

percentage yield the ratio, expressed as a percentage, of the actual yield to the theoretical yield

period a row in the periodic table

periodic law a rule, developed from many observations, stating that when the elements are arranged in order of increasing atomic number, their properties show a periodic recurrence and gradual change

petrochemical a compound manufactured from a fossil fuel

photochemical smog a hazy cloud of air pollutants formed by the reaction of emissions of factories and vehicles with sunlight

phytoremediation the process of using plants to treat contaminated land or water

polar covalent bond a covalent bond formed between atoms with significantly different electronegativities resulting in a bond with localized positive and negative charges or poles

polar molecule a molecule in which the uneven distribution of electrons results in a positive charge at one end and a negative charge at the other end

polyatomic ion an ion, made up of more than one atom, that acts as a single entity

polyatomic ionic compound a compound that consists of ions of more than two elements

potable water water that is suitable for drinking

precipitate a solid produced as a result of the reaction of two solutions

pressure (*P*) force per unit area; unit is the pascal, unit symbol Pa; 1 Pa = 1 N/m^2

primary standard a highly pure and stable chemical used to determine the precise concentration of acids or bases

proton a positively charged particle in the atom's nucleus

Q

qualitative analysis the process of identifying substances present in a sample; no measurements are involved

quantitative analysis the process of measuring the quantity of a substance in a sample, providing numerical data

R

radioactive having the potential to emit nuclear radiation upon decay

radioactive decay the spontaneous disintegration of unstable isotopes

radioisotope an isotope that spontaneously decays to produce two or more smaller nuclei and radiation

recycling converting a product back into material that can be used to make new goods

remediation the process of treating contaminated land or water so that it is safe for use again

renewable resource a natural substance that replenishes itself as it is used

S

saturated solution a solution that contains the maximum quantity of solute at a given temperature and pressure

single displacement reaction a reaction in which an element displaces another element in a compound, producing a new compound and a new element; general pattern: $A + BC \rightarrow AC + B$

smelting the chemical process that extracts a metal from its ore using heat and chemicals

soft water water that has a relatively low concentration of Ca^{2+} and Mg^{2+}

solubility the quantity of solute that dissolves in a given quantity of solvent at a given temperature

solubility curve a graph of the solubility of a substance over a range of temperatures

solute the substance that dissolves in a solvent

solution a homogenous mixture of two or more substances

solvent the substance that dissolves the solute

spectator ions ions that are not involved in a chemical reaction

standard ambient temperature and pressure (SATP) 25 °C and 100 kPa

standard pressure 101.325 kPa (often rounded to 101 kPa)

standard solution a solution for which the precise concentration is known

standard temperature and pressure (STP) 0 °C and 101.325 kPa

stock solution a concentrated solution that is used to prepare dilute solutions for actual use

stoichiometric amount an amount of reactants that is in the same proportion as the reactant coefficients in the balanced chemical equation

stoichiometry the study of the mass and amount relationships between reactants and products in a chemical reaction

strong acid a substance that ionizes completely in water; strong acids have strong acidic properties (e.g., low pH and vigorous reactivity)

structural formula a representation of the number, types, and arrangement of atoms in a molecule, with dashes representing covalent bonds

supersaturated solution a solution that contains more than the maximum quantity of solute that it should at a given temperature and pressure

surface tension a phenomenon, caused by forces of attraction between molecules, that leads to the formation of a skin-like film on the surface of a liquid

surfactant a compound that can reduce the surface tension of a solvent; surfactants have both a hydrophobic part and a hydrophilic part

synthesis reaction a reaction in which two reactants combine to make a larger or more complex product; general pattern: $A + B \rightarrow AB$

T

temperature a measure of the average kinetic energy of the entities of a substance

theoretical knowledge knowledge that explains scientific observations

theoretical yield the amount or mass of product predicted based on the stoichiometry of the chemical equation

theory an explanation or model based on observation, experimentation, and reasoning

thermal pollution an increase in water temperature, usually as a result of warm water being added to an aquatic ecosystem

titrant the solution in the burette during a titration

titration a procedure used to determine the concentration of a solution using a standardized solution

total ionic equation a chemical equation in which all highly soluble ionic compounds are written as dissociated ions

transpiration the evaporation of water from the leaves of a plant

U

universal gas constant (*R*) the constant in the ideal gas law equation that relates the pressure, volume, amount, and temperature of an ideal gas

unsaturated solution a solution in which more solute can dissolve at a given temperature and pressure

upcycling the process of converting an industrial material or product into something of similar or greater value

V

valence the charge of an ion; the combining capacity of an atom determined by the number of electrons that it will lose, add, or share when it reacts with other atoms

valence electron an electron in the outermost energy level or orbit

valence shell the outermost energy level or orbit of an atom or ion

van der Waals forces the weak forces of attraction between molecules, including dipole–dipole forces and London dispersion forces

W

water cycle the flow of water on, above, and below the surface of Earth

weak acid a substance that only partially ionizes in water; weak acids have weak acidic properties (e.g., moderate pH and mild reactivity)

Z

zero-sum rule the statement that the sum of the positive charges equals the sum of the negative charges in an ionic compound

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