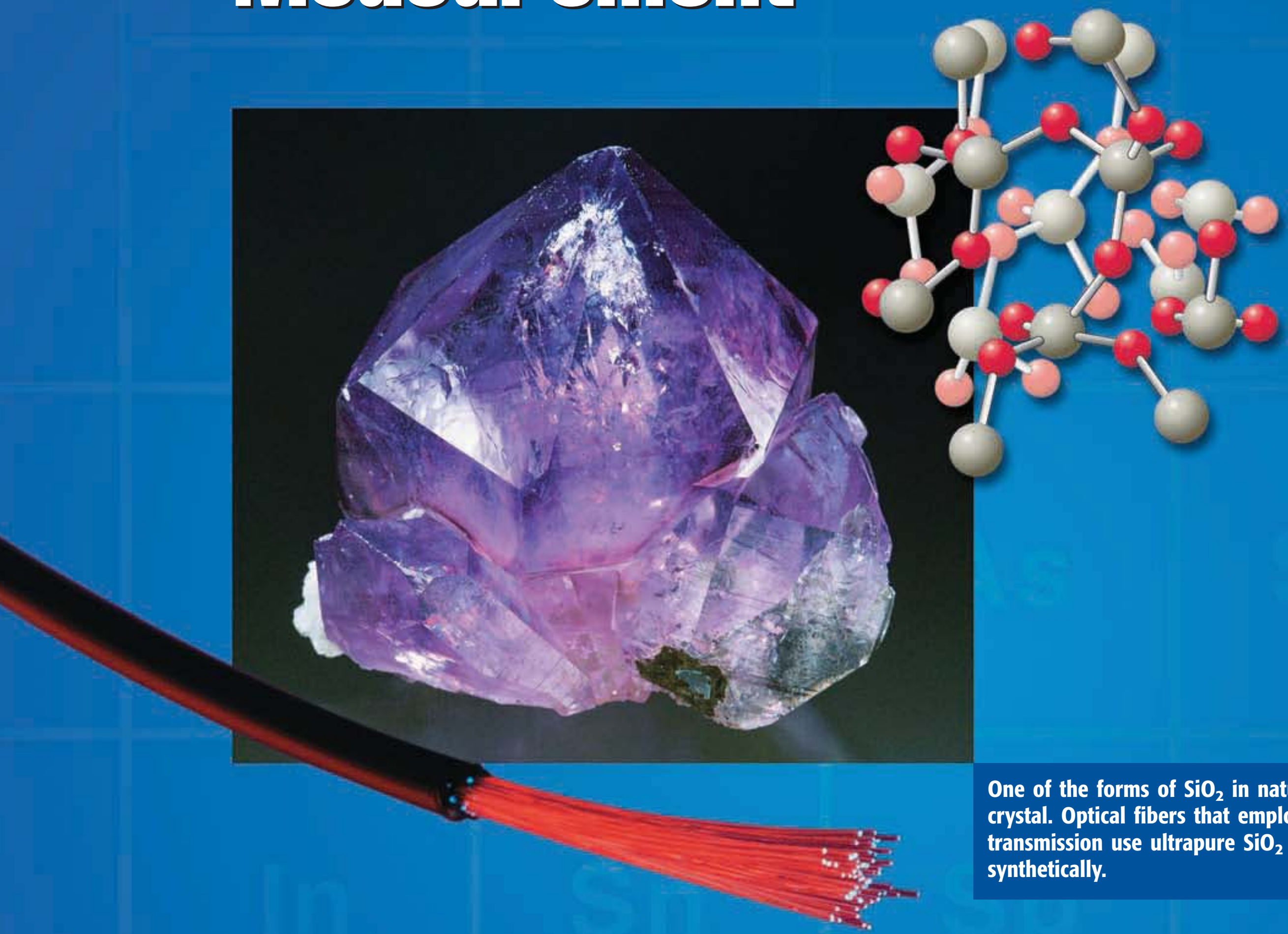
1 Chemistry and Measurement



One of the forms of SiO_2 in nature is the quartz crystal. Optical fibers that employ light for data transmission use ultrapure SiO_2 that is produced synthetically.

Contents and Concepts

We start by defining the science called chemistry and introducing some fundamental concepts.

Making and recording measurements of the properties and chemical behavior of matter is the foundation of chemistry.

In 1964 Barnett Rosenberg and his coworkers at Michigan State University were studying the effects of electricity on bacterial growth. They inserted platinum wire electrodes into a live bacterial culture and allowed an electric current to pass. After 1 to 2 hours, they noted that cell division in the bacteria stopped. The researchers were very surprised by this result, but even more surprised by the explanation. They were able to show that cell division was inhibited by a substance containing platinum, produced from the platinum electrodes by the electric current. A substance such as this one, the researchers thought, might be useful as an anticancer drug, because cancer involves runaway cell division. Later research confirmed this view, and today the platinum-containing substance *cisplatin* is a leading anticancer drug (Figure 1.1).

FIGURE 1.1
Barnett Rosenberg
Discoverer of the anticancer activity of cisplatin.

This story illustrates three significant reasons to study chemistry. First, chemistry has important practical applications. The development of lifesaving drugs is one, and a complete list would touch upon most areas of modern technology.

Second, chemistry is an intellectual enterprise, a way of explaining our material world. When Rosenberg and his coworkers saw that cell division in the bacteria had ceased, they systematically looked for the chemical substance that caused it to cease. They sought a chemical explanation for the occurrence.

Finally, chemistry figures prominently in other fields. Rosenberg's experiment began as a problem in biology; through the application of chemistry, it led to an advance in medicine. Whatever your career plans, you will find that your knowledge of chemistry is a useful intellectual tool for making important decisions.

See page 30 for the Media Summary.

An Introduction to Chemistry

All of the objects around you—this book, your pen or pencil, and the things of nature such as rocks, water, and plant and animal substances—constitute the *matter* of the universe. Each of the particular kinds of matter, such as a certain kind of paper or plastic or metal, is referred to as a *material.* We can define chemistry as the science of the composition and structure of materials and of the changes that materials undergo.

One chemist may hope that by understanding certain materials, he or she will be able to find a cure for a disease or a solution for an environmental ill. Another chemist may simply want to understand a phenomenon. Because chemistry deals with all materials, it is a subject of enormous breadth. It would be difficult to exaggerate the influence of chemistry on modern science and technology or on our ideas about our planet and the universe. In the section that follows, we will take a brief glimpse at modern chemistry and see some of the ways it has influenced technology, science, and modern thought.

1.1 Modern Chemistry: A Brief Glimpse

For thousands of years, human beings have fashioned natural materials into useful products. Modern chemistry certainly has its roots in this endeavor. After the discovery of fire, people began to notice changes in certain rocks and minerals exposed to high temperatures. From these observations came the development of ceramics, glass, and metals, which today are among our most useful materials. Dyes and medicines were other early products obtained from natural substances. For example, the ancient

Phoenicians extracted a bright purple dye, known as Tyrian purple, from a species of sea snail. One ounce of Tyrian purple required over 200,000 snails. Because of its brilliant hue and scarcity, the dye became the choice of royalty.

Although chemistry has its roots in early technology, chemistry as a field of study based on scientific principles came into being only in the latter part of the eighteenth century. Chemists began to look at the precise quantities of substances they used in their experiments. From this work came the central principle of modern chemistry: the materials around us are composed of exceedingly small particles called *atoms,* and the precise arrangement of these atoms into *molecules* or more complicated structures accounts for the many different characteristics of materials. Once chemists understood this central principle, they could begin to fashion molecules to order. They could *synthesize* molecules; that is, they could build large molecules from small ones. Tyrian purple, for example, was eventually synthesized from the simpler molecule aniline; see Figure 1.2. Chemists could also correlate molecular structure with the characteristics of materials and so begin to fashion materials with special characteristics.

The liquid-crystal displays (LCDs) that are used on everything from watches and cell phones to computer monitors and televisions are an example of an application that depends on the special characteristics of materials (Figure 1.3). The liquid crystals used in these displays are a form of matter intermediate in characteristics between those of liquids and those of solid crystals—hence the name. Many of these liquid crystals are composed of rodlike molecules that tend to align themselves something like the wood matches in a matchbox. The liquid crystals are held in alignment in layers by plates that have microscopic grooves. The molecules are attached to small electrodes or transistors. When the molecules are subjected to an electric charge from the transistor or electrode, they change alignment to point in a new direction. When they change direction, they change how light passes through their layer. When the liquid-crystal layer is combined with a light source and color filters, incremental changes of alignment of the molecules throughout the display allow for images that have high contrast and millions of colors. Figure 1.4 shows a model of one of the molecules that forms a liquid crystal; note the rodlike shape of the molecule. Chemists have designed many similar molecules for liquid-crystal applications. <

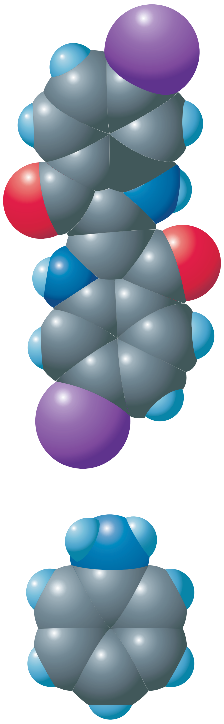


FIGURE 1.2 ▲

Molecular models of Tyrian purple and aniline

Tyrian purple *(top)* is a dye that was obtained by the early Phoenicians from a species of sea snail. The dye was eventually synthesized from aniline *(bottom)*.

Liquid crystals and liquid-crystal displays are described in the essay at the end of Section 11.7.

FIGURE 1.3 ▶

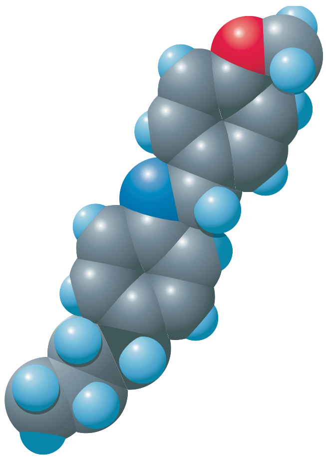
A cell phone that uses a liquid-crystal display

These liquid-crystal displays are used in a variety of electronic devices.

FIGURE 1.4 ▼

Model of a molecule that forms a liquid crystal

Note that the molecule has a rodlike shape.



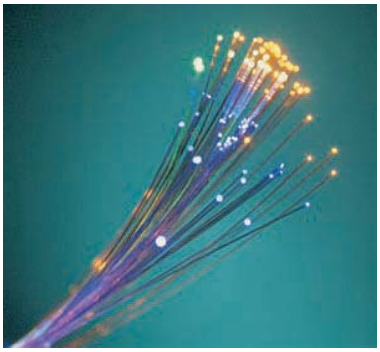


FIGURE 1.5
Optical fibers
A bundle of optical fibers that can be used to transmit data via pulses of light.

Chemists continue to develop new materials and to discover new properties of old ones. Electronics and communications, for example, have been completely transformed by technological advances in materials. Optical-fiber cables have replaced long-distance telephone cables made of copper wire. Optical fibers are fine threads of extremely pure glass. Because of their purity, these fibers can transmit laser light pulses for miles compared with only a few inches in ordinary glass. Not only is optical-fiber cable cheaper and less bulky than copper cable carrying the same information, but through the use of different colors of light, optical-fiber cable can carry voice, data, and video information at the same time (Figure 1.5). At the ends of an optical-fiber cable, devices using other new materials convert the light pulses to electrical signals and back, while computer chips constructed from still other materials process the signals.

Chemistry has also affected the way we think of the world around us. For example, biochemists and molecular biologists—scientists who study the molecular basis of living organisms—have made a remarkable finding: all forms of life appear to share many of the same molecules and molecular processes. Consider the information of inheritance, the genetic information that is passed on from one generation of organism to the next. Individual organisms, whether bacteria or human beings, store this information in a particular kind of molecule called deoxyribonucleic acid, or DNA (Figure 1.6).

DNA consists of two intertwined molecular chains; each chain consists of links of four different types of molecular pieces, or bases. Just as you record information on a page by stringing together characters (letters, numbers, spaces, and so on), an organism stores the information for reproducing itself in the order of these bases in its DNA. In a multicellular organism, such as a human being, every cell contains the same DNA.

One of our first projects will be to look at this central concept of chemistry, the atomic theory of matter. We will do that in the next chapter, but first we must lay the groundwork for this discussion. We will need some basic vocabulary to talk about science and to describe materials; then we will need to discuss measurement and units, because measurement is critical for quantitative work.

1.2 Experiment and Explanation

Experiment and explanation are the heart of chemical research. A chemist makes observations under circumstances in which variables, such as temperature and amounts of substances, can be controlled. An **experiment** is *an observation of natural phenomena carried out in a controlled manner so that the results can be duplicated and rational conclusions obtained.* In the chapter opening it was mentioned that Rosenberg studied the effects of electricity on bacterial growth. Temperature and amounts of nutrients in a given volume of bacterial medium are important variables in such experiments. Unless these variables are controlled, the work cannot be duplicated, nor can any reasonable conclusion be drawn.

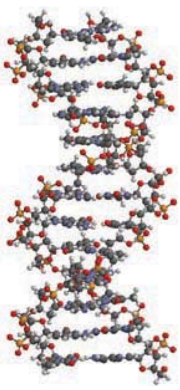
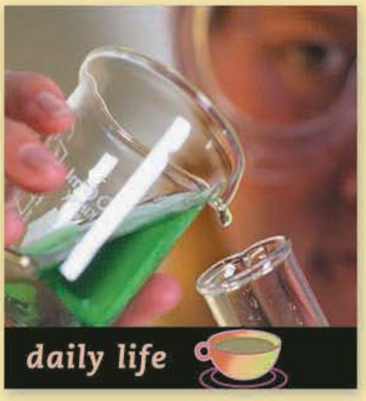


FIGURE 1.6
A side view of a fragment of a DNA molecule
DNA contains the hereditary information of an organism that is passed on from one generation to the next.

After a series of experiments, perhaps a researcher sees some relationship or regularity in the results. For instance, Rosenberg noted that in each experiment in which an electric current was passed through a bacterial culture by means of platinum wire electrodes, the bacteria ceased dividing. If the regularity or relationship is fundamental and we can state it simply, we call it a law. A **law** is *a concise statement or mathematical equation about a fundamental relationship or regularity of nature.* An example is the law of conservation of mass, which says that the mass, or quantity of matter, remains constant during any chemical change.

At some point in a research project, a scientist tries to make sense of the results by devising an explanation. Explanations help us organize knowledge and predict future events. A **hypothesis** is *a tentative explanation of some regularity of nature.* Having seen that bacteria ceased to divide when an electric current from platinum wire electrodes passed through the culture, Rosenberg was eventually able to propose the hypothesis that certain platinum compounds were responsible. If a hypothesis is

A Chemist Looks at . . .



The Birth of the Post-it Note®

Have you ever used a Post-it and wondered where the idea for those little sticky notes came from? You have a chemist to thank for their invention. The story of the Post-it Note illustrates how the creativity and insights of a scientist can result in a product that is as common in the office as the stapler or pen.

In the early 1970s, Art Fry, a 3M scientist, was standing in the choir at his church trying to keep track of all the little bits of paper that marked the music selections for the service. During the service, a number of the markers fell out of the music, making him lose his place. While standing in front of the congregation, he realized that he needed a bookmark that would stick to the book, wouldn't hurt the book, and could be easily detached. To make his plan work, he required an adhesive that would not *permanently* stick things together. Finding the appropriate adhesive was not as simple as it may seem, because most adhesives at that time were created to stick things together permanently.

Still thinking about his problem the next day, Fry consulted a colleague, Spencer Silver, who was studying adhesives at the 3M research labs. That study consisted of conducting a series of tests on a range of adhesives to determine the strength of the bond they formed. One of the adhesives that Silver created for the study was an adhesive that always remained sticky. Fry recognized that this adhesive was just what he needed for his bookmark. His first bookmark, invented the day after the initial idea, consisted of a strip of Silver's tacky adhesive applied to the edge of a piece of paper.

Part of Fry's job description at 3M was to spend time working on creative ideas such as his bookmark. As a result, he continued to experiment with the bookmark to improve its properties of sticking, detaching, and not hurting the surface to which it was attached. One day, while doing some paperwork, he wrote a question on one of his experimental strips of paper and sent it to his boss stuck to the top of a file folder. His boss then answered the question on the note and returned it attached to some other documents. During a later discussion over coffee, Fry and his boss realized that they had invented a new way for people to communicate: the Post-it Note was born. Today the Post-it Note is one of the top-selling office products in the United States.

■ See Problems 1.43 and 1.44.

to be useful, it should suggest new experiments that become tests of the hypothesis. Rosenberg could test his hypothesis by looking for the platinum compound and testing for its ability to inhibit cell division.

If a hypothesis successfully passes many tests, it becomes known as a theory. A **theory** is *a tested explanation of basic natural phenomena.* An example is the molecular theory of gases—the theory that all gases are composed of very small particles called molecules. This theory has withstood many tests and has been fruitful in suggesting many experiments. Note that we cannot prove a theory absolutely. It is always possible that further experiments will show the theory to be limited or that someone will develop a better theory. For example, the physics of the motion of objects devised by Isaac Newton withstood experimental tests for more than two centuries, until physicists discovered that the equations do not hold for objects moving near the speed of light. Later physicists showed that very small objects also do not follow Newton's equations. Both discoveries resulted in revolutionary developments in physics. The first led to the theory of relativity, the second to quantum mechanics, which has had an immense impact on chemistry.

The two aspects of science, experiment and explanation, are closely related. A scientist performs experiments and observes some regularity; someone explains this regularity and proposes more experiments; and so on. From his experiments, Rosenberg explained that certain platinum compounds inhibit cell division. This explanation led him to do new experiments on the anticancer activity of these compounds.

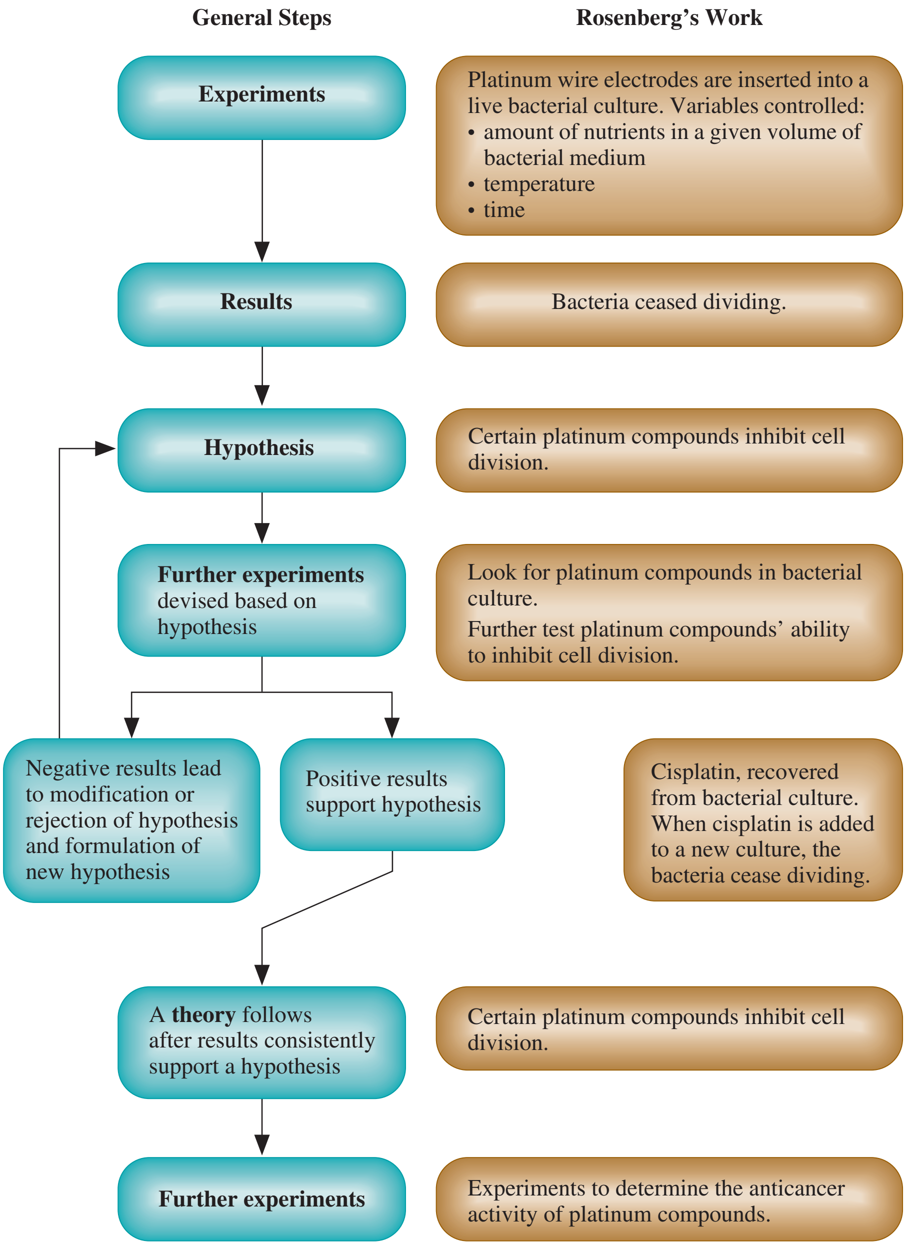


FIGURE 1.7

A representation of the scientific method
This flow diagram shows the general steps in the scientific method. At the right, Rosenberg's work on the development of an anticancer drug illustrates the steps.

FIGURE 1.8

Laboratory balance
A modern single-pan balance. The mass of the material on the pan appears on the digital readout.

The *general* process of advancing scientific knowledge through observation; the framing of laws, hypotheses, or theories; and the conducting of more experiments is called the *scientific method* (Figure 1.7). It is not a method for carrying out a *specific* research program, because the design of experiments and the explanation of results draw on the creativity and individuality of a researcher.

1.3 Law of Conservation of Mass 质量

Modern chemistry emerged in the eighteenth century, when chemists began to use the balance systematically as a tool in research. Balances measure **mass,** which is *the quantity of matter in a material* (Figure 1.8). **Matter** is the general term for the material things around us; we can define it as *whatever occupies space and can be perceived by our senses.*

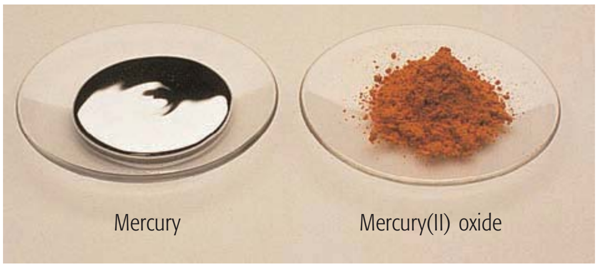


FIGURE 1.9

Heating mercury metal in air
Mercury metal reacts with oxygen to yield mercury(II) oxide. The color of the oxide varies from red to yellow, depending on the particle size.

FIGURE 1.10

Heated mercury(II) oxide
When you heat mercury(II) oxide, it decomposes to mercury and oxygen gas.

Antoine Lavoisier (1743–1794), a French chemist, was one of the first to insist on the use of the balance in chemical research. By weighing substances before and after chemical change, he demonstrated the **law of conservation of mass,** which states that *the total mass remains constant during a chemical change (chemical reaction).* <

Chemical reactions may involve a gain or loss of heat and other forms of energy. According to Einstein, mass and energy are equivalent. Thus, when energy is lost as heat, mass is also lost, but changes of mass in chemical reactions (billionths of a gram) are too small to detect.

In a series of experiments, Lavoisier applied the law of conservation of mass to clarify the phenomenon of burning, or combustion. He showed that when a material burns, a component of air (which he called oxygen) combines chemically with the material. For example, when the liquid metal mercury is heated in air, it burns or combines with oxygen to give a red-orange substance, whose modern name is mercury(II) oxide. We can represent the chemical change as follows:

$$\text{Mercury} + \text{oxygen} \longrightarrow \text{mercury(II) oxide}$$

The arrow means "is changed to." See Figure 1.9.

By strongly heating the red-orange substance, Lavoisier was able to decompose it to yield the original mercury and oxygen gas (Figure 1.10). The following example illustrates how the law of conservation of mass can be used to study this reaction.

Example 1.1 Using the Law of Conservation of Mass

You heat 2.53 grams of metallic mercury in air, which produces 2.73 grams of a red-orange residue. Assume that the chemical change is the reaction of the metal with oxygen in air.

$$\text{Mercury} + \text{oxygen} \longrightarrow \text{red-orange residue}$$

What is the mass of oxygen that reacts? When you strongly heat the red-orange residue, it decomposes to give back the mercury and release the oxygen, which you collect. What is the mass of oxygen you collect?

Problem Strategy You apply the law of conservation of mass to the reaction. According to this law, the total mass remains constant during a chemical reaction; that is,

$$\text{Mass of substances before reaction} = \text{mass of substances after reaction}$$

(*continued*)

(continued)

Solution From the law of conservation of mass,

$$\text{Mass of mercury} + \text{mass of oxygen} = \text{mass of red-orange residue}$$

Substituting, you obtain

$$2.53 \text{ grams} + \text{mass of oxygen} = 2.73 \text{ grams}$$

or

$$\text{Mass of oxygen} = (2.73 - 2.53) \text{ grams} = \mathbf{0.20\ grams}$$

The mass of oxygen collected when the red-orange residue decomposes equals the mass of oxygen that originally reacted **(0.20 grams).**

Answer Check Arithmetic errors account for many mistakes. You should always check your arithmetic, either by carefully redoing the calculation or, if possible, by doing the arithmetic in a slightly different way. Here, you obtained the answer by subtracting numbers. You can check the result by addition: the sum of the masses of mercury and oxygen, 2.53 + 0.20 grams, should equal the mass of the residue, 2.73 grams.

Exercise 1.1 You place 1.85 grams of wood in a vessel with 9.45 grams of air and seal the vessel. Then you heat the vessel strongly so that the wood burns. In burning, the wood yields ash and gases. After the experiment, you weigh the ash and find that its mass is 0.28 gram. What is the mass of the gases in the vessel at the end of the experiment?

■ See Problems 1.37, 1.38, 1.39, and 1.40.

Lavoisier set out his views on chemistry in his *Traité Élémentaire de Chimie* (*Basic Treatise on Chemistry*) in 1789. The book was very influential, especially among younger chemists, and set the stage for modern chemistry.

Before leaving this section, you should note the distinction between the terms *mass* and *weight* in precise usage. The weight of an object is the force of gravity exerted on it. The weight is proportional to the mass of the object divided by the square of the distance between the center of mass of the object and that of the earth. < Because the earth is slightly flattened at the poles, an object weighs more at the North Pole, where it is closer to the center of the earth, than at the equator. The mass of an object is the same wherever it is measured.

The force of gravity F between objects whose masses are m_1 and m_2 is Gm_1m_2/r^2, where G is the gravitational constant and r is the distance between the centers of mass of the two objects.

物质

1.4 Matter: Physical State and Chemical Constitution

We describe iron as a silvery-colored metal that melts at 1535°C (2795°F). Once we have collected enough descriptive information about many different kinds of matter, patterns emerge that suggest ways of classifying it. There are two principal ways of classifying matter: by its physical state as a solid, liquid, or gas and by its chemical constitution as an element, compound, or mixture.

Solids, Liquids, and Gases

Commonly, a given kind of matter exists in different physical forms under different conditions. Water, for example, exists as ice (solid water), as liquid water, and as steam

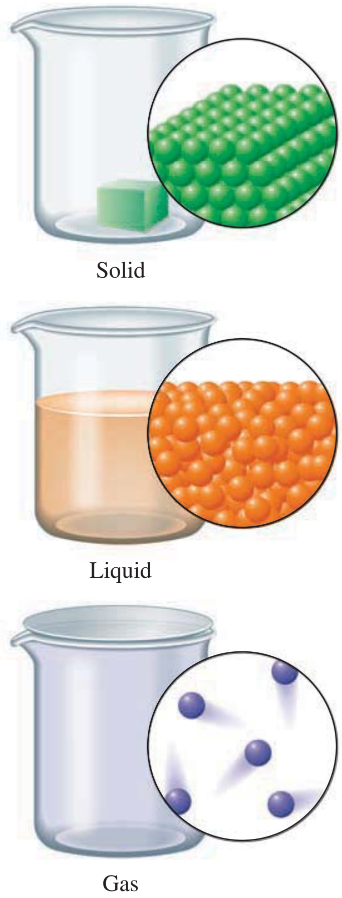


FIGURE 1.11

Molecular representations of solid, liquid, and gas

The top beaker contains a solid with a molecular view of the solid; the molecular view depicts the closely packed, immobile atoms that make up the solid structure. The middle beaker contains a liquid with a molecular view of the liquid; the molecular view depicts atoms that are close together but moving freely. The bottom beaker contains a gas with a molecular view of the gas; the molecular view depicts atoms that are far apart and moving freely.

(gaseous water) (Figure 1.11). The main identifying characteristic of solids is their rigidity: they tend to maintain their shapes when subjected to outside forces. Liquids and gases, however, are *fluids;* that is, they flow easily and change their shapes in response to slight outside forces.

What distinguishes a gas from a liquid is the characteristic of *compressibility* (and its opposite, *expansibility*). A gas is easily compressible, whereas a liquid is not. You can put more and more air into a tire, which increases only slightly in volume. In fact, a given quantity of gas can fill a container of almost any size. A small quantity would expand to fill the container; a larger quantity could be compressed to fill the same space. By contrast, if you were to try to force more liquid water into a closed glass bottle that was already full of water, it would burst.

These two characteristics, rigidity (or fluidity) and compressibility (or expansibility), can be used to frame definitions of the three common states of matter:

solid *the form of matter characterized by rigidity;* a solid is relatively incompressible and has fixed shape and volume.

liquid *the form of matter that is a relatively incompressible fluid;* a liquid has a fixed volume but no fixed shape.

gas *the form of matter that is an easily compressible fluid;* a given quantity of gas will fit into a container of almost any size and shape.

The term *vapor* is often used to refer to the gaseous state of any kind of matter that normally exists as a liquid or a solid.

These three forms of matter—solid, liquid, gas—comprise the common **states of matter.**

Elements, Compounds, and Mixtures

To understand how matter is classified by its chemical constitution, we must first distinguish between physical and chemical changes and between physical and chemical properties. A **physical change** is *a change in the form of matter but not in its chemical identity.* Changes of physical state are examples of physical changes. The process of dissolving one material in another is a further example of a physical change. For instance, you can dissolve sodium chloride (table salt) in water. The result is a clear liquid, like pure water, though many of its other characteristics are different from those of pure water. The water and sodium chloride in this liquid retain their chemical identities and can be separated by some method that depends on physical changes.

Distillation is one way to separate the sodium chloride and water components of this liquid. You place the liquid in a flask to which a device called a *condenser* is attached (see Figure 1.12 on page 10). The liquid in the flask is heated to bring it to a boil. (Boiling entails the formation of bubbles of the vapor in the body of the liquid.) Water vapor forms and passes from the flask into the cooled condenser, where the vapor changes back to liquid water. The liquid water is collected in another flask, called a *receiver.* The original flask now contains the solid sodium chloride. Thus, by means of physical changes (the change of liquid water to vapor and back to liquid), you have separated the sodium chloride and water that you had earlier mixed together.

A **chemical change,** or **chemical reaction,** is *a change in which one or more kinds of matter are transformed into a new kind of matter or several new kinds of matter.* The rusting of iron, during which iron combines with oxygen in the air to form a new material called rust, is a chemical change. The original materials (iron and oxygen) combine chemically and cannot be separated by any physical means. To recover the iron and oxygen from rust requires a chemical change or a series of chemical changes.

FIGURE 1.12

Separation by distillation
You can separate an easily vaporized liquid from another substance by distillation.

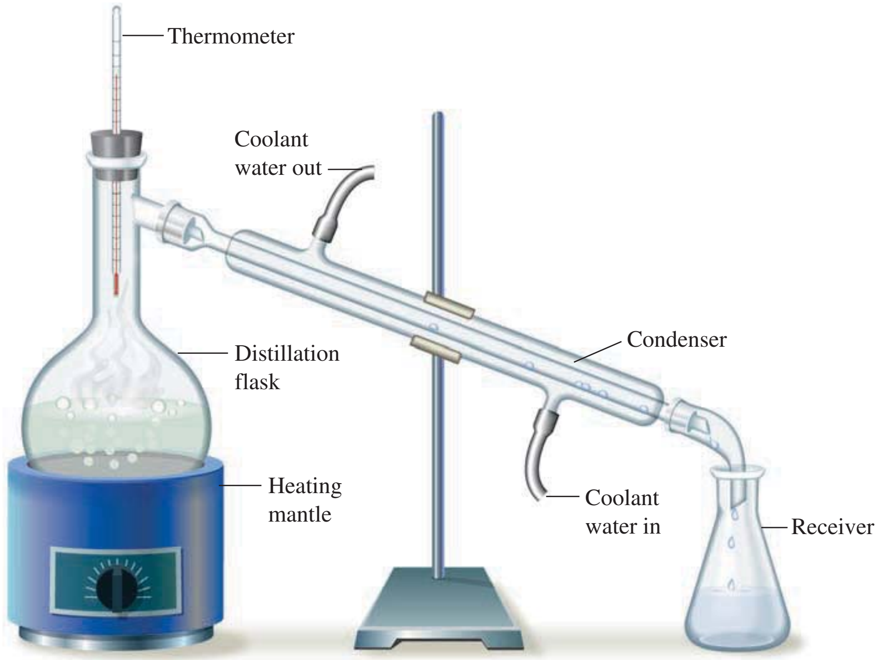


FIGURE 1.13

Reaction of sodium with water
Sodium metal flits around the water surface as it reacts briskly, giving off hydrogen gas. The other product is sodium hydroxide, which changes a substance added to the water (phenolphthalein) from colorless to pink.

We characterize or identify a material by its various properties, which may be either physical or chemical. A **physical property** is *a characteristic that can be observed for a material without changing its chemical identity.* Examples are physical state (solid, liquid, or gas), melting point, and color. A **chemical property** is *a characteristic of a material involving its chemical change.* A chemical property of iron is its ability to react with oxygen to produce rust.

Substances The various materials we see around us are either substances or mixtures of substances. A **substance** is *a kind of matter that cannot be separated into other kinds of matter by any physical process.* Earlier you saw that when sodium chloride is dissolved in water, it is possible to separate the sodium chloride from the water by the physical process of distillation. However, sodium chloride is itself a substance and cannot be separated by physical processes into new materials. Similarly, pure water is a substance.

No matter what its source, a substance always has the same characteristic properties. Sodium is a solid metal having a melting point of 98°C. The metal also reacts vigorously with water (Figure 1.13). No matter how sodium is prepared, it always has these properties. Similarly, whether sodium chloride is obtained by burning sodium in chlorine or from seawater, it is a white solid melting at 801°C.

Exercise 1.2 Potassium is a soft, silvery-colored metal that melts at 64°C. It reacts vigorously with water, with oxygen, and with chlorine. Identify all of the physical properties given in this description. Identify all of the chemical properties given.

■ See Problems 1.47, 1.48, 1.49, and 1.50.

FIGURE 1.14

Some elements

Center: Sulfur. *From upper right, clockwise:* Arsenic, iodine, magnesium, bismuth, mercury.

Elements Millions of substances have been characterized by chemists. Of these, a very small number are known as elements, from which all other substances are made. Lavoisier was the first to establish an experimentally useful definition of an element. He defined an **element** as *a substance that cannot be decomposed by any chemical reaction into simpler substances.* In 1789 Lavoisier listed 33 substances as elements, of which more than 20 are still so regarded. Today 116 elements are known. Some elements are shown in Figure 1.14. <

In Chapter 2, we will redefine an element in terms of atoms.

Compounds Most substances are compounds. A **compound** is *a substance composed of two or more elements chemically combined.* By the end of the eighteenth century, Lavoisier and others had examined many compounds and showed that all of them were composed of the elements in definite proportions by mass. Joseph Louis Proust (1754–1826), by his painstaking work, convinced the majority of chemists of the general validity of the **law of definite proportions** (also known as the **law of constant composition**): *a pure compound, whatever its source, always contains definite or constant proportions of the elements by mass.* For example, 1.0000 gram of sodium chloride always contains 0.3934 gram of sodium and 0.6066 gram of chlorine, chemically combined. Sodium chloride has definite proportions of sodium and chlorine; that is, it has constant or definite composition. <

It is now known that some compounds do not follow the law of definite proportions. These nonstoichiometric compounds, as they are called, are described briefly in Chapter 11.

Mixtures Most of the materials around us are mixtures. A **mixture** is *a material that can be separated by physical means into two or more substances.* Unlike a pure compound, a mixture has variable composition. When you dissolve sodium chloride in water, you obtain a mixture; its composition depends on the relative amount of sodium chloride dissolved. You can separate the mixture by the physical process of distillation. <

Chromatography, another example of a physical method used to separate mixtures, is described in the essay at the end of this section.

Mixtures are classified into two types. A **heterogeneous mixture** is *a mixture that consists of physically distinct parts, each with different properties.* Figure 1.15 shows a heterogeneous mixture of potassium dichromate and iron filings. Another example is salt and sugar that have been stirred together. If you were to look closely, you would see the separate crystals of sugar and salt. A **homogeneous mixture** (also known as a **solution**) is *a mixture that is uniform in its properties throughout given samples.* When sodium chloride is dissolved in water, you obtain a homogeneous mixture, or solution. Air is a gaseous solution, principally of two

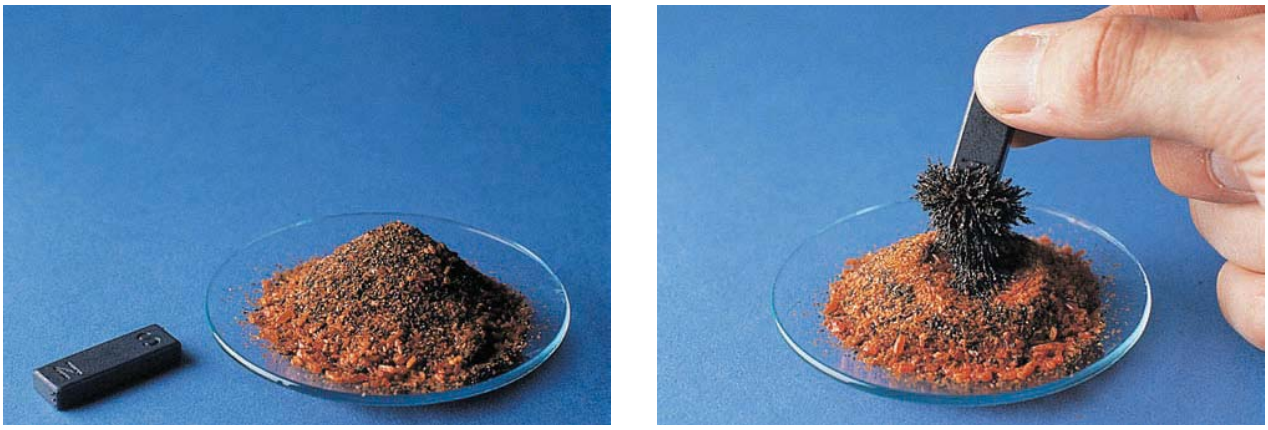


FIGURE 1.15

A heterogeneous mixture
Left: The mixture on the watch glass consists of potassium dichromate (orange crystals) and iron filings. *Right:* A magnet separates the iron filings from the mixture.

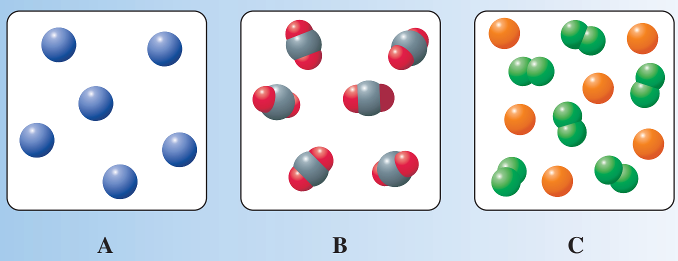
elementary substances, nitrogen and oxygen, which are physically mixed but not chemically combined.

A **phase** is *one of several different homogeneous materials present in the portion of matter under study.* A heterogeneous mixture of salt and sugar is said to be composed of two different phases: one of the phases is salt; the other is sugar. Similarly, ice cubes in water are said to be composed of two phases: one phase is ice; the other is liquid water. Ice floating in a solution of sodium chloride in water also consists of two phases, ice and the liquid solution. Note that a phase may be either a pure substance in a particular state or a solution in a particular state (solid, liquid, or gaseous). Also, the portion of matter under consideration may consist of several phases of the same substance or several phases of different substances.

Figure 1.16 summarizes the relationships among elements, compounds, and mixtures. Materials are either substances or mixtures. Substances can be mixed by physical processes, and other physical processes can be used to separate the mixtures into substances. Substances are either elements or compounds. Elements may react chemically to yield compounds, and compounds may be decomposed by chemical reactions into elements.

Concept Check 1.1

Matter can be represented as being composed of individual units. For example, the smallest individual unit of matter can be represented as a single circle, •, and chemical combinations of these units of matter as connected circles, ••, with each element represented by a different color. Using this model, place the appropriate label—element, compound, or mixture—on each container.



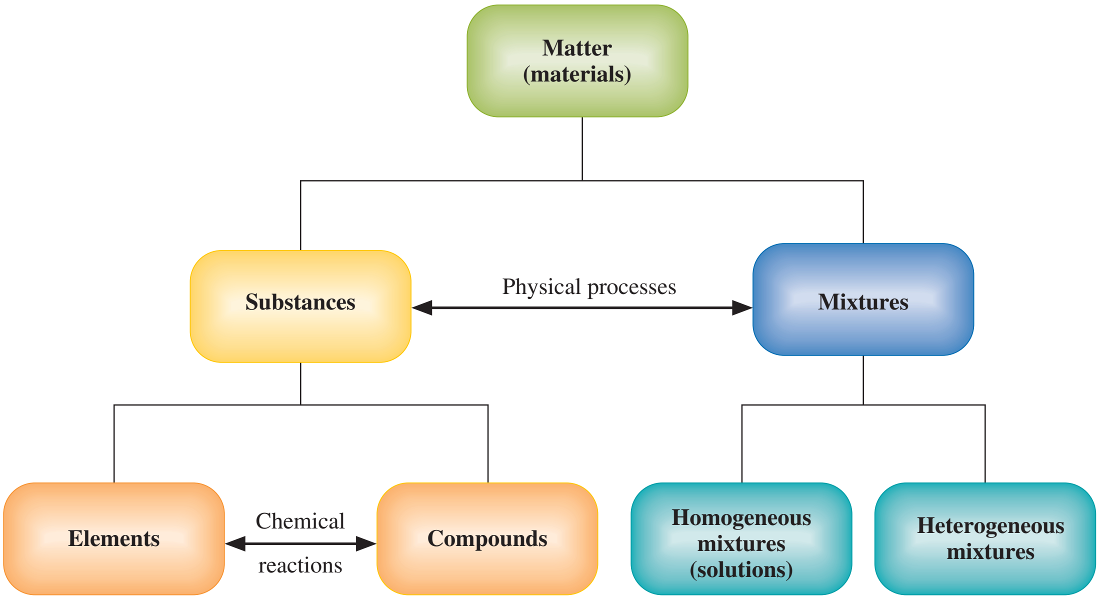


FIGURE 1.16

Relationships among elements, compounds, and mixtures
Mixtures can be separated by physical processes into substances, and substances can be combined physically into mixtures. Compounds can be separated by chemical reactions into their elements, and elements can be combined chemically to form compounds.

Physical Measurements

Chemists characterize and identify substances by their particular properties. To determine many of these properties requires physical measurements. Cisplatin, the substance featured in the chapter opening, is a yellow substance whose solubility in water (the amount that dissolves in a given quantity of water) is 0.252 gram in 100 grams of water. You can use the solubility, as well as other physical properties, to identify cisplatin. In a modern chemical laboratory, measurements often are complex, but many experiments begin with simple measurements of mass, volume, time, and so forth. In the next few sections, we will look at the measurement process. We first consider the concept of *precision* of a measurement.

1.5 Measurement and Significant Figures

2.54 centimeters = 1 inch

Measurement is the comparison of a physical quantity to be measured with a **unit** of measurement—that is, with *a fixed standard of measurement.* On a centimeter scale, the centimeter unit is the standard of comparison. < A steel rod that measures 9.12 times the centimeter unit has a length of 9.12 centimeters. To record the measurement, you must be careful to give both the *measured number* (9.12) and the *unit* (centimeters).

If you repeat a particular measurement, you usually do not obtain precisely the same result, because each measurement is subject to experimental error. The measured values vary slightly from one another. Suppose you perform a series of identical measurements of a quantity. The term **precision** refers to *the closeness of the set of values obtained from identical measurements of a quantity.* **Accuracy** is a related term; it refers to *the closeness of a single measurement to its true value.* To illustrate

Instrumental Methods

Separation of Mixtures by Chromatography

Chromatography is a group of similar separation techniques. Each depends on how fast a substance moves, in a stream of gas or liquid, past a stationary phase to which the substance may be slightly attracted. An example is provided by a simple experiment in *paper chromatography* (see Figure 1.17). In this experiment, a line of ink is drawn near one edge of a sheet of paper, and the paper is placed upright with this edge in a solution of methanol and water. As the solution creeps up the paper, the ink moves upward, separating into a series of different-colored bands that correspond to the different dyes in the ink. All the dyes are attracted to the wet paper fibers, but with different strengths of attraction. As the solution moves upward, the dyes less strongly attracted to the paper fibers move more rapidly.

The Russian botanist Mikhail Tswett was the first to understand the basis of chromatography and to apply it systematically as a method of separation. In 1906 Tswett separated pigments in plant leaves by *column chromatography*. He first dissolved the pigments from the leaves in petroleum ether, a liquid similar to gasoline. After packing a glass tube or column with powdered chalk, he poured the solution of pigments into the top of the column (see Figure 1.18). When he washed the column by pouring in more petroleum ether, it began to show distinct yellow and green bands. These bands, each containing a pure pigment, became well separated as they moved down the column, so that the pure pigments could be obtained. The name *chromatography* originates from this early separation of colored substances (the stem *chromato-* means "color"), although the technique is not limited to colored substances.

Gas chromatography (GC) is a more recent separation method. Here the moving stream is a gaseous mixture of vaporized substances plus a gas such as helium, which is called the *carrier*. The stationary material is either a solid or a liquid adhering to a solid, packed in a column. As the gas passes through the column, substances in the mixture are attracted differently to the stationary column packing and thus are separated. Gas chromatography is a rapid, small-scale method of separating mixtures. It is also important in the analysis of mixtures because the time it takes for a substance at a given temperature to travel through the column to a detector (called the *retention time*) is fixed. You can therefore use retention times to help identify substances. Figure 1.19 shows a gas chromatograph and a portion of a computer plot *(chromatogram)*. Each peak on the chromatogram corresponds to a specific substance. The peaks were automatically recorded by the instrument as the different substances in the mixture passed the detector. Chemists have analyzed complicated mixtures by gas chromatography. Analysis of chocolate, for example, shows that it contains over 800 flavor compounds.

FIGURE 1.17

An illustration of paper chromatography
A line of ink has been drawn along the lower edge of a sheet of paper. The dyes in the ink separate as a solution of methanol and water creeps up the paper.

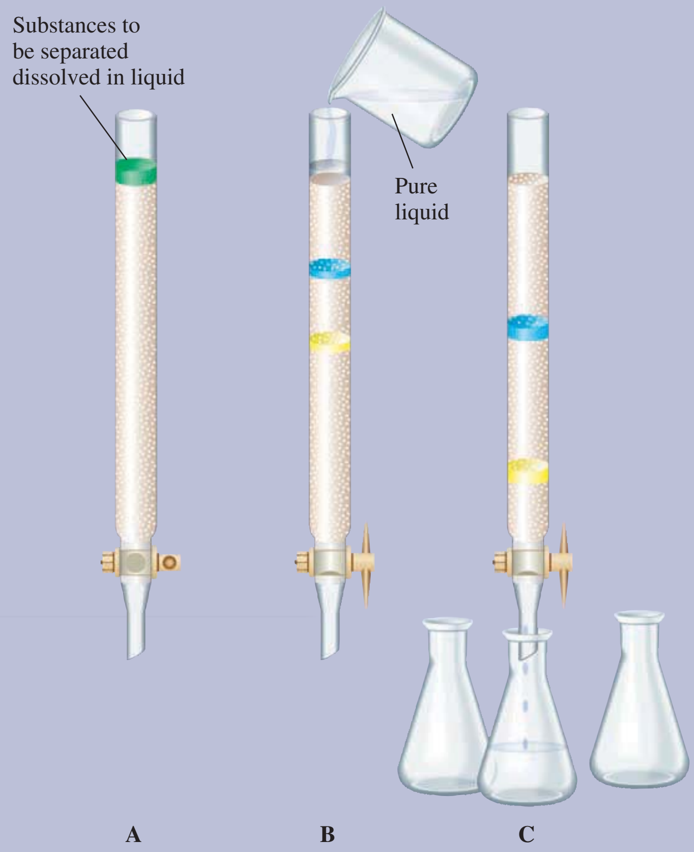


FIGURE 1.18

Column chromatography
(A) A solution containing substances to be separated is poured into the top of a column, which contains powdered chalk. *(B)* Pure liquid is added to the column, and the substances begin to separate into bands. *(C)* The substances separate further on the column. Each substance is collected in a separate flask as it comes off the column.

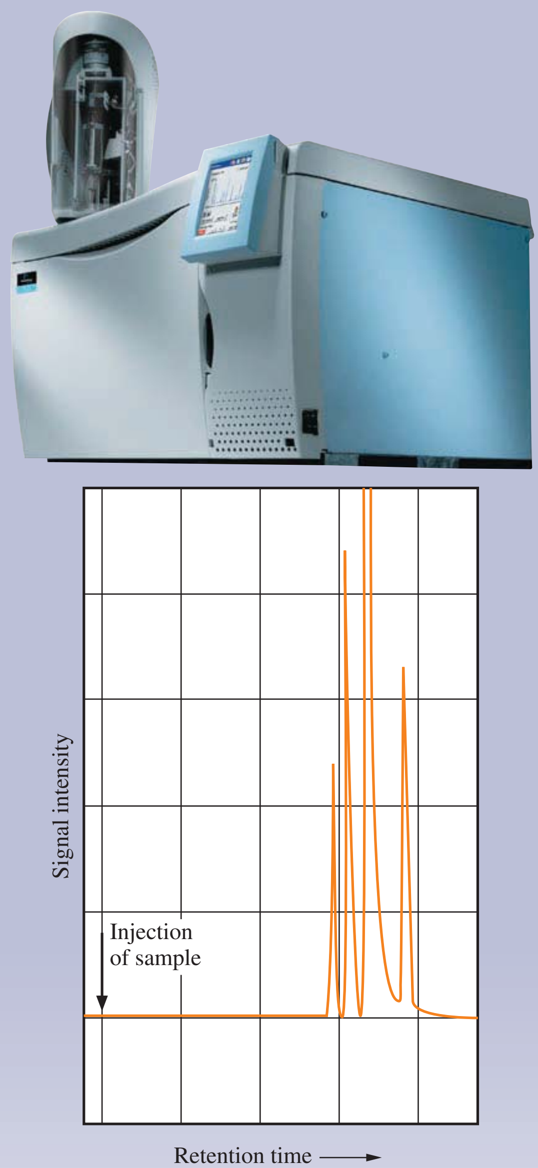


FIGURE 1.19

Gas chromatography
Top: A modern gas chromatograph. *Bottom:* This is a chromatogram of a hexane mixture, showing its separation into four different substances. Such hexane mixtures occur in gasoline; hexane is also used as a solvent to extract the oil from certain vegetable seeds.

■ See Problems 1.145 and 1.146.

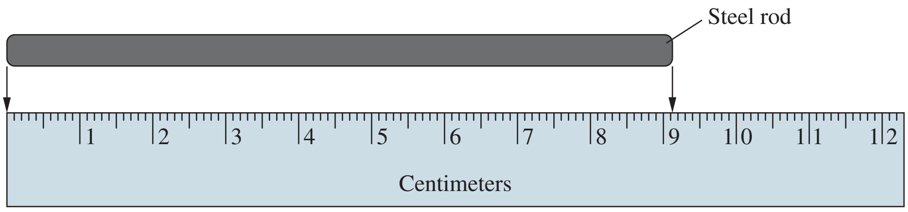


FIGURE 1.20

Precision of measurement with a centimeter ruler
The length of the rod is just over 9.1 cm. On successive measurements, we estimate the length by eye at 9.12, 9.11, and 9.13 cm. We record the length as between 9.11 cm and 9.13 cm.

the idea of precision, consider a simple measuring device, the centimeter ruler. In Figure 1.20, a steel rod has been placed near a ruler subdivided into tenths of a centimeter. You can see that the rod measures just over 9.1 cm (cm = centimeter). With care, it is possible to estimate by eye to hundredths of a centimeter. Here you might give the measurement as 9.12 cm. Suppose you measure the length of this rod twice more. You find the values to be 9.11 cm and 9.13 cm. Thus, you record the length of the rod as being somewhere between 9.11 cm and 9.13 cm. The spread of values indicates the precision with which a measurement can be made by this centimeter ruler. <

Measurements that are of high precision are usually accurate. It is possible, however, to have a systematic error in a measurement. Suppose that, in calibrating a ruler, the first centimeter is made too small by 0.1 cm. Then, although the measurements of length on this ruler are still precise to 0.01 cm, they are accurate to only 0.1 cm.

To indicate the precision of a measured number (or result of calculations on measured numbers), we often use the concept of significant figures. **Significant figures** are *those digits in a measured number (or in the result of a calculation with measured numbers) that include all certain digits plus a final digit having some uncertainty.* When you measured the rod, you obtained the values 9.12 cm, 9.11 cm, and 9.13 cm. You could report the result as the average, 9.12 cm. The first two digits (9.1) are certain; the next digit (2) is estimated, so it has some uncertainty. It would be incorrect to write 9.120 cm for the length of the rod. This would say that the last digit (0) has some uncertainty but that the other digits (9.12) are certain, which is not true.

Number of Significant Figures

The term **number of significant figures** refers to *the number of digits reported for the value of a measured or calculated quantity, indicating the precision of the value.* Thus, there are three significant figures in 9.12 cm, whereas 9.123 cm has four. To count the number of significant figures in a given measured quantity, you observe the following rules:

1. All digits are significant except zeros at the beginning of the number and possibly terminal zeros (one or more zeros at the end of a number). Thus, 9.12 cm, 0.912 cm, and 0.00912 cm all contain three significant figures.
2. Terminal zeros ending at the right of the decimal point are significant. Each of the following has three significant figures: 9.00 cm, 9.10 cm, 90.0 cm.
3. Terminal zeros in a number without an explicit decimal point may or may not be significant. If someone gives a measurement as 900 cm, you do not know whether one, two, or three significant figures are intended. If the person writes 900. cm (note the decimal point), the zeros are significant. More generally, you can remove any uncertainty in such cases by expressing the measurement in scientific notation.

Scientific notation is *the representation of a number in the form $A \times 10^n$, where A is a number with a single nonzero digit to the left of the decimal point and n is an integer, or whole number.* In scientific notation, the measurement 900 cm precise to two significant figures is written 9.0×10^2 cm. If precise to three significant figures, it is written 9.00×10^2 cm. Scientific notation is also convenient for expressing very large or very small quantities. It is much easier (and simplifies calculations) to write the speed of light as 3.00×10^8 (rather than as 300,000,000) meters per second. <

See Appendix A for a review of scientific notation.

FIGURE 1.21

Significant figures and calculators
Not all of the figures that appear on a calculator display are significant. In performing the calculation 100.0 × 0.0634 ÷ 25.31, the calculator display shows 0.250493875. We would report the answer as 0.250, however, because the factor 0.0634 has the least number of significant figures (three).

Significant Figures in Calculations

Measurements are often used in calculations. How do you determine the correct number of significant figures to report for the answer to a calculation? The following are two rules that we use:

1. **Multiplication and division.** When multiplying or dividing measured quantities, give as many significant figures in the answer as there are in the measurement with *the least number of significant figures.*
2. **Addition and subtraction.** When adding or subtracting measured quantities, give the same number of decimal places in the answer as there are in the measurement with *the least number of decimal places.*

Let us see how you apply these rules. Suppose you have a substance believed to be cisplatin, and, in an effort to establish its identity, you measure its solubility (the amount that dissolves in a given quantity of water). You find that 0.0634 gram of the substance dissolves in 25.31 grams of water. The amount dissolving in 100.0 grams is

$$100.0 \text{ grams of water} \times \frac{0.0634 \text{ gram cisplatin}}{25.31 \text{ grams of water}}$$

Performing the arithmetic on a pocket calculator (Figure 1.21), you get 0.250493875 for the numerical part of the answer (100.0 × 0.0634 ÷ 25.31). It would be incorrect to give this number as the final answer. The measurement 0.0634 gram has the least number of significant figures (three). Therefore, you report the answer to three significant figures—that is, 0.250 gram.

Now consider the addition of 184.2 grams and 2.324 grams. On a calculator, you find that 184.2 + 2.324 = 186.524. But because the quantity 184.2 grams has the least number of decimal places—one, whereas 2.324 grams has three—the answer is 186.5 grams.

Exact Numbers

So far we have discussed only numbers that involve uncertainties. However, you will also encounter exact numbers. An **exact number** is *a number that arises when you count items or sometimes when you define a unit.* For example, when you say there are 9 coins in a bottle, you mean exactly 9, not 8.9 or 9.1. Also, when you say there are 12 inches to a foot, you mean exactly 12. Similarly, the inch is defined to be exactly 2.54 centimeters. The conventions of significant figures do not apply to exact numbers. Thus, the 2.54 in the expression "1 inch equals 2.54 centimeters" should not be interpreted as a measured number with three significant figures. In effect, the 2.54 has an infinite number of significant figures, but of course it would be impossible to write out an infinite number of digits. You should make a mental note of any numbers in a calculation that are exact, because they have no effect on the number of significant figures in a calculation. The number of significant figures in a calculation result depends only on the numbers of significant figures in quantities having uncertainties. For example, suppose you want the total mass of 9 coins when each coin has a mass of 3.0 grams. The calculation is

$$3.0 \text{ grams} \times 9 = 27 \text{ grams}$$

You report the answer to two significant figures because 3.0 grams has two significant figures. The number 9 is exact and does not determine the number of significant figures.

Rounding

In reporting the solubility of your substance in 100.0 grams of water as 0.250 gram, you *rounded* the number you read off the calculator (0.2504938). **Rounding** is *the* *procedure of dropping nonsignificant digits in a calculation result and adjusting the last digit reported.* The general procedure is as follows: Look at the leftmost digit to be dropped. Then . . .

1. If this digit is 5 or greater, add 1 to the last digit to be retained and drop all digits farther to the right. Thus, rounding 1.2151 to three significant figures gives 1.22.
2. If this digit is less than 5, simply drop it and all digits farther to the right. Rounding 1.2143 to three significant figures gives 1.21.

In doing a calculation of two or more steps, it is desirable to retain nonsignificant digits for intermediate answers. This ensures that accumulated small errors from rounding do not appear in the final result. If you use a calculator, you can simply enter numbers one after the other, performing each arithmetic operation and rounding only the final answer. < To keep track of the correct number of significant figures, you may want to record intermediate answers with a line under the last significant figure, as shown in the solution to part d of the following example.

The answers for Exercises and Problems will follow this practice. For most Example Problems where intermediate answers are reported, all answers, intermediate and final, will be rounded.

Example 1.2 Using Significant Figures in Calculations

Perform the following calculations and round the answers to the correct number of significant figures (units of measurement have been omitted).

a. $\dfrac{2.568 \times 5.8}{4.186}$ b. $5.41 - 0.398$

c. $3.38 - 3.01$ d. $4.18 - 58.16 \times (3.38 - 3.01)$

Problem Strategy Parts a, b, and c are relatively straightforward; use the rule for significant figures that applies in each case. In a multistep calculation as in d, proceed step by step, applying the relevant rule to each step. First, do the operations within parentheses. By convention, you perform multiplications and divisions before additions and subtractions. Do not round intermediate answers, but do keep track of the least significant digit—say, by underlining it.

Solution a. The factor 5.8 has the fewest significant figures; therefore, the answer should be reported to two significant figures. Round the answer to **3.6.** b. The number with the least number of decimal places is 5.41. Therefore, round the answer to two decimal places, to **5.01.** c. The answer is **0.37.** Note how you have lost one significant figure in the subtraction. d. You first do the subtraction within parentheses, underlining the least significant digits.

$$4.1\underline{8} - 58.1\underline{6} \times (3.3\underline{8} - 3.0\underline{1}) = 4.1\underline{8} - 58.1\underline{6} \times 0.3\underline{7}$$

Following convention, you do the multiplication before the subtraction.

$$4.1\underline{8} - 58.1\underline{6} \times 0.3\underline{7} = 4.1\underline{8} - 2\underline{1}.5192 = -1\underline{7}.3392$$

The final answer is **−17.**

Answer Check When performing any calculation, you should always report answers to the correct number of significant figures. If a set of calculations involves *only* multiplication and division, it is not necessary to track the significant figures through intermediate calculations; you just need to use the measurement with the least number of significant figures as the guide. However, when addition or subtraction is involved, make sure that you have noted the correct number of significant figures in each calculated step in order to ensure that your answer is correctly reported.

Exercise 1.3 Give answers to the following arithmetic setups. Round to the correct number of significant figures.

a. $\dfrac{5.61 \times 7.891}{9.1}$ b. $8.91 - 6.435$

c. $6.81 - 6.730$ d. $38.91 \times (6.81 - 6.730)$

■ See Problems 1.61 and 1.62.

Concept Check 1.2

a. When you report your weight to someone, how many significant figures do you typically use?

b. What is your weight with two significant figures?

c. Indicate your weight and the number of significant figures you would obtain if you weighed yourself on a truck scale that can measure in 50 kg or 100 lb increments.

1.6 SI Units

The first measurements were probably based on the human body (the length of the foot, for example). In time, fixed standards developed, but these varied from place to place. Each country or government (and often each trade) adopted its own units. As science became more quantitative in the seventeenth and eighteenth centuries, scientists found that the lack of standard units was a problem. < They began to seek a simple, international system of measurement. In 1791 a study committee of the French Academy of Sciences devised such a system. Called the *metric system,* it became the official system of measurement for France and was soon used by scientists throughout the world. Most nations have since adopted the metric system or, at least, have set a schedule for changing to it.

In the system of units that Lavoisier used in the eighteenth century, there were 9216 grains to the pound (the livre). English chemists of the same period used a system in which there were 7000 grains to the pound—unless they were trained as apothecaries, in which case there were 5760 grains to the pound!

SI Base Units and SI Prefixes

In 1960 the General Conference of Weights and Measures adopted the **International System of units** (or **SI,** after the French *le Système International d'Unités*), which is *a particular choice of metric units.* This system has seven **SI base units,** *the SI units from which all others can be derived.* Table 1.1 lists these base units and the symbols used to represent them. In this chapter, we will discuss four base quantities: length, mass, time, and temperature. <

The amount of substance is discussed in Chapter 3, and the electric current (ampere) is introduced in Chapter 19. Luminous intensity will not be used in this book.

One advantage of any metric system is that it is a decimal system. In SI, a larger or smaller unit for a physical quantity is indicated by an **SI prefix,** which is *a prefix used in the International System to indicate a power of 10.* For example, the base unit of length in SI is the meter (somewhat longer than a yard), and 10^{-2} meter is called a *centi*meter. Thus 2.54 centimeters equals 2.54×10^{-2} meters. The SI prefixes used in this book are presented in Table 1.2.

TABLE 1.1 SI Base Units

Quantity	Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Temperature	kelvin	K
Amount of substance	mole	mol
Electric current	ampere	A
Luminous intensity	candela	cd

TABLE 1.2 Selected SI Prefixes

Prefix	Multiple	Symbol
mega	10^6	M
kilo	10^3	k
deci	10^{-1}	d
centi	10^{-2}	c
milli	10^{-3}	m
micro	10^{-6}	μ*
nano	10^{-9}	n
pico	10^{-12}	p

*Greek letter mu, pronounced "mew."

Length, Mass, and Time

The meter was originally defined in terms of a standard platinum–iridium bar kept at Sèvres, France. In 1983, the meter was defined as the distance traveled by light in a vacuum in 1/299,792,458 seconds.

The **meter (m)** is *the SI base unit of length.* < By combining it with one of the SI prefixes, you can get a unit of appropriate size for any length measurement. For the very small lengths used in chemistry, the nanometer (nm; 1 nanometer = 10^{-9} m) or the picometer (pm; 1 picometer = 10^{-12} m) is an acceptable SI unit. *A non-SI unit of length* traditionally used by chemists is the **angstrom (Å),** which equals 10^{-10} m. (An oxygen atom, one of the minute particles of which the substance oxygen is composed, has a diameter of about 1.3 Å. If you could place oxygen atoms adjacent to one another, you could line up over 75 million of them in 1 cm.)

The present standard of mass is the platinum–iridium kilogram mass kept at the International Bureau of Weights and Measures in Sèvres, France.

The **kilogram (kg)** is *the SI base unit of mass,* equal to about 2.2 pounds. < This is an unusual base unit in that it contains a prefix. In forming other SI mass units, prefixes are added to the word *gram* (g) to give units such as the *milli*gram (mg; 1 mg = 10^{-3} g).

The **second (s)** is *the SI base unit of time* (Figure 1.22). Combining this unit with prefixes such as *milli-, micro-, nano-,* and *pico-,* you create units appropriate for measuring very rapid events. The time required for the fastest chemical processes is about a picosecond, which is on the order of how fast supercomputers can perform a single calculation. When you measure times much longer than a few hundred seconds, you revert to *minutes* and *hours,* an obvious exception to the prefix–base format of the International System.

Exercise 1.4 Express the following quantities using an SI prefix and a base unit. For instance, 1.6×10^{-6} m = 1.6 μm. A quantity such as 0.000168 g could be written 0.168 mg or 168 μg.

a. 1.84×10^{-9} m b. 5.67×10^{-12} s c. 7.85×10^{-3} g

d. 9.7×10^{3} m e. 0.000732 s f. 0.000000000154 m

■ See Problems 1.65 and 1.66.

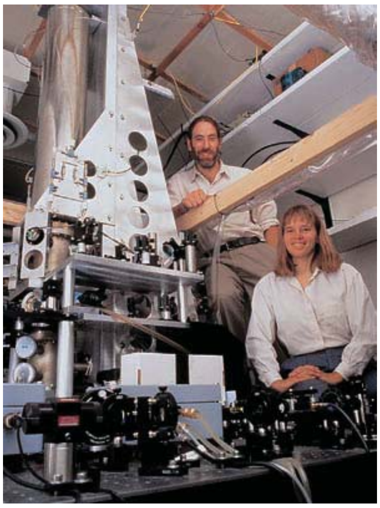


FIGURE 1.22 ▲

NIST F-1 cesium fountain clock

First put into service in 1999 at the National Institute of Standards and Technology (NIST) in Boulder, Colorado, this clock represents a new generation of super-accurate atomic clocks. The clock is so precise that it loses only one second every 20 million years.

Temperature

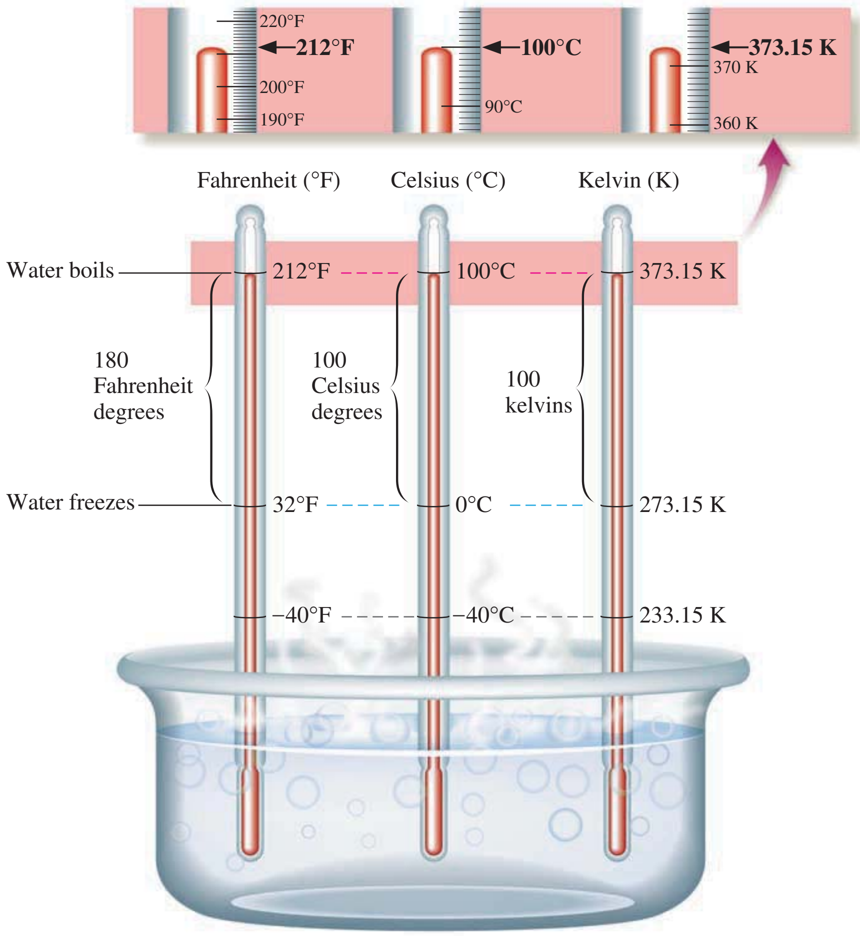
Temperature is difficult to define precisely, but we all have an intuitive idea of what we mean by it. It is a measure of "hotness." A hot object placed next to a cold one becomes cooler, while the cold object becomes hotter. Heat energy passes from a hot object to a cold one, and the quantity of heat passed between the objects depends on the difference in temperature between the two. Therefore, temperature and heat are different, but related, concepts.

A thermometer is a device for measuring temperature. The common type consists of a glass capillary containing a column of liquid whose length varies with temperature. A scale alongside the capillary gives a measure of the temperature. The **Celsius scale** (formerly the centigrade scale) is *the temperature scale in general scientific use.* On this scale, the freezing point of water is 0°C and the boiling point of water at normal barometric pressure is 100°C. However, *the SI base unit of temperature* is the **kelvin (K),** a unit on an *absolute temperature* scale. (See the first margin note on the next page.) On any absolute scale, the lowest temperature that can be attained theoretically is zero. The Celsius and the Kelvin scales have equal-size units (that is, a change of 1°C is equivalent to a change of 1 K), where 0°C is a temperature equivalent to 273.15 K. Thus, it is easy to convert from one scale to the other, using the formula

$$T_K = \left(t_C \times \frac{1\ \mathrm{K}}{1^\circ\mathrm{C}}\right) + 273.15\ \mathrm{K}$$

where T_K is the temperature in kelvins and t_C is the temperature in degrees Celsius. A temperature of 20°C (about room temperature) equals 293 K.

FIGURE 1.23
Comparison of temperature scales
Room temperature is about 68°F, 20°C, and 293 K. Water freezes at 32°F, 0°C, and 273.15 K. Water boils under normal pressure at 212°F, 100°C, and 373.15 K.



Note that the degree sign (°) is not used with the Kelvin scale, and the unit of temperature is simply the kelvin (not capitalized).

The Fahrenheit scale is at present the common temperature scale in the United States. Figure 1.23 compares Kelvin, Celsius, and Fahrenheit scales. As the figure shows, 0°C is the same as 32°F (both exact), and 100°C corresponds to 212°F (both exact). Therefore, there are 212 − 32 = 180 Fahrenheit degrees in the range of 100 Celsius degrees. That is, there are exactly 9 Fahrenheit degrees for every 5 Celsius degrees. Knowing this, and knowing that 0°C equals 32°F, we can derive a formula to convert degrees Celsius to degrees Fahrenheit.

These two temperature conversion formulas are often written K = °C + 273.15 and °F = (1.8 × °C) + 32. Although these yield the correct number, they do not take into account the units.

$$t_F = \left(t_C \times \frac{9°\text{F}}{5°\text{C}}\right) + 32°\text{F}$$

where t_F is the temperature in Fahrenheit and t_C is the temperature in Celsius. By rearranging this, we can obtain a formula for converting degrees Fahrenheit to degrees Celsius:

$$t_C = \frac{5°\text{C}}{9°\text{F}} \times (t_F - 32°\text{F})$$

Example 1.3 Converting from One Temperature Scale to Another

The hottest place on record in North America is Death Valley in California. It reached a temperature of 134°F in 1913. What is this temperature reading in degrees Celsius? in kelvins?

(*continued*)

(continued)

Problem Strategy This calculation involves conversion of a temperature in degrees Fahrenheit (°F) to degrees Celsius (°C) and kelvins (K). This is a case where using formulas is the most reliable method for arriving at the answer.

Solution Substituting, we find that

$$t_C = \frac{5°C}{9°F} \times (t_F - 32°F) = \frac{5°C}{9°F} \times (134°F - 32°F) = \mathbf{56.7°C}$$

In kelvins,

$$T_K = \left(t_C \times \frac{1\ K}{1°C}\right) + 273.15\ K = \left(56.7°C \times \frac{1\ K}{1°C}\right) + 273.15\ K = \mathbf{329.9\ K}$$

Answer Check An easy way to check temperature conversions is to use the appropriate equation and convert your answer back to the original temperature to see whether it matches.

Exercise 1.5 a. A person with a fever has a temperature of 102.5°F. What is this temperature in degrees Celsius? b. A cooling mixture of dry ice and isopropyl alcohol has a temperature of −78°C. What is this temperature in kelvins?

■ See Problems 1.69, 1.70, 1.71, and 1.72.

Concept Check 1.3

a. Estimate and express the length of your leg in an appropriate metric unit.

b. What would be a reasonable height, in meters, of a three-story building?

c. How would you be feeling if your body temperature was 39°C?

d. Would you be comfortable sitting in a room at 23°C in a short-sleeved shirt?

1.7 Derived Units

Once base units have been defined for a system of measurement, you can derive other units from them. You do this by using the base units in equations that define other physical quantities. For example, *area* is defined as length times length. Therefore,

$$\text{SI unit of area} = (\text{SI unit of length}) \times (\text{SI unit of length})$$

From this, you see that the SI unit of area is meter × meter, or m^2. Similarly, *speed* is defined as the rate of change of distance with time; that is, speed = distance/time. Consequently,

$$\text{SI unit of speed} = \frac{\text{SI unit of distance}}{\text{SI unit of time}}$$

The SI unit of speed is meters per second (that is, meters divided by seconds). The unit is symbolized m/s or $m \cdot s^{-1}$. The unit of speed is an example of an **SI derived unit,** which is *a unit derived by combining SI base units.* Table 1.3 defines a number of derived units. Volume and density are discussed in this section; pressure and energy are discussed later (in Sections 5.1 and 6.1, respectively).

Volume

Volume is defined as length cubed and has the SI unit of cubic meter (m^3). This is too large a unit for normal laboratory work, so we use either cubic decimeters (dm^3)

TABLE 1.3 Derived Units

Quantity	Definition of Quantity	SI Unit
Area	Length squared	m^2
Volume	Length cubed	m^3
Density	Mass per unit volume	kg/m^3
Speed	Distance traveled per unit time	m/s
Acceleration	Speed changed per unit time	m/s^2
Force	Mass times acceleration of object	$kg \cdot m/s^2$ (= newton, N)
Pressure	Force per unit area	$kg/(m \cdot s^2)$ (= pascal, Pa)
Energy	Force times distance traveled	$kg \cdot m^2/s^2$ (= joule, J)

or cubic centimeters (cm^3, also written cc). Traditionally, chemists have used the **liter (L),** which is *a unit of volume equal to a cubic decimeter* (approximately one quart). In fact, most laboratory glassware (Figure 1.24) is calibrated in liters or milliliters (1000 mL = 1 L). Because 1 dm equals 10 cm, a cubic decimeter, or one liter, equals $(10 \text{ cm})^3 = 1000 \text{ cm}^3$. Therefore, a milliliter equals a cubic centimeter. In summary

$$1 \text{ L} = 1 \text{ dm}^3 \qquad \text{and} \qquad 1 \text{ mL} = 1 \text{ cm}^3$$

Density

The **density** of an object is its *mass per unit volume.* You can express this as

$$d = \frac{m}{V}$$

where d is the density, m is the mass, and V is the volume. Suppose an object has a mass of 15.0 g and a volume of 10.0 cm^3. Substituting, you find that

$$d = \frac{15.0 \text{ g}}{10.0 \text{ cm}^3} = 1.50 \text{ g/cm}^3$$

The density of the object is 1.50 g/cm^3 (or 1.50 $g \cdot cm^{-3}$).

Density is an important characteristic property of a material. Water, for example, has a density of 1.000 g/cm^3 at 4°C and a density of 0.998 g/cm^3 at 20°C. Lead has

FIGURE 1.24

Some laboratory glassware

Left to right: A 600-mL beaker; a 100-mL graduated cylinder; a 250-mL volumetric flask; a 250-mL Erlenmeyer flask. *Front:* A 5-mL pipet.

FIGURE 1.25 (*left*)

The relative densities of copper and mercury

Copper floats on mercury because the density of copper is less than that of mercury. (This is a copper penny; recently minted pennies are copper-clad zinc.)

FIGURE 1.26 (*right*)

Relative densities of some liquids

Shown are three liquids (dyed so that they will show up clearly): *ortho*-xylene, water, and 1,1,1-trichloroethane. Water (blue layer) is less dense than 1,1,1-trichloroethane (colorless) and floats on it. *Ortho*-xylene (top, golden layer) is less dense than water and floats on it.

The density of solid materials on earth ranges from about 1 g/cm^3 to 22.5 g/cm^3 (osmium metal). In the interior of certain stars, the density of matter is truly staggering. Black neutron stars—stars composed of neutrons, or atomic cores compressed by gravity to a superdense state—have densities of about 10^{15} g/cm^3.

a density of 11.3 g/cm^3 at 20°C. (Figures 1.25 and 1.26 dramatically show some relative densities.) < Oxygen gas has a density of 1.33×10^{-3} g/cm^3 at normal pressure and 20°C. (Like other gases under normal conditions, oxygen has a density that is about 1000 times smaller than those of liquids and solids.) Because the density is characteristic of a substance, it can be helpful in identifying it. Example 1.4 illustrates this point. Density can also be useful in determining whether a substance is pure. Consider a gold bar whose purity is questioned. The metals likely to be mixed with gold, such as silver or copper, have lower densities than gold. Therefore, an adulterated (impure) gold bar can be expected to be far less dense than pure gold.

Example 1.4 Calculating the Density of a Substance

A colorless liquid, used as a solvent (a liquid that dissolves other substances), is believed to be one of the following:

Substance	*Density (in g/mL)*
n-butyl alcohol	0.810
ethylene glycol	1.114
isopropyl alcohol	0.785
toluene	0.866

To identify the substance, a chemist determined its density. By pouring a sample of the liquid into a graduated cylinder, she found that the volume was 35.1 mL. She also found that the sample weighed 30.5 g. What was the density of the liquid? What was the substance?

Problem Strategy The solution to this problem lies in finding the density of the unknown substance. Once the density of the unknown substance is known, you can compare it to the list of known substances presented in the problem and look for a match. Density is the relationship of the mass of a substance per volume of that substance. Expressed as an equation, density is the mass divided by the volume: $d = m/v$.

Solution You substitute 30.5 g for the mass and 35.1 mL for the volume into the equation.

$$d = \frac{m}{V} = \frac{30.5\ \text{g}}{35.1\ \text{mL}} = \mathbf{0.869\ g/mL}$$

The density of the liquid equals that of **toluene** (within experimental error).

Answer Check Always be sure to report the density in the units used when performing the calculation. Density is not always reported in units of g/ml or g/cm^3, for example; gases are often reported with the units of g/L.

Exercise 1.6 A piece of metal wire has a volume of 20.2 cm^3 and a mass of 159 g. What is the density of the metal? We know that the metal is manganese, iron, or nickel, and these have densities of 7.21 g/cm^3, 7.87 g/cm^3, and 8.90 g/cm^3, respectively. From which metal is the wire made?

■ See Problems 1.73, 1.74, 1.75, and 1.76.

In addition to characterizing a substance, the density provides a useful relationship between mass and volume. For example, suppose an experiment calls for a certain mass of liquid. Rather than weigh the liquid on a balance, you might instead measure out the corresponding volume. Example 1.5 illustrates this idea.

Example 1.5 Using the Density to Relate Mass and Volume

An experiment requires 43.7 g of isopropyl alcohol. Instead of measuring out the sample on a balance, a chemist dispenses the liquid into a graduated cylinder. The density of isopropyl alcohol is 0.785 g/mL. What volume of isopropyl alcohol should he use?

Problem Strategy The formula that defines density provides a relationship between the density of a substance and the mass and volume of the substance: $d = m/V$. Because this relationship is an equation, it can be used to solve for one variable when the other two are known. Examining this problem reveals that the mass (m) and density (d) of the substance are known, so the equation can be used to solve for the volume (V) of the liquid.

Solution You rearrange the formula defining the density to obtain the volume.

$$V = \frac{m}{d}$$

Then you substitute into this formula:

$$V = \frac{43.7\ \text{g}}{0.785\ \text{g/mL}} = \mathbf{55.7\ mL}$$

Answer Check Note that the density of the alcohol is close to 1 g/mL, which is common for many substances. In cases like this, you can perform a very quick check of your answer, because the numerical value for the volume of the liquid should not be much different from the numerical value of the mass of the liquid.

Exercise 1.7 Ethanol (grain alcohol) has a density of 0.789 g/cm^3. What volume of ethanol must be poured into a graduated cylinder to equal 30.3 g?

■ See Problems 1.77, 1.78, 1.79, and 1.80.

Concept Check 1.4

You are working in the office of a precious metals buyer. A miner brings you a nugget of metal that he claims is gold. You suspect that the metal is a form of "fool's gold," called marcasite, that is composed of iron and sulfur. In the back of your office, you have a chunk of pure gold. What simple experiments could you perform to determine whether the miner's nugget is gold?

1.8 Units and Dimensional Analysis (Factor-Label Method)

In performing numerical calculations with physical quantities, it is good practice to enter each quantity as a number with its associated unit. Both the numbers and the units are then carried through the indicated algebraic operations. The advantages of this are twofold:

1. The units for the answer will come out of the calculations.
2. If you make an error in arranging factors in the calculation (for example, if you use the wrong formula), this will become apparent because the final units will be nonsense.

Dimensional analysis (or the **factor-label method**) is *the method of calculation in which one carries along the units for quantities.* As an illustration, suppose you want to find the volume V of a cube, given s, the length of a side of the cube. Because

$V = s^3$, if $s = 5.00$ cm, you find that $V = (5.00\text{ cm})^3 = 5.00^3\text{ cm}^3$. There is no guesswork about the unit of volume here; it is cubic centimeters (cm^3).

Suppose, however, that you wish to express the volume in liters (L), a metric unit that equals 10^3 cubic centimeters (approximately one quart). You can write this equality as

$$1\text{ L} = 10^3\text{ cm}^3$$

If you divide both sides of the equality by the right-hand quantity, you get

$$\frac{1\text{ L}}{10^3\text{ cm}^3} = \frac{\cancel{10^3\text{ cm}^3}}{\cancel{10^3\text{ cm}^3}} = 1$$

Observe that you treat units in the same way as algebraic quantities. Note too that the right-hand side now equals 1 and no units are associated with it. Because it is always possible to multiply any quantity by 1 without changing that quantity, you can multiply the previous expression for the volume by the factor $1\text{ L}/10^3\text{ cm}^3$ without changing the actual volume. You are changing only the way you express this volume:

$$V = 5.00^3\,\cancel{\text{cm}^3} \times \underbrace{\frac{1\text{ L}}{10^3\,\cancel{\text{cm}^3}}}_{\substack{\text{converts}\\ \text{cm}^3\text{ to L}}} = 125 \times 10^{-3}\text{ L} = 0.125\text{ L}$$

It takes more room to explain conversion factors than it does to use them. With practice, you will be able to write the final conversion step without the intermediate algebraic manipulations outlined here.

The ratio $1\text{ L}/10^3\text{ cm}^3$ is called a **conversion factor** because it is *a factor equal to 1 that converts a quantity expressed in one unit to a quantity expressed in another unit.* < Note that the numbers in this conversion factor are *exact,* because 1 L equals exactly 1000 cm^3. Such exact conversion factors do not affect the number of significant figures in an arithmetic result. In the previous calculation, the quantity 5.00 cm (the measured length of the side) does determine or limit the number of significant figures.

The next two examples illustrate how the conversion-factor method may be used to convert one metric unit to another.

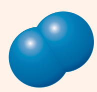
Example 1.6 Converting Units: Metric Unit to Metric Unit

Nitrogen gas is the major component of air. A sample of nitrogen gas in a glass bulb weighs 243 mg. What is this mass in SI base units of mass (kilograms)?

Problem Strategy This problem requires converting a mass in milligrams to kilograms. Finding the correct relationship for performing such a conversion directly in one step might be more difficult than doing the conversion in multiple steps. In this case, conversion factors for converting milligrams to grams and from grams to kilograms are relatively easy to derive, so it makes sense to do this conversion in two steps. First convert milligrams to grams; then convert grams to kilograms. To convert from milligrams to grams, note that the prefix *milli-* means 10^{-3}. To convert from grams to kilograms, note that the prefix *kilo-* means 10^3.

Solution Since $1\text{ mg} = 10^{-3}\text{ g}$, you can write

$$243\,\cancel{\text{mg}} \times \frac{10^{-3}\text{ g}}{1\,\cancel{\text{mg}}} = 2.43 \times 10^{-1}\text{ g}$$



Nitrogen molecular model

Then, because the prefix *kilo-* means 10^3, you write

$$1\text{ kg} = 10^3\text{ g}$$

and

$$2.43 \times 10^{-1}\,\cancel{\text{g}} \times \frac{1\text{ kg}}{10^3\,\cancel{\text{g}}} = \mathbf{2.43 \times 10^{-4}\ kg}$$

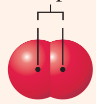
(*continued*)

(continued)

Note, however, that you can do the two conversions in one step, as follows:

$$243\ \cancel{\text{mg}} \times \underbrace{\frac{10^{-3}\ \cancel{\text{g}}}{1\ \cancel{\text{mg}}}}_{\substack{\text{converts}\\ \text{mg to g}}} \times \underbrace{\frac{1\ \text{kg}}{10^{3}\ \cancel{\text{g}}}}_{\substack{\text{converts}\\ \text{g to kg}}} = 2.43 \times 10^{-4}\ \text{kg}$$

Answer Check Often, mistakes lead to answers that are easily detected by giving them a *reasonableness check*. For example, if while solving this problem you mistakenly write the conversion factors in inverted order (although this should not happen if you use units in your calculation), the final answer will be very large. Or perhaps you inadvertently copy the exponent from your calculator with the wrong sign, giving the answer as 2.43×10^4 kg. A large answer is not reasonable, given that you are converting from small mass units (mg) to relatively large mass units (kg). Both of these errors can be caught simply by taking the time to make sure that your answer is reasonable. After completing any calculations, be sure to check for reasonableness in your answers.



Oxygen molecular model

Exercise 1.8 The oxygen molecule (the smallest particle of oxygen gas) consists of two oxygen atoms a distance of 121 pm apart. How many millimeters is this distance?

■ See Problems 1.81, 1.82, 1.83, and 1.84.

Example 1.7 Converting Units: Metric Volume to Metric Volume

The world's oceans contain approximately 1.35×10^9 km^3 of water. What is this volume in liters?

Problem Strategy This problem requires several conversions. A relationship that will prove helpful here is the definition of the liter: 1 dm^3 = 1 L. This provides a clue that if we can convert cubic kilometers (km^3) to cubic decimeters (dm^3), we can use this relationship to convert to liters. When solving problems that involve cubic units such as this one, it is often helpful to think of the conversion factors that you would use if the units were not cubed. In this case, convert from kilometers to meters and then from meters to decimeters. Next, think about how these conversion factors would look if they were cubed. For example, 10^3 m = 1 km, so $(10^3\ \text{m})^3 = (1\ \text{km})^3$ is the cubic relationship. Likewise, 10^{-1} m = 1 dm, so $(10^{-1}\ \text{m})^3 = (1\ \text{dm})^3$ is the cubic relationship. Now these cubic conversion factors can be combined into three steps to solve the problem.

Solution

$$1.35 \times 10^9\ \cancel{\text{km}^3} \times \underbrace{\left(\frac{10^3\ \cancel{\text{m}}}{1\ \cancel{\text{km}}}\right)^3}_{\substack{\text{converts}\\ \text{km}^3 \text{ to m}^3}} \times \underbrace{\left(\frac{1\ \text{dm}}{10^{-1}\ \cancel{\text{m}}}\right)^3}_{\substack{\text{converts}\\ \text{m}^3 \text{ to dm}^3}} = 1.35 \times 10^{21}\ \text{dm}^3$$

Because a cubic decimeter is equal to a liter, the volume of the oceans is

$$\mathbf{1.35 \times 10^{21}\ L}$$

Answer Check One of the most common errors made with this type of problem is a failure to cube both the numerator and the denominator in the conversion factors. For example, $(10^3\ \text{m}/1\ \text{km})^3 = 10^9\ \text{m}^3/1\ \text{km}^3$.

(continued)

(continued)

Exercise 1.9 A large crystal is constructed by stacking small, identical pieces of crystal, much as you construct a brick wall by stacking bricks. A unit cell is the smallest such piece from which a crystal can be made. The unit cell of a crystal of gold metal has a volume of 67.6 Å^3. What is this volume in cubic decimeters?

■ See Problems 1.85 and 1.86.

The conversion-factor method can be used to convert any unit to another unit, provided a conversion equation exists between the two units. Relationships between certain U.S. units and metric units are given in Table 1.4. You can use these to convert between U.S. and metric units. Suppose you wish to convert 0.547 lb to grams. From Table 1.4, note that 1 lb = 0.4536 kg, or 1 lb = 453.6 g, so the conversion factor from pounds to grams is 453.6 g/1 lb. Therefore,

$$0.547\,\cancel{\text{lb}} \times \frac{453.6\text{ g}}{1\,\cancel{\text{lb}}} = 248\text{ g}$$

The next example illustrates a conversion requiring several steps.

TABLE 1.4 Relationships of Some U.S. and Metric Units

Length	Mass	Volume
1 in. = 2.54 cm (exact)	1 lb = 0.4536 kg	1 qt = 0.9464 L
1 yd = 0.9144 m (exact)	1 lb = 16 oz (exact)	4 qt = 1 gal (exact)
1 mi = 1.609 km	1 oz = 28.35 g	
1 mi = 5280 ft (exact)		

Example 1.8 Converting Units: Any Unit to Another Unit

How many centimeters are there in 6.51 miles? Use the exact definitions 1 mi = 5280 ft, 1 ft = 12 in., and 1 in. = 2.54 cm.

Problem Strategy This problem involves converting from U.S. to metric units: specifically, from miles to centimeters. One path to a solution is using three conversions: miles to feet, feet to inches, and inches to centimeters. Table 1.4 has the information you need to develop the conversion factors.

Solution From the definitions, you obtain the following conversion factors:

$$1 = \frac{5280\text{ ft}}{1\text{ mi}} \qquad 1 = \frac{12\text{ in.}}{1\text{ ft}} \qquad 1 = \frac{2.54\text{ cm}}{1\text{ in.}}$$

Then,

$$6.51\,\cancel{\text{mi}} \times \underbrace{\frac{5280\,\cancel{\text{ft}}}{1\,\cancel{\text{mi}}}}_{\text{converts mi to ft}} \times \underbrace{\frac{12\,\cancel{\text{in.}}}{1\,\cancel{\text{ft}}}}_{\text{converts ft to in.}} \times \underbrace{\frac{2.54\text{ cm}}{1\,\cancel{\text{in.}}}}_{\text{converts in. to cm}} = \mathbf{1.05 \times 10^6\ cm}$$

All of the conversion factors are exact, so the number of significant figures in the result is determined by the number of significant figures in 6.51 mi.

(continued)

(continued)

Answer Check Because miles are a relatively large unit of length and centimeters are a relatively small unit of length, one would expect that converting from miles to centimeters would yield a very large number. This is the case here, so the answer is reasonable.

Exercise 1.10 Using the definitions 1 in. = 2.54 cm and 1 yd = 36 in. (both exact), obtain the conversion factor for yards to meters. How many meters are there in 3.54 yd?

■ See Problems 1.87, 1.88, 1.89, and 1.90.

A Checklist for Review

Important Terms

Note: The number in parentheses denotes the section in which the term is defined.

experiment (1.2)
law (1.2)
hypothesis (1.2)
theory (1.2)
mass (1.3)
matter (1.3)
law of conservation of mass (1.3)
solid (1.4)
liquid (1.4)
gas (1.4)
states of matter (1.4)
physical change (1.4)
chemical change (chemical reaction) (1.4)
physical property (1.4)
chemical property (1.4)
substance (1.4)
element (1.4)
compound (1.4)
law of definite proportions (law of constant composition) (1.4)
mixture (1.4)
heterogeneous mixture (1.4)
homogeneous mixture (solution) (1.4)
phase (1.4)
unit (1.5)
precision (1.5)
accuracy (1.5)
significant figures (1.5)
number of significant figures (1.5)
scientific notation (1.5)
exact number (1.5)
rounding (1.5)
International System of units (SI) (1.6)
SI base units (1.6)
SI prefix (1.6)
meter (m) (1.6)
angstrom (Å) (1.6)
kilogram (kg) (1.6)
second (s) (1.6)
Celsius scale (1.6)
kelvin (K) (1.6)
SI derived unit (1.7)
liter (L) (1.7)
density (1.7)
dimensional analysis (factor-label method) (1.8)
conversion factor (1.8)

Key Equations

$$T_K = \left(t_C \times \frac{1\ \text{K}}{1^\circ\text{C}}\right) + 273.15\ \text{K}$$

$$t_C = \frac{5^\circ\text{C}}{9^\circ\text{F}} \times (t_F - 32^\circ\text{F})$$

$$d = m/V$$

Summary of Facts and Concepts

Chemistry is an experimental science in that the facts of chemistry are obtained by *experiment.* These facts are systematized and explained by *theory,* and theory suggests more experiments. The *scientific method* involves this interplay, in which the body of accepted knowledge grows as it is tested by experiment.

Chemistry emerged as a *quantitative science* with the work of the eighteenth-century French chemist Antoine Lavoisier. He made use of the idea that the mass, or quantity of matter, remains constant during a chemical reaction *(law of conservation of mass).*

Matter may be classified by its physical state as a *gas, liquid,* or *solid.* Matter may also be classified by its chemical constitution as an *element, compound,* or *mixture.* Materials are either substances or mixtures of substances. Substances are either elements or compounds, which are composed of two or more elements. Mixtures can be separated into substances by physical processes, but compounds can be separated into elements only by chemical reactions.

A quantitative science requires the making of measurements. Any measurement has limited *precision,* which you convey by writing the measured number to a certain number of *significant figures*. There are many different systems of measurement, but scientists generally use the metric system. The International System (SI) uses a particular selection of metric units. It employs seven *base units* combined with *prefixes* to obtain units of various size. Units for other quantities are derived from these. To obtain a *derived unit* in SI for a quantity such as the volume or density, you merely substitute base units into a defining equation for the quantity.

Dimensional analysis is a technique of calculating with physical quantities in which units are included and treated in the same way as numbers. You can thus obtain the *conversion factor* needed to express a quantity in new units.

Media Summary

Visit the **student website at college.hmco.com/pic/ebbing9e** to help prepare for class, study for quizzes and exams, understand core concepts, and visualize molecular-level interactions. The following media activities are available for this chapter:

Prepare for Class

- **Video Lessons** *Mini lectures from chemistry experts*

  An Introduction to Chemistry
  The Scientific Method
  States of Matter
  CIA Demonstration: Distillation
  Properties of Matter
  CIA Demonstration: Chromatography
  Precision and Accuracy
  CIA Demonstration: Precision and Accuracy with Glassware
  Significant Figures
  Scientific (Exponential) Notation
  The Measurement of Matter
  CIA Demonstration: Differences in Density Due to Temperature
  Dimensional Analysis

Improve Your Grade

- **Visualizations** *Molecular-level animations and lab demonstration videos*

  Structure of a Solid
  Structure of a Liquid
  Structure of a Gas
  Comparison of a Compound and a Mixture
  Comparison of a Solution and a Mixture
  Homogeneous Mixtures: Air and Brass

- **Tutorials** *Animated examples and interactive activities*

  Unit Conversions

- **Flashcards** *Key terms and definitions*

  Online Flashcards

- **Self-Assessment Questions** *Additional questions with full worked-out solutions*

  6 Self-Assessment Questions

ACE the Test

Multiple-choice quizzes
3 ACE Practice Tests

Access these resources using the passkey available free with new texts or for purchase separately.

Learning Objectives

1.1 Modern Chemistry: A Brief Glimpse

- Provide examples of the contributions of chemistry to humanity.

1.2 Experiment and Explanation

- Describe how chemistry is an experimental science.
- Understand how the scientific method is an approach to performing science.

1.3 Law of Conservation of Mass

- Explain the law of conservation of mass.
- Apply the law of the conservation of mass. Example 1.1

1.4 Matter: Physical State and Chemical Constitution

- Compare and contrast the three common states of matter: solid, liquid, and gas.
- Describe the classifications of matter: elements, compounds, and mixtures (heterogeneous and homogeneous).
- Understand the difference between chemical changes (chemical reactions) and physical changes.
- Distinguish between chemical properties and physical properties.

1.5 Measurement and Significant Figures

- Define and use the terms *precision* and *accuracy* when describing measured quantities.

- Learn the rules for determining significant figures in reported measurements.
- Know how to represent numbers using scientific notation.
- Apply the rules of significant figures to reporting calculated values.
- Be able to recognize exact numbers.
- Know when and how to apply the rules for rounding.
- Use significant figures in calculations. Example 1.2

1.6 SI Units

- Become familiar with the SI (metric) system of units, including the SI prefixes.
- Convert from one temperature scale to another. Example 1.3

1.7 Derived Units

- Define and provide examples of derived units.
- Calculate the density of a substance. Example 1.4
- Use density to relate mass and volume. Example 1.5

1.8 Units and Dimensional Analysis (Factor-Label Method)

- Apply dimensional analysis to solving numerical problems.
- Convert from one metric unit to another metric unit. Example 1.6
- Convert from one metric volume to another metric volume. Example 1.7
- Convert from any unit to another unit. Example 1.8

Self-Assessment and Review Questions

Key: These questions test your understanding of the ideas you worked with in the chapter. Each section ends with four multiple-choice questions that you can use to help assess whether you have understood the main concepts; these are answered in the Answers to Self-Assessment Questions in the back of the book. There are also six additional multiple-choice questions available at the student website for further practice.

1.1 Discuss some ways in which chemistry has changed technology. Give one or more examples of how chemistry has affected another science.
1.2 Define the terms *experiment* and *theory*. How are theory and experiment related? What is a hypothesis?
1.3 Illustrate the steps in the scientific method using Rosenberg's discovery of the anticancer activity of cisplatin.
1.4 Define the terms *matter* and *mass*. What is the difference between mass and weight?
1.5 State the law of conservation of mass. Describe how you might demonstrate this law.
1.6 A chemical reaction is often accompanied by definite changes in appearance. For example, heat and light may be emitted, and there may be a change of color of the substances. Figure 1.9 shows the reactions of the metal mercury with oxygen in air. Describe the changes that occur.
1.7 Characterize gases, liquids, and solids in terms of compressibility and fluidity.
1.8 Choose a substance and give several of its physical properties and several of its chemical properties.
1.9 Give examples of an element, a compound, a heterogeneous mixture, and a homogeneous mixture.
1.10 What phases or states of matter are present in a glass of bubbling carbonated beverage that contains ice cubes?
1.11 What distinguishes an element from a compound? Can a compound also be an element?
1.12 What is meant by the precision of a measurement? How is it indicated?
1.13 Two rules are used to decide how to round the result of a calculation to the correct number of significant figures. Use a calculation to illustrate each rule. Explain how you obtained the number of significant figures in the answers.
1.14 Distinguish between a measured number and an exact number. Give examples of each.
1.15 How does the International System (SI) obtain units of different size from a given unit? How does the International System obtain units for all possible physical quantities from only seven base units?
1.16 What is an absolute temperature scale? How are degrees Celsius related to kelvins?
1.17 Define *density*. Describe some uses of density.
1.18 Why should units be carried along with numbers in a calculation?
1.19 When the quantity 12.9 g is added to 2×10^{-02} g, how many significant figures should be reported in the answer?
a. one **b.** two **c.** three **d.** four **e.** five
1.20 You perform an experiment in the lab and determine that there are 36.3 inches in a meter. Using this experimental value, how many millimeters are there in 1.34 feet?
a. 4.43×10^{2} mm
b. 4.05×10^{2} mm
c. 44.3 mm
d. 4.43×10^{5} mm
e. 4.05×10^{8} mm
1.21 A 75.0-g sample of a pure liquid, liquid A, with a density of 3.00 g/mL is mixed with a 50.0-mL sample of a pure liquid, liquid B, with a density of 2.00 g/mL. What is the total volume of the mixture? (Assume there is no reaction upon the mixing of A and B.)
a. 275 mL
b. 175 mL
c. 125 mL
d. 100. mL
e. 75 mL
1.22 Which of the following represents the *smallest* mass?
a. 23 cg
b. 2.3×10^{3} μg
c. 0.23 mg
d. 0.23 g
e. 2.3×10^{-2} kg

Concept Explorations

Key: Concept explorations are comprehensive problems that provide a framework that will enable you to explore and learn many of the critical concepts and ideas in each chapter. If you master the concepts associated with these explorations, you will have a better understanding of many important chemistry ideas and will be more successful in solving all types of chemistry problems. These problems are well suited for group work and for use as in-class activities.

1.23 Physical and Chemical Changes

Say you are presented with two beakers, beaker A and beaker B, each containing a white, powdery compound.

a. From your initial observations, you suspect that the two beakers contain the same compound. Describe, in general terms, some experiments in a laboratory that you could do to help prove or disprove that the beakers contain the same compound.

b. Would it be easier to prove that the compounds are the same or to prove that they different? Explain your reasoning.

c. Which of the experiments that you listed above are the most convincing in determining whether the compounds are the same? Justify your answer.

d. A friend states that the best experiment for determining whether the compounds are the same is to see if they both dissolve in water. He proceeds to take 10.0 g of each compound and places them in separate beakers, each containing 100 mL of water. Both compounds completely dissolve. He then states, "Since the same amount of both substances dissolved in the same volume of water, they must both have the same chemical composition." Is he justified in making this claim? Why or why not?

1.24 Significant Figures

Part 1

a. Consider three masses that you wish to add together: 3 g, 1.4 g, and 3.3 g. These numbers represent measured values. Add the numbers together and report your answer to the correct number of significant figures.

b. Now perform the addition in a stepwise fashion in the following manner. Add 3 g and 1.4 g, reporting this sum to the correct number of significant figures. Next, take the number from the first step and add it to 3.3 g, reporting this sum to the correct number of significant figures.

c. Compare your answers from performing the addition in the two distinct ways presented in parts a and b. Does one of the answers represent a "better" way of reporting the results of the addition? If your answer is yes, explain why your choice is better.

d. A student performs the calculation $(5.0 \times 5.143\ \text{g}) + 2.80\ \text{g}$ and, being mindful of significant figures, reports an answer of 29 g. Is this the correct answer? If not, what might this student have done incorrectly?

e. Another student performs the calculation $(5 \times 5.143\ \text{g}) + 2.80$ and reports an answer of 29 g. Is this the correct answer? If not, what might this student have done incorrectly?

f. Yet another student performs the calculation $(5.00 \times 5.143\ \text{g}) + 2.80$ and reports an answer of 28.5 g. Is this the correct answer? If not, what did this student probably do incorrectly?

g. Referring to the calculations above, outline a procedure or rule(s) that will always enable you to report answers using the correct number of significant figures.

Part 2

a. A student wants to determine the volume of 27.2 g of a substance. He looks up the density of the material in a reference book, where it is reported to be 2.4451 g/cm^3. He performs the calculation in the following manner:

$$27.2\ \text{g} \times 1.0\ \text{cm}^3/2.4\ \text{g} = 11.3\ \text{cm}^3$$

Is the calculated answer correct? If not, explain why it is not correct.

b. Another student performs the calculation in the following manner:

$$27.2\ \text{g} \times 1.00\ \text{cm}^3/2.45\ \text{g} = 11.1\ \text{cm}^3$$

Is this a "better" answer than that of the first student? Is this the "best" answer, or could it be "improved"? Explain.

c. Say that you have ten ball bearings, each having a mass of 1.234 g and a density of 3.1569 g/cm^3. Calculate the volume of these ten ball bearings. In performing the calculation, present your work as unit conversions, and report your answer to the correct number of significant figures.

d. Explain how the answer that you calculated in part c is the "best" answer to the problem?

Conceptual Problems

Key: These problems are designed to check your understanding of the concepts associated with some of the main topics presented in the chapter. A strong conceptual understanding of chemistry is the foundation for both applying chemical knowledge and solving chemical problems. These problems vary in level of difficulty and often can be used as a basis for group discussion.

1.25

a. Sodium metal is partially melted. What are the two phases present?

b. A sample of sand is composed of granules of quartz (silicon dioxide) and seashells (calcium carbonate). The sand is mixed with water. What phases are present?

1.26 A material is believed to be a compound. Suppose you have several samples of this material obtained from various places around the world. Comment on what you would expect to find upon observing the melting point and color for each sample. What would you expect to find upon determining the elemental composition for each sample?

1.27 You need a thermometer that is accurate to ±5°C to conduct some experiments in the temperature range of 0°C to 100°C. You find one in your lab drawer that has lost its markings.

a. What experiments could you do to make sure your thermometer is suitable for your experiments?

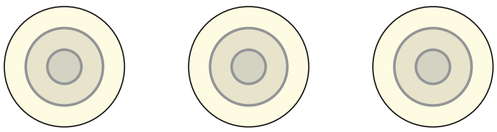
b. Assuming that the thermometer works, what procedure could you follow to put a scale on your thermometer that has the desired accuracy?

1.28 Imagine that you get the chance to shoot five arrows at each of the targets depicted below. On each of the targets, indicate the pattern that the five arrows make when

a. You have poor accuracy and good precision.

b. You have poor accuracy and poor precision.

c. You have good accuracy and good precision.



1.29 Say you live in a climate where the temperature ranges from −100°F to 20°F and you want to define a new temperature scale, YS (YS is the "Your Scale" temperature scale), which defines this range as 0.0°YS to 100.0°YS.

a. Come up with an equation that would allow you to convert between °F and °YS.

b. Using your equation, what would be the temperature in °F if it were 66°YS?

1.30 You are presented with a piece of metal in a jar. It is your job to determine what the metal is. What are some physical properties that you could measure in order to determine the type of metal? You suspect that the metal might be sodium; what are some chemical properties that you could investigate? (See Section 1.4 for some ideas.)

1.31 You have two identical boxes with interior dimensions of 8.0 cm × 8.0 cm × 8.0 cm. You completely fill one of the boxes with wooden spheres that are 1.6 cm in diameter. The other box gets filled with wooden cubes that are 1.6 cm on each edge. After putting the lid on the filled boxes, you then measure the density of each. Which one is more dense?

1.32 Consider the following compounds and their densities.

Substance	*Density (g/mL)*	*Substance*	*Density (g/mL)*
Isopropyl alcohol	0.785	Toluene	0.866
n-Butyl alcohol	0.810	Ethylene glycol	1.114

You create a column of the liquids in a glass cylinder with the most dense material on the bottom layer and the least dense on the top. You do not allow the liquids to mix.

a. First you drop a plastic bead that has a density of 0.24 g/cm^3 into the column. What do you expect to observe?

b. Next you drop a different plastic bead that has a volume of 0.043 mL and a mass of 3.92×10^{-2} g into the column. What would you expect to observe in this case?

c. You drop another bead into the column and observe that it makes it all the way to the bottom of the column. What can you conclude about the density of this bead?

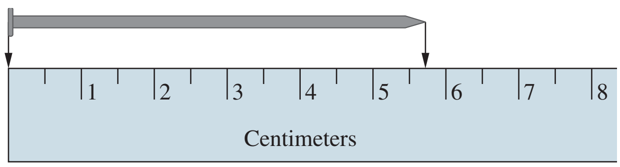
1.33

a. Which of the following items have a mass of about 1 g?

a grain of sand
a paper clip
a nickel
a 5.0-gallon bucket of water
a brick
a car

b. What is the approximate mass (using SI mass units) of each of the items in part a?

1.34 What is the length of the nail reported to the correct number of significant figures?



1.35 For these questions, be sure to apply the rules for significant figures.

a. You are conducting an experiment where you need the volume of a box; you take the length, height, and width measurements and then multiply the values together to find the volume. You report the volume of the box as 0.310 m^3. If two of your measurements were 0.7120 m and 0.52145 m, what was the other measurement?

b. If you were to add the two measurements from the first part of the problem to a third length measurement with the reported result of 1.509 m, what was the value of the third measurement?

1.36 You are teaching a class of second graders some chemistry, and they are confused about an object's mass versus its density. Keeping in mind that most second graders don't understand fractions, how would you explain these two ideas and illustrate how they differ? What examples would you use as a part of your explanation?

Practice Problems

Key: These problems are for practice in applying problem-solving skills. They are divided by topic, and some are keyed to exercises (see the ends of the exercises). The problems are arranged in matched pairs; the odd-numbered problem of each pair is listed first and its answer is given in the back of the book.

Conservation of Mass

1.37 A 15.9-g sample of sodium carbonate is added to a solution of acetic acid weighing 20.0 g. The two substances react, releasing carbon dioxide gas to the atmosphere. After reaction, the contents of the reaction vessel weighs 29.3 g. What is the mass of carbon dioxide given off during the reaction?

1.38 Some iron wire weighing 5.6 g is placed in a beaker and covered with 15.0 g of dilute hydrochloric acid. The acid reacts with the metal and gives off hydrogen gas, which escapes into the surrounding air. After reaction, the contents of the beaker weighs 20.4 g. What is the mass of hydrogen gas produced by the reaction?

1.39 Zinc metal reacts with yellow crystals of sulfur in a fiery reaction to produce a white powder of zinc sulfide. A chemist determines that 65.4 g of zinc reacts with 32.1 g of sulfur. How many grams of zinc sulfide could be produced from 20.0 g of zinc metal?

1.40 Aluminum metal reacts with bromine, a red-brown liquid with a noxious odor. The reaction is vigorous and produces aluminum bromide, a white crystalline substance. A sample of 27.0 g of aluminum yields 266.7 g of aluminum bromide. How many grams of bromine react with 15.0 g of aluminum?

Solids, Liquids, and Gases

1.41 Give the normal state (solid, liquid, or gas) of each of the following.

a. sodium hydrogen carbonate (baking soda)
b. isopropyl alcohol (rubbing alcohol)
c. oxygen
d. copper

1.42 Give the normal state (solid, liquid, or gas) of each of the following.

a. potassium hydrogen tartrate (cream of tartar)
b. lead
c. carbon (diamond)
d. bromine

Chemical and Physical Changes; Properties of Substances

1.43 Which of the following are physical changes and which are chemical changes?

a. melting of sodium chloride ph
b. pulverizing of rock salt ph
c. burning of sulfur che
d. dissolving of salt in water che

1.44 For each of the following, decide whether a physical or a chemical change is involved.

a. dissolving of sugar in water
b. rusting of iron
c. burning of wood
d. evaporation of alcohol

1.45 A sample of mercury(II) oxide was heated to produce mercury metal and oxygen gas. Then the liquid mercury was cooled to −40°C, where it solidified. A glowing wood splint was thrust into the oxygen, and the splint burst into flame. Identify each physical change and each chemical change.

1.46 Solid iodine, contaminated with salt, was heated until the iodine vaporized. The violet vapor of iodine was then cooled to yield the pure solid. Solid iodine and zinc metal powder were mixed and ignited to give a white powder. Identify each physical change and each chemical change.

1.47 The following are properties of substances. Decide whether each is a physical property or a chemical property.

a. Chlorine gas liquefies at −35°C under normal pressure.
b. Hydrogen burns in chlorine gas.
c. Bromine melts at −7.2°C.
d. Lithium is a soft, silvery-colored metal.
e. Iron rusts in an atmosphere of moist air.

1.48 Decide whether each of the following is a physical property or a chemical property of the substance.

a. Salt substitute, potassium chloride, dissolves in water.
b. Seashells, calcium carbonate, fizz when immersed in vinegar.
c. The gas hydrogen sulfide smells like rotten eggs.
d. Fine steel wool (Fe) can be burned in air.
e. Pure water freezes at 0°C.

1.49 Iodine is a solid having somewhat lustrous, blue-black crystals. The crystals vaporize readily to a violet-colored gas. Iodine combines with many metals. For example, aluminum combines with iodine to give aluminum iodide. Identify the physical and the chemical properties of iodine that are cited.

1.50 Mercury(II) oxide is an orange-red solid with a density of 11.1 g/cm^3. It decomposes when heated to give mercury and oxygen. The compound is insoluble in water (does not dissolve in water). Identify the physical and the chemical properties of mercury(II) oxide that are cited.

Elements, Compounds, and Mixtures

1.51 Consider the following separations of materials. State whether a physical process or a chemical reaction is involved in each separation.

a. Sodium chloride is obtained from seawater by evaporation of the water.
b. Mercury is obtained by heating the substance mercury(II) oxide; oxygen is also obtained.
c. Pure water is obtained from ocean water by evaporating the water, then condensing it.
d. Iron is produced from an iron ore that contains the substance iron(III) oxide.
e. Gold is obtained from river sand by panning (allowing the heavy metal to settle in flowing water).

1.52 All of the following processes involve a separation of either a mixture into substances or a compound into elements. For each, decide whether a physical process or a chemical reaction is required.

a. Sodium metal is obtained from the substance sodium chloride.
b. Iron filings are separated from sand by using a magnet.
c. Sugar crystals are separated from a sugar syrup by evaporation of water.
d. Fine crystals of silver chloride are separated from a suspension of the crystals in water.
e. Copper is produced when zinc metal is placed in a solution of copper(II) sulfate, a compound.

1.53 Label each of the following as a substance, a heterogeneous mixture, or a solution.
a. seawater **b.** sulfur
c. fluorine **d.** beach sand

1.54 Indicate whether each of the following materials is a substance, a heterogeneous mixture, or a solution.
a. milk **b.** bromine
c. gasoline **d.** aluminum

1.55 Which of the following are pure substances and which are mixtures? For each, list all of the different phases present.
a. bromine liquid and its vapor
b. paint, containing a liquid solution and a dispersed solid pigment
c. partially molten iron
d. baking powder containing sodium hydrogen carbonate and potassium hydrogen tartrate

1.56 Which of the following are pure substances and which are mixtures? For each, list all of the different phases present.
a. a sugar solution with sugar crystals at the bottom
b. ink containing a liquid solution with fine particles of carbon
c. a sand containing quartz (silicon dioxide) and calcite (calcium carbonate)
d. liquid water and steam at 100°C

Significant Figures

1.57 How many significant figures are there in each of the following measurements?
a. 73.0000 g **b.** 0.0503 kg
c. 6.300 cm **d.** 0.80090 m
e. 5.10×10^{-7} m **f.** 2.010 s

1.58 How many significant figures are there in each of the following measurements?
a. 301.0 kg **b.** 0.05930 g
c. 0.224800 m **d.** 3.100 s
e. 4.380×10^{-8} m **f.** 9.100×10^{4} cm

1.59 The circumference of the earth at the equator is 40,000 km. This value is precise to two significant figures. Write this in scientific notation to express correctly the number of significant figures.

1.60 The astronomical unit equals the mean distance between the earth and the sun. This distance is 150,000,000 km, which is precise to three significant figures. Express this in scientific notation to the correct number of significant figures.

1.61 Do the indicated arithmetic and give the answer to the correct number of significant figures.
a. $\dfrac{8.71 \times 0.0301}{0.031}$
b. $0.71 + 92.2$
c. $934 \times 0.00435 + 107$
d. $(847.89 - 847.73) \times 14673$

1.62 Do the indicated arithmetic and give the answer to the correct number of significant figures.
a. $\dfrac{08.71 \times 0.57}{5.871}$
b. $8.937 - 8.930$
c. $8.937 + 8.930$
d. $0.00015 \times 54.6 + 1.002$

1.63 One sphere has a radius of 5.10 cm; another has a radius of 5.00 cm. What is the difference in volume (in cubic centimeters) between the two spheres? Give the answer to the correct number of significant figures. The volume of a sphere is $(4/3)\pi r^3$, where $\pi = 3.1416$ and r is the radius.

1.64 A solid cylinder of iron of circular cross section with a radius of 1.500 cm has a ruler etched along its length. What is the volume of iron contained between the marks labeled 3.10 cm and 3.50 cm? The volume of a cylinder is $\pi r^2 l$, where $\pi = 3.1416$, r is the radius, and l is the length.

SI Units

1.65 Write the following measurements, without scientific notation, using the appropriate SI prefix.
a. 5.89×10^{-12} s **b.** 0.2010 m
c. 2.560×10^{-9} g **d.** 6.05×10^{3} m

1.66 Write the following measurements, without scientific notation, using the appropriate SI prefix.
a. 4.851×10^{-6} g **b.** 3.16×10^{-2} m
c. 2.591×10^{-9} s **d.** 8.93×10^{-12} g

1.67 Using scientific notation, convert:
a. 6.15 ps to s **b.** 3.781 μm to m
c. 1.546 Å to m **d.** 9.7 mg to g

1.68 Using scientific notation, convert:
a. 6.20 km to m **b.** 1.98 ns to s
c. 2.54 cm to m **d.** 5.23 µg to g

Temperature Conversion

1.69 Convert:
a. 68°F to degrees Celsius
b. −23°F to degrees Celsius
c. 26°C to degrees Fahrenheit
d. −70°C to degrees Fahrenheit

1.70 Convert:
a. 51°F to degrees Celsius
b. −7°F to degrees Celsius
c. −41°C to degrees Fahrenheit
d. 22°C to degrees Fahrenheit

1.71 Salt and ice are stirred together to give a mixture to freeze ice cream. The temperature of the mixture is −21.1°C. What is this temperature in degrees Fahrenheit?

1.72 Liquid nitrogen can be used for the quick freezing of foods. The liquid boils at −196°C. What is this temperature in degrees Fahrenheit?

Density

1.73 A certain sample of the mineral galena (lead sulfide) weighs 12.4 g and has a volume of 1.64 cm^3. What is the density of galena?

1.74 A flask contains 25.0 mL of diethyl ether weighing 17.84 g. What is the density of the ether?

1.75 A liquid with a volume of 8.5 mL has a mass of 6.71 g. The liquid is either octane, ethanol, or benzene, the densities of which are 0.702 g/cm^3, 0.789 g/cm^3, and 0.879 g/cm^3, respectively. What is the identity of the liquid?

1.76 A mineral sample has a mass of 5.95 g and a volume of 7.9 cm^3. The mineral is either sphalerite (density = 4.0 g/cm^3), cassiterite (density = 6.99 g/cm^3), or cinnabar (density = 8.10 g/cm^3). Which is it?

1.77 Platinum has a density of 21.4 g/cm^3. What is the mass of 5.9 cm^3 of this metal?

1.78 What is the mass of a 43.8-mL sample of gasoline, which has a density of 0.70 g/cm^3?

1.79 Ethanol has a density of 0.789 g/cm^3. What volume must be poured into a graduated cylinder to give 19.8 g of alcohol?

1.80 Bromine is a red-brown liquid with a density of 3.10 g/mL. A sample of bromine weighing 88.5 g occupies what volume?

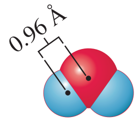
Unit Conversions

1.81 Sodium hydrogen carbonate, known commercially as baking soda, reacts with acidic materials such as vinegar to release carbon dioxide gas. An experiment calls for 0.480 kg of sodium hydrogen carbonate. Express this mass in milligrams.

1.82 The acidic constituent in vinegar is acetic acid. A 10.0-mL sample of a certain vinegar contains 501 mg of acetic acid. What is this mass of acetic acid expressed in micrograms?

1.83 The different colors of light have different wavelengths. The human eye is most sensitive to light whose wavelength is 555 nm (greenish-yellow). What is this wavelength in centimeters?

1.84 Water consists of molecules (groups of atoms). A water molecule has two hydrogen atoms, each connected to an oxygen atom. The distance between any one hydrogen atom and the oxygen atom is 0.96 Å. What is this distance in meters?



1.85 The total amount of fresh water on earth is estimated to be 3.73×10^8 km^3. What is this volume in cubic meters? in liters?

1.86 A submicroscopic particle suspended in a solution has a volume of 1.3 μm^3. What is this volume in liters?

1.87 How many grams are there in 3.58 short tons? Note that 1 g = 0.03527 oz (ounces avoirdupois), 1 lb (pound) = 16 oz, and 1 short ton = 2000 lb. (These relations are exact.)

1.88 The calorie, the Btu (British thermal unit), and the joule are units of energy; 1 calorie = 4.184 joules (exact), and 1 Btu = 252.0 calories. Convert 3.15 Btu to joules.

1.89 The first measurement of sea depth was made in 1840 in the central South Atlantic, where a plummet was lowered 2425 fathoms. What is this depth in meters? Note that 1 fathom = 6 ft, 1 ft = 12 in., and 1 in. = 2.54×10^{-2} m. (These relations are exact.)

1.90 The estimated amount of recoverable oil from the field at Prudhoe Bay in Alaska is 1.3×10^{10} barrels. What is this amount of oil in cubic meters? One barrel = 42 gal (exact), 1 gal = 4 qt (exact), and 1 qt = 9.46×10^{-4} m^3.

1.91 A fish tank is 20.0 in. long, 20.0 in. deep, and 10.0 in. high. What is the maximum volume of water, in liters, that the fish tank can hold?

1.92 The population density of worms in a particular field is 25 worms per cubic meter of soil. How many worms would there be in the top meter of soil in a field that has dimensions of 1.00 km by 2.0 km?

General Problems

Key: These problems provide more practice but are not divided by topic or keyed to exercises. Each section ends with essay questions, each of which is color-coded to refer to the *A Chemist Looks at* (gold) or *Instrumental Methods* (purple) chapter essay on which it is based. Odd-numbered problems and the even-numbered problems that follow are similar; answers to all odd-numbered problems except the essay questions are given in the back of the book.

1.93 Sodium metal reacts vigorously with water. A piece of sodium weighing 19.70 g was added to a beaker containing 126.22 g of water. During reaction, hydrogen gas was produced and bubbled from the solution. The solution, containing sodium hydroxide, weighed 145.06 g. How many grams of hydrogen gas were produced?

1.94 An antacid tablet weighing 0.853 g contained calcium carbonate as the active ingredient, in addition to an inert binder. When an acid solution weighing 56.519 g was added to the tablet, carbon dioxide gas was released, producing a fizz. The resulting solution weighed 57.152 g. How many grams of carbon dioxide were produced?

1.95 When a mixture of aluminum powder and iron(III) oxide is ignited, it produces molten iron and aluminum oxide. In an experiment, 5.40 g of aluminum was mixed with 18.50 g of iron(III) oxide. At the end of the reaction, the mixture contained 11.17 g of iron, 10.20 g of aluminum oxide, and an undetermined amount of unreacted iron(III) oxide. No aluminum was left. What is the mass of the iron(III) oxide?

1.96 When chlorine gas is bubbled into a solution of sodium bromide, the sodium bromide reacts to give bromine, a red-brown liquid, and sodium chloride (ordinary table salt). A solution was made by dissolving 20.6 g of sodium bromide in 100.0 g of water. After passing chlorine through the solution, investigators analyzed the mixture. It contained 16.0 g of bromine and 11.7 g of sodium chloride. How many grams of chlorine reacted?

1.97 A beaker weighed 53.10 g. To the beaker was added 5.348 g of iron pellets and 56.1 g of hydrochloric acid. What was the total mass of the beaker and the mixture (before reaction)? Express the answer to the correct number of significant figures.

1.98 A graduated cylinder weighed 68.1 g. To the cylinder was added 58.2 g of water and 5.279 g of sodium chloride. What was the total mass of the cylinder and the solution? Express the answer to the correct number of significant figures.

1.99 Describe each of the following as a physical or chemical property of each listed chemical substance.

a. baking soda reacts with vinegar
b. ice melts at 0°C
c. graphite is a soft, black solid
d. hydrogen burns in air

1.100 Describe each of the following as a physical or chemical property of each listed chemical substance.

a. chlorine is a green gas
b. zinc reacts with acids
c. magnesium has a density of 1.74 g/cm^3
d. iron can rust

1.101 Analyses of several samples of a material containing only iron and oxygen gave the following results. Could this material be a compound?

	Mass of Sample	*Mass of Iron*	*Mass of Oxygen*
Sample A	1.518 g	1.094 g	0.424 g
Sample B	2.056 g	1.449 g	0.607 g
Sample C	1.873 g	1.335 g	0.538 g

1.102 A red-orange solid contains only mercury and oxygen. Analyses of three different samples gave the following results. Are the data consistent with the hypothesis that the material is a compound?

	Mass of Sample	*Mass of Mercury*	*Mass of Oxygen*
Sample A	1.0410 g	0.9641g	0.0769 g
Sample B	1.5434 g	1.4293 g	0.1141 g
Sample C	1.2183 g	1.1283 g	0.0900 g

1.103 A cubic box measures 39.3 cm on an edge. What is the volume of the box in cubic centimeters? Express the answer to the correct number of significant figures.

1.104 A cylinder with circular cross section has a radius of 2.56 cm and a height of 56.32 cm. What is the volume of the cylinder? Express the answer to the correct number of significant figures.

1.105 An aquarium has a rectangular cross section that is 47.8 in. by 12.5 in.; it is 19.5 in. high. How many U.S. gallons does the aquarium contain? One U.S. gallon equals exactly 231 in^3.

1.106 A spherical tank has a radius of 175.0 in. Calculate the volume of the tank in cubic inches; then convert this to Imperial gallons. The volume of a sphere is (4/3) πr^3, where r is the radius. One Imperial gallon equals 277.4 in^3.

1.107 Obtain the difference in volume between two spheres, one of radius 5.61 cm, the other of radius 5.85 cm. The volume V of a sphere is (4/3) πr^3, where r is the radius. Give the result to the correct number of significant figures.

1.108 What is the difference in surface area between two circles, one of radius 7.98 cm, the other of radius 8.50 cm? The surface area of a circle of radius r is πr^2. Obtain the result to the correct number of significant figures.

1.109 Perform the following arithmetic setups and express the answers to the correct number of significant figures.

a. $\dfrac{56.1 - 51.1}{6.58}$ **b.** $\dfrac{56.1 + 51.1}{6.58}$
c. $(9.1 + 8.6) \times 26.91$ **d.** $0.0065 \times 3.21 + 0.0911$

1.110 Perform the following arithmetic setups and report the answers to the correct number of significant figures.

a. $\dfrac{9.345 - 9.005}{9.811}$ **b.** $\dfrac{9.345 + 9.005}{9.811}$
c. $(7.50 + 7.53) \times 3.71$ **d.** $0.71 \times 0.36 + 17.36$

1.111 For each of the following, write the measurement in terms of an appropriate prefix and base unit.

a. The mass of calcium per milliliter in a sample of blood serum is 0.0912 g.
b. The radius of an oxygen atom is about 0.000000000066 m.
c. A particular red blood cell measures 0.0000071 m.
d. The wavelength of a certain ultraviolet radiation is 0.000000056 m.

1.112 For each of the following, write the measurement in terms of an appropriate prefix and base unit.

a. The mass of magnesium per milliliter in a sample of blood serum is 0.0186 g.
b. The radius of a carbon atom is about 0.000000000077 m.
c. The hemoglobin molecule, a component of red blood cells, is 0.0000000065 m in diameter.
d. The wavelength of a certain infrared radiation is 0.00000085 m.

1.113 Write each of the following in terms of the SI base unit (that is, express the prefix as the power of 10).

a. 1.07 ps **b.** 5.8 μm
c. 319 nm **d.** 15.3 ms

1.114 Write each of the following in terms of the SI base unit (that is, express the prefix as the power of 10).

a. 6.6 mK **b.** 275 pm
c. 22.1 ms **d.** 45 μm

1.115 Tungsten metal, which is used in lightbulb filaments, has the highest melting point of any metal (3410°C). What is this melting point in degrees Fahrenheit?

1.116 Titanium metal is used in aerospace alloys to add strength and corrosion resistance. Titanium melts at 1677°C. What is this temperature in degrees Fahrenheit?

1.117 Calcium carbonate, a white powder used in toothpastes, antacids, and other preparations, decomposes when heated to about 825°C. What is this temperature in degrees Fahrenheit?

1.118 Sodium hydrogen carbonate (baking soda) starts to decompose to sodium carbonate (soda ash) at about 50°C. What is this temperature in degrees Fahrenheit?

1.119 Gallium metal can be melted by the heat of one's hand. Its melting point is 29.8°C. What is this temperature in kelvins? in degrees Fahrenheit?

1.120 Mercury metal is liquid at normal temperatures but freezes at −38.9°C. What is this temperature in kelvins? in degrees Fahrenheit?

1.121 Zinc metal can be purified by distillation (transforming the liquid metal to vapor, then condensing the vapor back to liquid). The metal boils at normal atmospheric pressure at 1666°F. What is this temperature in degrees Celsius? in kelvins?

1.122 Iodine is a bluish-black solid. It forms a violet-colored vapor when heated. The solid melts at 236°F. What is this temperature in degrees Celsius? in kelvins?

1.123 The density of magnesium metal (used in fireworks) is 1.74 g/cm^3. Express this density in SI units (kg/m^3).

1.124 Vanadium metal is added to steel to impart strength. The density of vanadium is 5.96 g/cm^3. Express this in SI units (kg/m^3).

1.125 The density of quartz mineral was determined by adding a weighed piece to a graduated cylinder containing 51.2 mL water. After the quartz was submerged, the water level was 65.7 mL. The quartz piece weighed 38.4 g. What was the density of the quartz?

1.126 Hematite (iron ore) weighing 70.7 g was placed in a flask whose volume was 53.2 mL. The flask with hematite was then carefully filled with water and weighed. The hematite and water weighed 109.3 g. The density of the water was 0.997 g/cm^3. What was the density of the hematite?

1.127 Some bottles of colorless liquids were being labeled when the technicians accidentally mixed them up and lost track of their contents. A 15.0-mL sample withdrawn from one bottle weighed 22.3 g. The technicians knew that the liquid was either acetone, benzene, chloroform, or carbon tetrachloride (which have densities of 0.792 g/cm^3, 0.899 g/cm^3, 1.489 g/cm^3, and 1.595 g/cm^3, respectively). What was the identity of the liquid?

1.128 A solid will float on any liquid that is more dense than it is. The volume of a piece of calcite weighing 35.6 g is 12.9 cm^3. On which of the following liquids will the calcite float: carbon tetrachloride (density = 1.60 g/cm^3), methylene bromide (density = 2.50 g/cm^3), tetrabromoethane (density = 2.96 g/cm^3), methylene iodide (density = 3.33 g/cm^3)?

1.129 Platinum metal is used in jewelry; it is also used in automobile catalytic converters. What is the mass of a cube of platinum 4.40 cm on an edge? The density of platinum is 21.4 g/cm^3.

1.130 Ultrapure silicon is used to make solid-state devices, such as computer chips. What is the mass of a circular cylinder of silicon that is 12.40 cm long and has a radius of 4.00 cm? The density of silicon is 2.33 g/cm^3.

1.131 Vinegar contains acetic acid (about 5% by mass). Pure acetic acid has a strong vinegar smell but is corrosive to the skin. What volume of pure acetic acid has a mass of 35.00 g? The density of acetic acid is 1.053 g/mL.

1.132 Ethyl acetate has a characteristic fruity odor and is used as a solvent in paint lacquers and perfumes. An experiment requires 0.035 kg of ethyl acetate. What volume is this (in liters)? The density of ethyl acetate is 0.902 g/mL.

1.133 Convert:
- **a.** 8.45 kg to micrograms
- **b.** 318 μs to milliseconds
- **c.** 93 km to nanometers
- **d.** 37.1 mm to centimeters

1.134 Convert:
- **a.** 125 Å to micrometers
- **b.** 32.4 kg to milligrams
- **c.** 16.8 cm to millimeters
- **d.** 2.2 ns to microseconds

1.135 Convert:
- **a.** 5.91 kg of chrome yellow to milligrams
- **b.** 753 mg of vitamin A to micrograms
- **c.** 90.1 MHz (megahertz), the wavelength of an FM signal, to kilohertz
- **d.** 498 mJ (the joule, J, is a unit of energy) to kilojoules

1.136 Convert:
- **a.** 7.19 μg of cyanocobalamin (vitamin B_{12}) to milligrams
- **b.** 104 pm, the radius of a sulfur atom, to angstroms
- **c.** 0.010 mm, the diameter of a typical blood capillary, to centimeters
- **d.** 0.0605 kPa (the pascal, Pa, is a unit of pressure) to centipascals

1.137 The largest of the Great Lakes is Lake Superior, which has a volume of 12,230 km^3. What is this volume in liters?

1.138 The average flow of the Niagara River is 0.501 km^3 of water per day. What is this volume in liters?

1.139 A room measures 10.0 ft × 11.0 ft and is 9.0 ft high. What is its volume in liters?

1.140 A cylindrical settling tank is 5.0 ft deep and has a radius of 15.0 ft. What is the volume of the tank in liters?

1.141 The masses of diamonds and gems are measured in carats. A carat is defined as 200 mg. If a jeweler has 275 carats of diamonds, how many grams does she have?

1.142 One year of world production of gold was 49.6×10^6 troy ounces. One troy ounce equals 31.10 g. What was the world production of gold in metric tons (10^6 g) for that year?

1.143 What are some characteristics of the adhesive used for Post-it Notes?

1.144 All good experiments start with a scientific question. What was the scientific question that Art Fry was trying to answer when he embarked on finding the adhesive for the Post-it Note?

1.145 What do the various chromatographic separation techniques have in common?

1.146 Describe how gas chromatography works.

Strategy Problems

Key: As noted earlier, all of the practice and general problems are matched pairs. This section is a selection of problems that are not in matched-pair format. These challenging problems require that you employ many of the concepts and strategies that were developed in the chapter. In some cases, you will have to integrate several concepts and operational skills in order to solve the problem successfully.

1.147 When the quantity 5×10^{-2} mg is subtracted from 4.6 mg, how many significant figures should be reported in the answer?

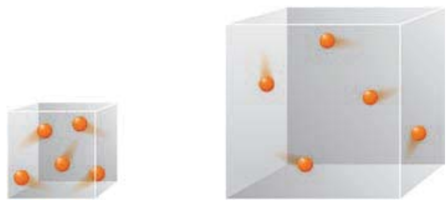
1.148 A 33.0-g sample of an unknown liquid at 20.0°C is heated to120°C. During this heating, the density of the liquid changes from 0.854 g/cm^3 to 0.797 g/cm^3. What volume would this sample occupy at 120°C?

1.149 A 175-g sample of a pure liquid, liquid A, with a density of 3.00 g/mL is mixed with a 50.0-mL sample of a pure liquid, liquid B, with a density of 2.00 g/mL. What is the total volume of the mixture? (Assume there is no reaction upon the mixing of A and B.)

1.150 On a long trip you travel 832 miles in 21 hours. During this trip, you use 29 gallons of gasoline.

a. Calculate your average speed in miles per hour.
b. Calculate your average speed in kilometers per hour.
c. Calculate your gas mileage in kilometers per liter.

1.151 The figures below represent a gas trapped in containers. The orange balls represent individual gas atoms. Container A on the left has a volume that is one-half the volume of container B on the right.



a. How does the mass of gas in one container compare with the mass of gas in the other?
b. Which container has the greater density of gas?
c. If the volume of container A were increased to the same volume as container B, how would the density of the gas in container A change?

1.152 An ice cube measures 3.50 cm on each edge and weighs 39.45 g.

a. Calculate the density of ice.
b. Calculate the mass of 400.4 mL of water in an ice cube.

1.153 The total length of all the DNA molecules contained in a human body is 1×10^{10} miles. The population of the United States is about 300 million. What is the total length of all the DNA of the U.S. population in lightyears? (A lightyear is the distance that light travels in a year and is 9.46×10^{15} m.)

1.154 Prospectors are considering searching for gold on a plot of land that contains 1.16 g of gold per bucket of soil. If the volume of the bucket is 4.08 L, how many grams of gold are there likely to be in 2.38×10^3 cubic feet of soil?

1.155 A solution is prepared by dissolving table salt, sodium chloride, in water at room temperature.

a. Assuming there is no significant change in the volume of water during the preparation of the solution, how would the density of the solution compare to that of pure water?
b. If you were to boil the solution for several minutes and then allow it to cool to room temperature, how would the density of the solution compare to the density in part a?
c. If you took the solution prepared in part a and added more water, how would this affect the density of the solution?

1.156 Water and saline (salt) solution have in common that they are both homogeneous. How do these materials differ? Be specific and use chemical terms to describe the two systems.

Cumulative-Skills Problems

Key: These problems require two or more operational skills you learned in this chapter. In later chapters, the problems under this heading will combine skills introduced in previous chapters with those given in the current one.

1.157 When 10.0 g of marble chips (calcium carbonate) is treated with 50.0 mL of hydrochloric acid (density 1.096 g/mL), the marble dissolves, giving a solution and releasing carbon dioxide gas. The solution weighs 60.4 g. How many liters of carbon dioxide gas are released? The density of the gas is 1.798 g/L.

1.158 Zinc ore (zinc sulfide) is treated with sulfuric acid, leaving a solution with some undissolved bits of material and releasing hydrogen sulfide gas. If 10.8 g of zinc ore is treated with 50.0 mL of sulfuric acid (density 1.153 g/mL), 65.1 g of solution and undissolved material remains. In addition, hydrogen sulfide (density 1.393 g/L) is evolved. What is the volume (in liters) of this gas?

1.159 A steel sphere has a radius of 1.58 in. If this steel has a density of 7.88 g/cm^3, what is the mass of the sphere in grams?

1.160 A weather balloon filled with helium has a diameter of 3.00 ft. What is the mass in grams of the helium in the balloon at 21°C and normal pressure? The density of helium under these conditions is 0.166 g/L.

1.161 The land area of Greenland is 840,000 mi^2, with only 132,000 mi^2 free of perpetual ice. The average thickness of this ice is 5000 ft. Estimate the mass of the ice (assume two significant figures). The density of ice is 0.917 g/cm^3.

1.162 Antarctica, almost completely covered in ice, has an area of 5,500,000 mi^2 with an average height of 7500 ft. Without the ice, the height would be only 1500 ft. Estimate the mass of this ice (two significant figures). The density of ice is 0.917 g/cm^3.

1.163 A sample of an ethanol–water solution has a volume of 54.2 cm^3 and a mass of 49.6 g. What is the percentage of ethanol (by mass) in the solution? (Assume that there is no change in volume when the pure compounds are mixed.) The density of ethanol is 0.789 g/cm^3 and that of water is 0.998 g/cm^3. Alcoholic beverages are rated in *proof,* which is a measure of the relative amount of ethanol in the beverage. Pure ethanol is exactly 200 proof; a solution that is 50% ethanol by volume is exactly 100 proof. What is the proof of the given ethanol–water solution?

1.164 You have a piece of gold jewelry weighing 9.35 g. Its volume is 0.654 cm^3. Assume that the metal is an alloy (mixture) of gold and silver, which have densities of 19.3 g/cm^3 and 10.5 g/cm^3, respectively. Also assume that there is no change in volume when the pure metals are mixed. Calculate the percentage of gold (by mass) in the alloy. The relative amount of gold in an alloy is measured in *karats.* Pure gold is 24 karats; an alloy of 50% gold is 12 karats. State the proportion of gold in the jewelry in karats.

1.165 A sample of vermilion-colored mineral was weighed in air, then weighed again while suspended in water. An object is buoyed up by the mass of the fluid displaced by the object. In air, the mineral weighed 18.49 g; in water, it weighed 16.21 g. The densities of air and water are 1.205 g/L and 0.9982 g/cm^3, respectively. What is the density of the mineral?

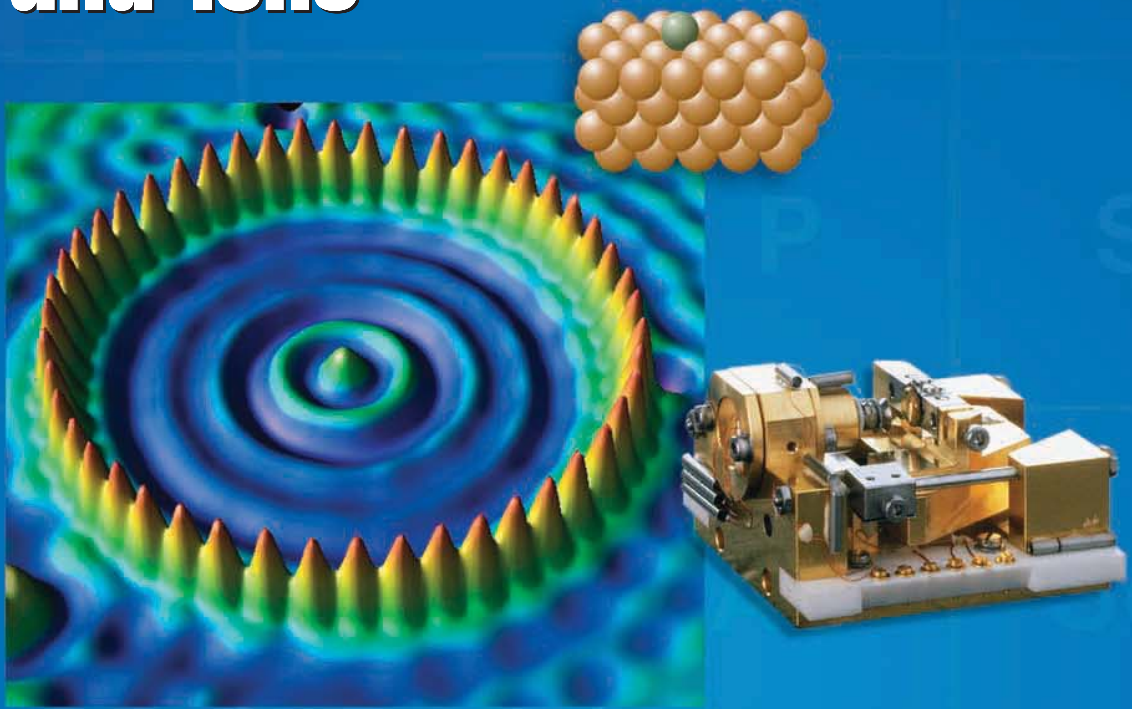
1.166 A sample of a bright blue mineral was weighed in air, then weighed again while suspended in water. An object is buoyed up by the mass of the fluid displaced by the object. In air, the mineral weighed 7.35 g; in water, it weighed 5.40 g. The densities of air and water are 1.205 g/L and 0.9982 g/cm^3, respectively. What is the density of the mineral?

1.167 A student gently drops an object weighing 15.8 g into an open vessel that is full of ethanol, so that a volume of ethanol spills out equal to the volume of the object. The experimenter now finds that the vessel and its contents weigh 10.5 g more than the vessel full of ethanol only. The density of ethanol is 0.789 g/cm^3. What is the density of the object?

1.168 An experimenter places a piece of a solid metal weighing 255 g into a graduated cylinder, which she then fills with liquid mercury. After weighing the cylinder and its contents, she removes the solid metal and fills the cylinder with mercury. She now finds that the cylinder and its contents weigh 101 g less than before. The density of mercury is 13.6 g/cm^3. What is the density of the solid metal?

2

Atoms, Molecules, and Ions



The probe of a scanning tunneling microscope (as sharp as a single atom!) was used to arrange iron atoms in a circle on the surface of a copper metal lattice and collect data to produce this image.

Contents and Concepts

The key concept in chemistry is that all matter is composed of very small particles called atoms. *We look at* atomic theory, *discuss* atomic structure, *and finally describe the* periodic table, *which organizes the elements.*

We explore how atoms combine in various ways to yield the millions of known substances.

Sodium is a soft, silvery metal. This metal cannot be handled with bare fingers, because it reacts with any moisture on the skin, causing a burn. Chlorine is a poisonous, greenish yellow gas with a choking odor. When molten sodium is placed into an atmosphere of chlorine, a dramatic reaction occurs: the sodium bursts into flame and a white, crystalline powder forms (see Figure 2.1). What is particularly interesting about this reaction is that this chemical combination of two toxic substances produces sodium chloride, the substance that we commonly call table salt.

Sodium metal and chlorine gas are particular forms of matter. In burning, they undergo a chemical change—a chemical reaction—in which these forms of matter change to a form of matter with different chemical and physical properties. How do we explain the differences in properties of different forms of matter? And how do we explain chemical reactions such as the burning of sodium metal in chlorine gas? This chapter and the next take an introductory look at these basic questions in chemistry. In later chapters we will develop the concepts introduced here.

■ See page 75 for the Media Summary.

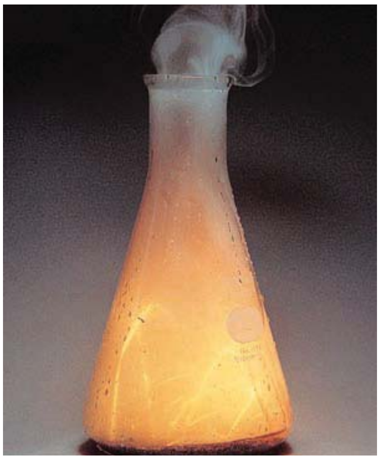



FIGURE 2.1

Reaction of sodium and chlorine
Top: Sodium metal and chlorine gas. *Bottom:* A small piece of molten sodium burning in a flask of chlorine. The product is sodium chloride, common table salt, along with lots of energy.

Atomic Theory and Atomic Structure

As noted in Chapter 1, all matter is composed of atoms. In this first part of Chapter 2, we explore the atomic theory of matter and the structure of atoms. We also examine the *periodic table,* which is an organizational chart designed to highlight similarities among the elements.

2.1 Atomic Theory of Matter

As we noted in Chapter 1, Lavoisier laid the experimental foundation of modern chemistry. But the British chemist John Dalton (1766–1844) provided the basic theory: all matter—whether element, compound, or mixture—is composed of small particles called atoms. The postulates, or basic statements, of Dalton's theory are presented in this section. Note that the terms *element, compound,* and *chemical reaction,* which were defined in Chapter 1 in terms of matter as we normally see it, are redefined here by the postulates of Dalton's theory in terms of atoms.

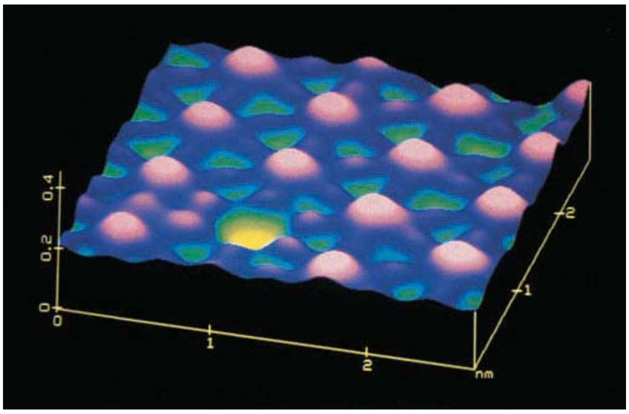


FIGURE 2.2

Image of iodine atoms on a platinum metal surface
This computer-screen image was obtained by a scanning tunneling microscope (discussed in Chapter 7); the color was added to the image by computer. Iodine atoms are the large peaks with pink tops. Note the "vacancy" in the array of iodine atoms. A scale shows the size of the atoms in nanometers.

Postulates of Dalton's Atomic Theory

The main points of Dalton's **atomic theory,** *an explanation of the structure of matter in terms of different combinations of very small particles,* are given by the following postulates:

1. All matter is composed of indivisible atoms. An **atom** is *an extremely small particle of matter that retains its identity during chemical reactions.* (See Figure 2.2.)
2. An **element** is *a type of matter composed of only one kind of atom,* each atom of a given kind having the same properties. Mass is one such property. Thus, the atoms of a given element have a characteristic mass. (We will need to revise this definition of element in Section 2.3 so that it is stated in modern terms.)
3. A **compound** is *a type of matter composed of atoms of two or more elements chemically combined in fixed proportions.* The relative numbers of any two kinds of atoms in a compound occur in simple ratios. Water, for example, a compound of the elements hydrogen and oxygen, consists of hydrogen and oxygen atoms in the ratio of 2 to 1.
4. A **chemical reaction** consists of *the rearrangement of the atoms present in the reacting substances to give new chemical combinations present in the substances formed by the reaction.* Atoms are not created, destroyed, or broken into smaller particles by any chemical reaction.

As you will see in Section 2.4, it is the average mass of an atom that is characteristic of each element on earth.

Today we know that atoms are not truly indivisible; they are themselves made up of particles. Nevertheless, Dalton's postulates are essentially correct.

Atomic Symbols and Models

It is convenient to use symbols for the atoms of the different elements. An **atomic symbol** is *a one- or two-letter notation used to represent an atom corresponding to a particular element.* Typically, the atomic symbol consists of the first letter, capitalized, from the name of the element, sometimes with an additional letter from the name in lowercase. For example, chlorine has the symbol Cl. Other symbols are derived from a name in another language (usually Latin). Sodium is given the symbol Na from its Latin name, *natrium.* Symbols of the elements are listed in the table on the back cover.

The atomic names (temporary) given to newly discovered elements are derived from the atomic number (discussed in Section 2.3) using the numerical roots: nil (0), un (1), bi (2), tri (3), quad (4), pent (5), hex (6), sept (7), oct (8), and enn (9) with the ending -ium. *For example, the element with atomic number 116 is ununhexium; its atomic symbol is Uuh.*

Early in the development of his atomic theory, Dalton built spheres to represent atoms and used combinations of these spheres to represent compounds. Chemists continue to use this idea of representing atoms by three-dimensional models, but because we now know so much more about atoms and molecules than Dalton did, these models have become more refined. Section 2.6 describes these models in more detail.

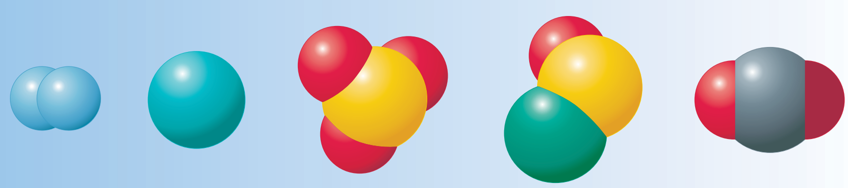
Deductions from Dalton's Atomic Theory

Note how atomic theory explains the difference between an element and a compound. Atomic theory also explains two laws we considered earlier. One of these is the law of conservation of mass, which states that the total mass remains constant during a chemical reaction. By postulate 2, every atom has a definite mass. Because a chemical reaction only rearranges the chemical combinations of atoms (postulate 4), the mass must remain constant. The other law explained by atomic theory is the law of definite proportions (constant composition). Postulate 3 defines a compound as a type of matter containing the atoms of two or more elements in definite proportions. Because the atoms have definite mass, compounds must have the elements in definite proportions by mass.

A good theory should not only explain known facts and laws but also predict new ones. The **law of multiple proportions,** deduced by Dalton from his atomic theory, states that *when two elements form more than one compound, the masses of one element in these compounds for a fixed mass of the other element are in ratios of small whole numbers.* To illustrate this law, consider the following. If we take a fixed mass of carbon, 1.000 gram, and react it with oxygen, we end up with two compounds: one that contains 1.3321 grams of oxygen for each 1.000 gram of carbon, and one that contains 2.6642 grams of oxygen per 1.000 gram of carbon. Note that the ratio of the amounts of oxygen in the two compounds is 2 to 1 (2.6642 ÷ 1.3321), which indicates that there must be twice as much oxygen in the second compound. Applying atomic theory, if we assume that the compound that has 1.3321 grams of oxygen to 1.000 gram of carbon is CO (the combination of one carbon atom and one oxygen atom), then the compound that contains twice as much oxygen per 1.000 gram of carbon must be CO_2 (the combination of one carbon atom and two oxygen atoms). The deduction of the law of multiple proportions from atomic theory was important in convincing chemists of the validity of the theory.

Concept Check 2.1

Like Dalton, chemists continue to model atoms using spheres. Modern models are usually drawn with a computer and use different colors to represent atoms of different elements. Which of the models below represents CO_2?



2.2 The Structure of the Atom

Although Dalton had postulated that atoms were indivisible particles, experiments conducted around the beginning of the last century showed that atoms themselves consist of particles. These experiments showed that an atom consists of two kinds of particles: a **nucleus,** *the atom's central core, which is positively charged and contains most of the atom's mass,* and one or more electrons. An **electron** is *a very light, negatively charged particle that exists in the region around the atom's positively charged nucleus.*

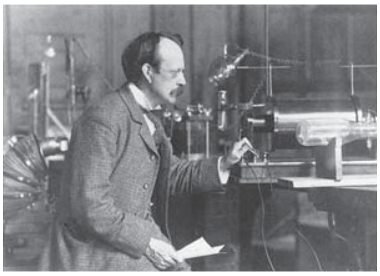


FIGURE 2.3

Joseph John Thomson (1856–1940)
J. J. Thomson's scientific ability was recognized early with his appointment as professor of physics in the Cavendish Laboratory at Cambridge University when he was not quite 28 years old. Soon after this appointment, Thomson began research on the discharge of electricity through gases. This work culminated in 1897 with the discovery of the electron. Thomson was awarded the Nobel Prize in physics in 1906.

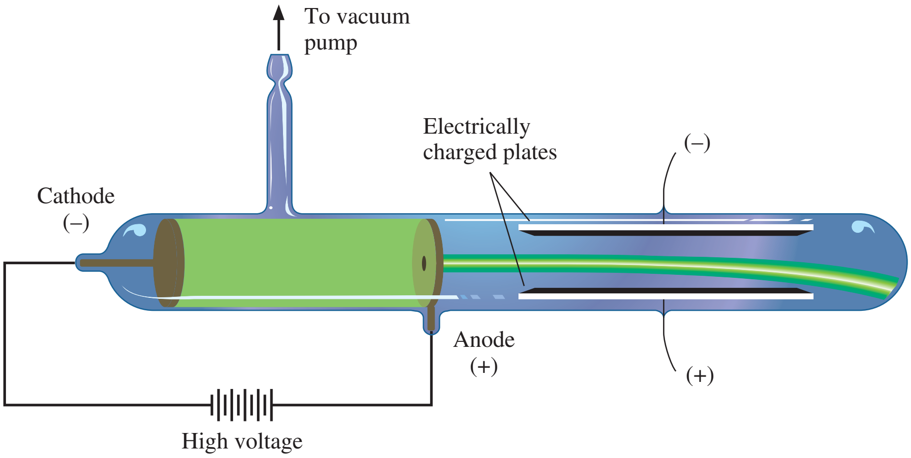


FIGURE 2.4

Formation of cathode rays
Cathode rays leave the cathode, or negative electrode, and are accelerated toward the anode, or positive electrode. Some of the rays pass through the hole in the anode to form a beam, which then is bent by the electric plates in the tube.

Discovery of the Electron

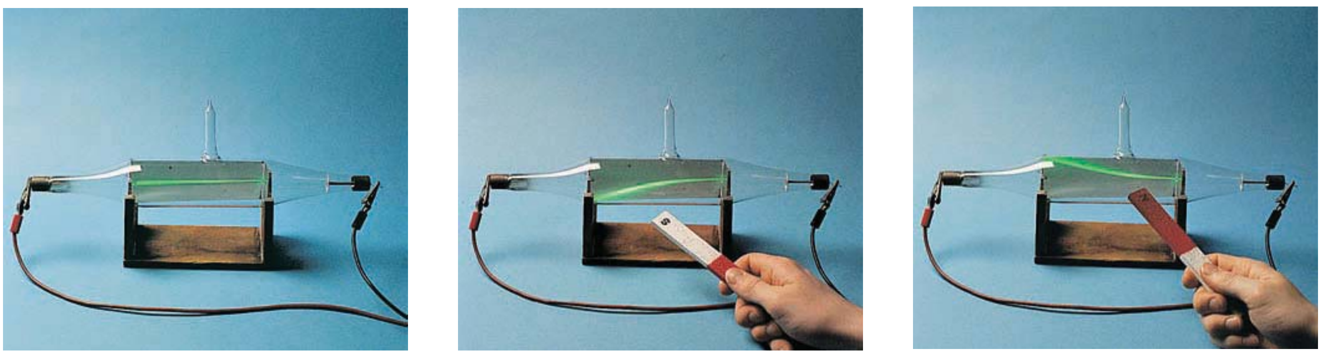
In 1897 the British physicist J. J. Thomson (Figure 2.3) conducted a series of experiments that showed that atoms were not indivisible particles. Figure 2.4 shows an experimental apparatus similar to the one used by Thomson. In this apparatus, two electrodes from a high-voltage source are sealed into a glass tube from which the air has been evacuated. The negative electrode is called the *cathode;* the positive one, the *anode.* When the high-voltage current is turned on, the glass tube emits a greenish light. Experiments showed that this greenish light is caused by the interaction of the glass with *cathode rays,* which are rays that originate from the cathode. <

A television tube uses the deflection of cathode rays by magnetic fields. A beam of cathode rays is directed toward a coated screen on the front of the tube, where by varying the magnetism generated by electromagnetic coils, the beam traces a luminescent image.

After the cathode rays leave the negative electrode, they move toward the anode, where some rays pass through a hole to form a beam (Figure 2.4). This beam bends away from the negatively charged plate and toward the positively charged plate. (Cathode rays are not directly visible, but do cause certain materials such as zinc sulfide to glow so you can observe them.) Figure 2.5 shows a similar experiment, in which cathode rays are seen to bend when a magnet is brought toward them. Thomson showed that the characteristics of cathode rays are independent of the material making up the cathode. From such evidence, he concluded that a cathode ray consists of a beam of negatively charged particles (or electrons) and that electrons are constituents of all matter.

FIGURE 2.5

Bending of cathode rays by a magnet
Left: The cathode-ray beam travels from right to left. It is visible where it falls on a zinc-sulfide screen. *Center:* The beam of negative particles bends downward as the south pole of the magnet is brought toward it. *Right:* When the magnet is turned around, the beam bends in the opposite direction.



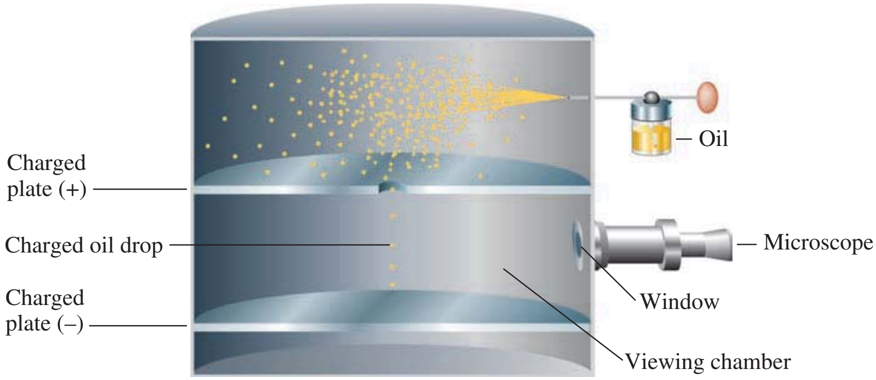


FIGURE 2.6

Millikan's oil-drop experiment
An atomizer, or spray bottle, introduces a fine mist of oil drops into the top chamber. Several drops happen to fall through a small hole into the lower chamber, where the experimenter follows the motion of one drop with a microscope. Some of these drops have picked up one or more electrons as a result of friction in the atomizer and have become negatively charged. A negatively charged drop will be attracted upward when the experimenter turns on a current to the electric plates. The drop's upward speed (obtained by timing its rise) is related to its mass-to-charge ratio, from which you can calculate the charge on the electron.

From his experiments, Thomson could also calculate the ratio of the electron's mass, m_e, to its electric charge, e. He could not obtain either the mass or the charge separately, however. In 1909 the U.S. physicist Robert Millikan performed a series of ingenious experiments in which he obtained the charge on the electron by observing how a charged drop of oil falls in the presence and in the absence of an electric field (Figure 2.6). From this type of experiment, the charge on the electron is found to be 1.602×10^{-19} coulombs (the *coulomb,* abbreviated C, is a unit of electric charge). If you use this charge with the most recent value of the mass-to-charge ratio of the electron, you obtain an electron mass of 9.109×10^{-31} kg, which is more than 1800 times smaller than the mass of the lightest atom (hydrogen). This shows quite clearly that the electron is indeed a subatomic particle.

The Nuclear Model of the Atom

Ernest Rutherford (1871–1937), a British physicist, put forth the idea of the *nuclear model* of the atom in 1911, based on experiments done in his laboratory by Hans Geiger and Ernest Marsden. These scientists observed the effect of bombarding thin gold foil (and other metal foils) with alpha radiation from radioactive substances such as uranium (Figure 2.7). Rutherford had already shown that alpha rays consist of positively charged particles. Geiger and Marsden found that most of the alpha particles passed through a metal foil as though nothing were there, but a few (about 1 in 8000) were scattered at large angles and sometimes almost backward.

According to Rutherford's model, most of the mass of the atom (99.95% or more) is concentrated in a positively charged center, or nucleus, around which the negatively charged electrons move. Although most of the mass of an atom is in its nucleus, the nucleus occupies only a very small portion of the space of the atom. Nuclei have diameters of about 10^{-15} m (10^{-3} pm), whereas atomic diameters are about 10^{-10} m (100 pm), a hundred thousand times larger. If you were to use a golf ball to represent the nucleus, the atom would be about 3 miles in diameter!

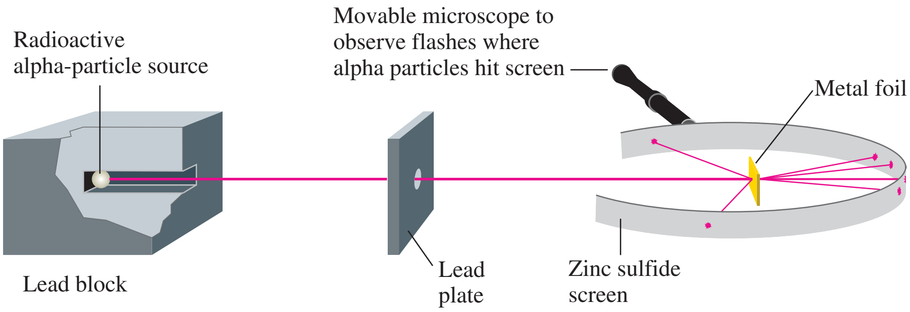


FIGURE 2.7

Alpha-particle scattering from metal foils
Alpha radiation is produced by a radioactive source and formed into a beam by a lead plate with a hole in it. (Lead absorbs the radiation.) Scattered alpha particles are made visible by a zinc sulfide screen, which emits tiny flashes where particles strike it. A movable microscope is used for viewing the flashes.

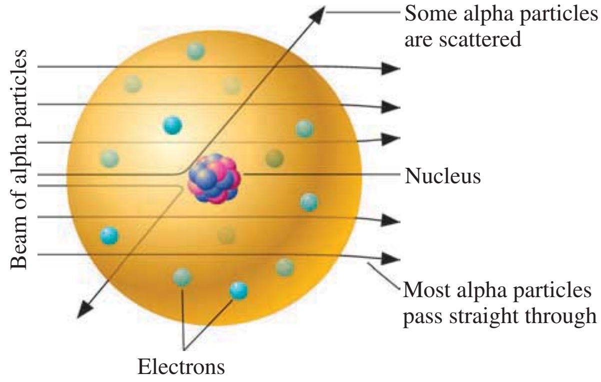


FIGURE 2.8

Representation of the scattering of alpha particles by a gold foil
Most of the alpha particles pass through the foil barely deflected. A few, however, collide with gold nuclei and are deflected at large angles. (The relative sizes of nuclei are smaller than can be drawn here.)

The nuclear model easily explains the results of bombarding gold and other metal foils with alpha particles. Alpha particles are much lighter than gold atoms. (Alpha particles are helium nuclei.) Most of the alpha particles pass through the metal atoms of the foil, undeflected by the lightweight electrons. When an alpha particle does happen to hit a metal-atom nucleus, however, it is scattered at a wide angle because it is deflected by the massive, positively charged nucleus (Figure 2.8).

Concept Check 2.2

What would be a feasible model for the atom if Geiger and Marsden had found that 7999 out of 8000 alpha particles were deflected back at the alpha-particle source?

2.3 Nuclear Structure; Isotopes

The nucleus of an atom also has a structure; the nucleus is composed of two different kinds of particles, protons and neutrons. The type of alpha-particle scattering experiment that led to the nuclear model of the atom was also instrumental in clarifying this structure of the nucleus.

An important property of the nucleus is its electric charge. One way to determine the value of the positive charge of a nucleus is by analyzing the distribution of alpha particles scattered from a metal foil; other experiments also provide the nuclear charge. < From such experiments, researchers discovered that each element has a unique nuclear charge that is an integer multiple of the magnitude of the electron charge. This integer, which is characteristic of an element, is called the *atomic number* (Z). A hydrogen atom nucleus, whose magnitude of charge equals that of the electron, has the smallest atomic number, which is 1.

The simplest way to obtain the nuclear charge is by analyzing the x rays emitted by the element when irradiated with cathode rays. This is discussed in the essay at the end of Section 8.2.

In 1919, Rutherford discovered that hydrogen nuclei, or what we now call protons, form when alpha particles strike some of the lighter elements, such as nitrogen. A **proton** is *a nuclear particle having a positive charge equal to that of the electron and a mass more than 1800 times that of the electron.* When an alpha particle collides with a nitrogen atom, a proton may be knocked out of the nitrogen nucleus. The protons in a nucleus give the nucleus its positive charge.

The **atomic number** (Z) is therefore *the number of protons in the nucleus of an atom.* Because you can determine the atomic number experimentally, you can determine unambiguously whether or not a sample is a pure element or if perhaps you have discovered a new element. We can now state the definition of an element with more precision than we could in Section 2.1. An **element** is *a substance whose*

TABLE 2.1 Properties of the Electron, Proton, and Neutron

Particle	Mass (kg)	Charge (C)	Mass (amu)*	Charge (e)
Electron	9.10939×10^{-31}	-1.60218×10^{-19}	0.00055	-1
Proton	1.67262×10^{-27}	$+1.60218 \times 10^{-19}$	1.00728	$+1$
Neutron	1.67493×10^{-27}	0	1.00866	0

*The atomic mass unit (amu) equals 1.66054×10^{-27} kg; it is defined in Section 2.4.

atoms all have the same atomic number. The inside back cover of the book lists the elements and their atomic numbers (see the Table of Atomic Numbers and Atomic Masses).

The neutron was also discovered by alpha-particle scattering experiments. When beryllium metal is irradiated with alpha rays, a strongly penetrating radiation is obtained from the metal. In 1932 the British physicist James Chadwick (1891–1974) showed that this penetrating radiation consists of neutral particles, or neutrons. The **neutron** is *a nuclear particle having a mass almost identical to that of the proton but no electric charge.* When beryllium nuclei are struck by alpha particles, neutrons are knocked out. Table 2.1 compares the masses and charges of the electron and the two nuclear particles, the proton and the neutron.

Now consider the nucleus of some atom, say of the naturally occurring sodium atom. The nucleus contains 11 protons and 12 neutrons. Thus, the charge on the sodium nucleus is $+11e$, which we usually write as simply $+11$, meaning $+11$ units of electron charge e. Similarly, the nucleus of a naturally occurring aluminum atom contains 13 protons and 14 neutrons; the charge on the nucleus is $+13$.

We characterize a nucleus by its atomic number (Z) and its mass number (A). The **mass number** (A) is *the total number of protons and neutrons in a nucleus.* The nucleus of the naturally occurring sodium atom has an atomic number of 11 and a mass number of 23 ($11 + 12$).

A **nuclide** is *an atom characterized by a definite atomic number and mass number.* The shorthand notation for any nuclide consists of the symbol of the element with the atomic number written as a subscript on the left and the mass number as a superscript on the left. You write the *nuclide symbol* for the naturally occurring sodium nuclide as follows:

$$\begin{matrix}\text{Mass number} \longrightarrow \\ \text{Atomic number} \longrightarrow\end{matrix} {}^{23}_{11}\text{Na}$$

An atom is normally electrically neutral, so it has as many electrons about its nucleus as the nucleus has protons; that is, the number of electrons in a neutral atom equals its atomic number. A sodium atom has a nucleus of charge $+11$, and around this nucleus are 11 electrons (with a charge of -11, giving the atom a charge of 0).

All nuclei of the atoms of a particular element have the same atomic number, but the nuclei may have different mass numbers. **Isotopes** are *atoms whose nuclei have the same atomic number but different mass numbers; that is, the nuclei have the same number of protons but different numbers of neutrons.* Sodium has only one naturally occurring isotope, which we denote by the same symbol we used for the nuclide (${}^{23}_{11}\text{Na}$); we also call the isotope sodium-23. Naturally occurring oxygen is a mixture of isotopes; it contains 99.759% oxygen-16, 0.037% oxygen-17, and 0.204% oxygen-18.

Figure 2.9 may help you to visualize the relationship among the different subatomic particles in the isotopes of an element. The size of the nucleus as represented

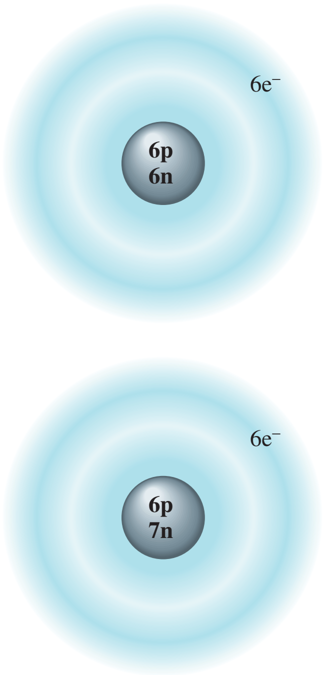


FIGURE 2.9

A representation of two isotopes of carbon

The drawing shows the basic particles making up the carbon-12 and carbon-13 isotopes. (The relative sizes of the nuclei are much exaggerated here.)

Isotopes were first suspected in about 1912 when chemically identical elements with different atomic masses were found in radioactive materials. The most convincing evidence, however, came from the mass spectrometer, discussed in Section 2.4.

in the figure is very much exaggerated in order to show the numbers of protons and neutrons. In fact, the space taken up by the nucleus is minuscule compared with the region occupied by electrons. Electrons, although extremely light particles, move throughout relatively large diffuse regions, or "shells," about the nucleus of an atom.

Example 2.1 Writing Nuclide Symbols

What is the nuclide symbol for a nucleus that contains 38 protons and 50 neutrons?

Problem Strategy To solve this problem, we need to keep in mind that the number of protons in the nucleus (the atomic number) is what uniquely identifies an element and that the mass number is the sum of the protons and neutrons.

Solution From the Table of Atomic Numbers and Atomic Masses on the inside back cover of this book, you will note that the element with atomic number 38 is strontium, symbol Sr. The mass number is 38 + 50 = 88. The symbol is $^{88}_{38}\mathbf{Sr}$.

Answer Check When writing elemental symbols, as you have done here, make sure that you always capitalize *only* the first letter of the symbol.

Exercise 2.1 A nucleus consists of 17 protons and 18 neutrons. What is its nuclide symbol?

■ See Problems 2.47 and 2.48.

2.4 Atomic Masses

A central feature of Dalton's atomic theory was the idea that an atom of an element has a characteristic mass. Now we know that a naturally occurring element may be a mixture of isotopes, each isotope having its own characteristic mass. However, the percentages of the different isotopes in most naturally occurring elements have remained essentially constant over time and in most cases are independent of the origin of the element. Thus, what Dalton actually calculated were *average* atomic masses (actually average *relative* masses, as we will discuss). Since you normally work with naturally occurring mixtures of elements (whether in pure form or as compounds), it is indeed these average atomic masses that you need in chemical work. <

Some variation of isotopic composition occurs in a number of elements, and this limits the significant figures in their average atomic masses.

Relative Atomic Masses

Dalton could not weigh individual atoms. What he could do was find the average mass of one atom relative to the average mass of another. We will refer to these relative atomic masses as *atomic masses.* To see how Dalton obtained his atomic masses, imagine that you burn hydrogen gas in oxygen. The product of this reaction is water, a compound of hydrogen and oxygen. By experiment, you find that 1.0000 gram of hydrogen reacts with 7.9367 g of oxygen to form water. To obtain the atomic mass of oxygen (relative to hydrogen), you need to know the relative numbers of hydrogen atoms and oxygen atoms in water. As you might imagine, Dalton had difficulty with this. Today we know that water contains two atoms of hydrogen for every atom of oxygen. So the atomic mass of oxygen is $2 \times 7.9367 = 15.873$ times that of the mass of the average hydrogen atom. <

FIGURE 2.10 ▲
A mass spectrometer
This instrument measures the masses of atoms and molecules.

Dalton had assumed the formula for water to be HO. Using this formula and accurate data, he would have obtained 7.9367 for the relative atomic weight of oxygen.

Atomic Mass Units

Dalton's hydrogen-based atomic mass scale was eventually replaced by a scale based on oxygen and then, in 1961, by the present carbon-12 mass scale. This scale depends on measurements of atomic mass by an instrument called a *mass spectrometer* (Figure 2.10), which we will describe briefly later in this section. You make accurate

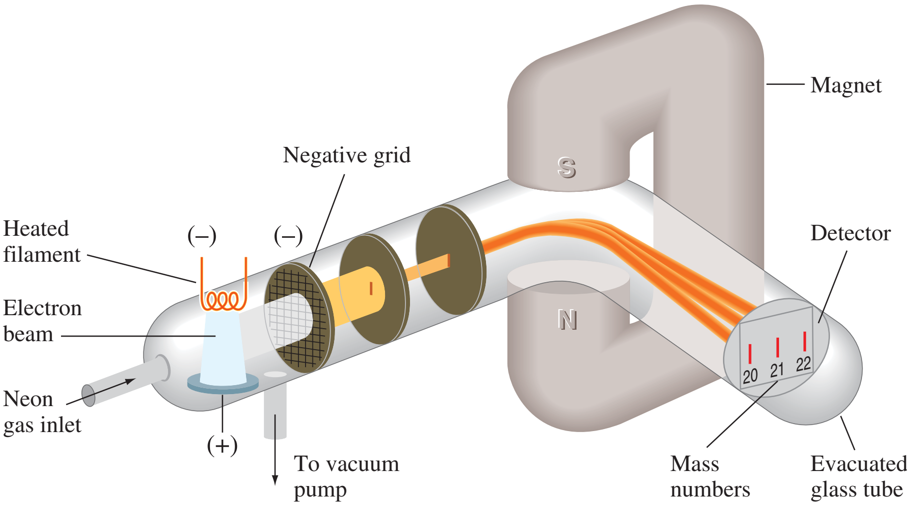


FIGURE 2.11

Diagram of a simple mass spectrometer, showing the separation of neon isotopes
Neon gas enters an evacuated chamber, where neon atoms form positive ions when they collide with electrons. Positively charged neon atoms, Ne^+, are accelerated from this region by the negative grid and pass between the poles of a magnet. The beam of positively charged atoms is split into three beams by the magnetic field according to the mass-to-charge ratios. The three beams then travel to a detector at the end of the tube. (The detector is shown here as a photographic plate; in modern spectrometers, the detector is electronic, and the mass positions are recorded on a computer.)

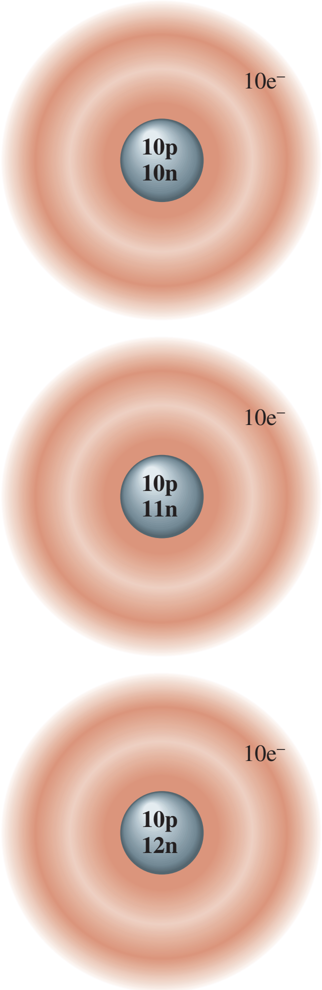


FIGURE 2.12

Representations of the three naturally occurring isotopes of Ne
Top to bottom: Ne-20, Ne-21, and Ne-22. Ne-20 is the most abundant isotope.

measurements of mass on this instrument by comparing the mass of an atom to the mass of a particular atom chosen as a standard. On the present atomic mass scale, the carbon-12 isotope is chosen as the standard and is arbitrarily assigned a mass of exactly 12 atomic mass units. One **atomic mass unit (amu)** is, therefore, *a mass unit equal to exactly one-twelfth the mass of a carbon-12 atom.*

On this modern scale, the **atomic mass** of an element is *the average atomic mass for the naturally occurring element, expressed in atomic mass units.* A complete table of atomic masses appears on the inside back cover of this book. (Atomic mass is sometimes referred to as atomic weight. Although the use of the term *weight* in *atomic weight* is not strictly correct, it is sanctioned by convention to mean "average atomic mass.")

Mass Spectrometry and Atomic Masses

Earlier we noted that you use a mass spectrometer to obtain accurate masses of atoms. The earliest mass spectrometers measured the mass-to-charge ratios of positively charged atoms using the same ideas that Thomson used to study electrons. Modern instruments may use other techniques, but all measure the mass-to-charge ratios of positively charged atoms (and molecules as well). Figure 2.11 shows a simplified sketch of a mass spectrometer running a neon sample (see the essay on page 98 for a discussion of modern mass spectrometry).

Mass spectrometers produce a *mass spectrum,* which shows the relative numbers of atoms for various masses. If a sample of neon, which has the three isotopes depicted in Figure 2.12, is introduced into a mass spectrometer, it produces the mass spectrum shown in Figure 2.13. This mass spectrum gives us all of the information we need to calculate the atomic mass of neon: the masses of all of the isotopes (neon-20, neon-21, and neon-22) and the relative number, or fractional abundance, of each isotope. The **fractional abundance** of an isotope is *the fraction of the total*

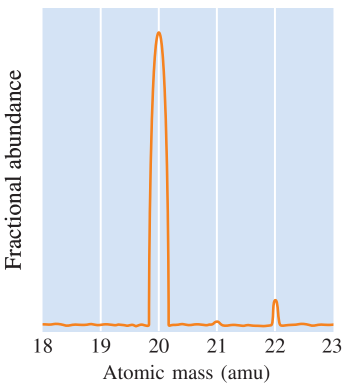


FIGURE 2.13

The mass spectrum of neon
Neon is separated into its isotopes Ne-20, Ne-21, and Ne-22. The height at each mass peak is proportional to the fraction of that isotope in the element.

number of atoms that is composed of a particular isotope. The fractional abundances of the neon isotopes in naturally occurring neon are neon-20, 0.9051; neon-21, 0.0027; and neon-22, 0.0922.

You calculate the atomic mass of an element by multiplying each isotopic mass by its fractional abundance and summing the values. If you do that for neon using the data given here, you will obtain 20.179 amu. The next example illustrates the calculation in full for chromium.

Example 2.2 Determining Atomic Mass from Isotopic Masses and Fractional Abundances

Chromium, Cr, has the following isotopic masses and fractional abundances:

Mass Number	*Isotopic Mass (amu)*	*Fractional Abundance*
50	49.9461	0.0435
52	51.9405	0.8379
53	52.9407	0.0950
54	53.9389	0.0236

What is the atomic mass of chromium?

Problem Strategy The type of average used to calculate the atomic mass of an element is similar to the "average" an instructor might use to obtain a student's final grade in a class. Suppose the student has a total exam grade of 76 and a total laboratory grade of 84. The instructor decides to give a weight of 70% to the exams and 30% to the laboratory. How would the instructor calculate the final grade? He or she would multiply each type of grade by its weight factor and add the results:

$$(76 \times 0.70) + (84 \times 0.30) = 78$$

The final grade is closer to the exam grade, because the instructor chose to give the exam grade greater weight.

Solution Multiply each isotopic mass by its fractional abundance, then sum:

$$\begin{aligned}
49.9461 \text{ amu} \times 0.0435 &= 2.17 \text{ amu} \\
51.9405 \text{ amu} \times 0.8379 &= 43.52 \text{ amu} \\
52.9407 \text{ amu} \times 0.0950 &= 5.03 \text{ amu} \\
53.9389 \text{ amu} \times 0.0236 &= \underline{1.27 \text{ amu}} \\
& 51.99 \text{ amu}
\end{aligned}$$

The atomic mass of chromium is **51.99 amu.**

Answer Check The average mass (atomic mass) should be near the mass of the isotope with greatest abundance: in this case, 51.9405 amu with fractional abundance of 0.8379. This provides a quick check on your answer to this type of problem; any "answer" that is far from this will be incorrect.

Exercise 2.2 Chlorine consists of the following isotopes:

Isotope	*Isotopic Mass (amu)*	*Fractional Abundance*
Chlorine-35	34.96885	0.75771
Chlorine-37	36.96590	0.24229

What is the atomic mass of chlorine?

■ See Problems 2.51, 2.52, 2.53, and 2.54.

2.5 Periodic Table of the Elements

In 1869 the Russian chemist Dmitri Mendeleev (1834–1907) (Figure 2.14) and the German chemist J. Lothar Meyer (1830–1895), working independently, made similar discoveries. They found that when they arranged the elements in order of atomic mass, they could place them in horizontal rows, one row under the other, so that the elements in each vertical column have similar properties. *A tabular arrangement of elements in rows and columns, highlighting the regular repetition of properties of the elements,* is called a **periodic table.**

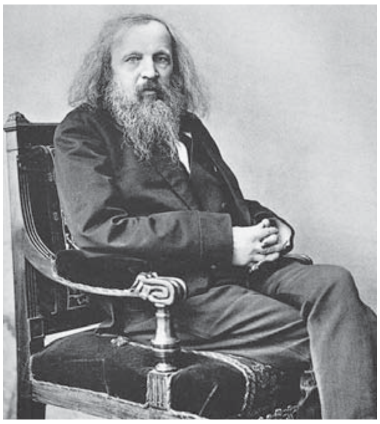


FIGURE 2.14

Dmitri Ivanovich Mendeleev (1834–1907) Mendeleev constructed a periodic table as part of his effort to systematize chemistry. He received many international honors for his work, but his reception in czarist Russia was mixed. He had pushed for political reforms and made many enemies as a result.

Eventually, more accurate determinations of atomic masses revealed discrepancies in this ordering of the elements. However, in the early part of this century, it was shown that the elements are characterized by their atomic numbers, rather than atomic masses. When the elements in the periodic table are ordered by atomic number, such discrepancies vanish.

A modern version of the periodic table, with the elements arranged by atomic number, is shown in Figure 2.15 (see also inside front cover). Each entry lists the atomic number, atomic symbol, and atomic mass of an element. This is a convenient way of tabulating such information, and you should become familiar with using the periodic table for that purpose. As we develop the subject matter of chemistry throughout the text, you will see how useful the periodic table is.

Periods and Groups

The basic structure of the periodic table is its division into rows and columns, or periods and groups. A **period** consists of *the elements in any one horizontal row of the periodic table.* A **group** consists of *the elements in any one column of the periodic table.*

The first period of elements consists of only hydrogen (H) and helium (He). The second period has 8 elements, beginning with lithium (Li) and ending with neon (Ne). There is then another period of 8 elements, and this is followed by a period having 18 elements, beginning with potassium (K) and ending with krypton (Kr). The fifth period also has 18 elements. The sixth period actually consists of 32 elements, but in order for the row to fit on a page, part of it appears at the bottom of the table. Otherwise the table would have to be expanded, with the additional elements placed after barium (Ba, atomic number 56). The seventh period,

FIGURE 2.15

A modern form of the periodic table This table is also given on the inside front cover of the book.

Main-Group Elements (groups 1–2); Transition Metals (groups 3–12); Main-Group Elements (groups 13–18)

Key: Atomic number / Symbol / Atomic mass (e.g., 1 **H** 1.00794)

Period	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIIIB	9 VIIIB	10 VIIIB	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1	1 **H** 1.00794																	2 **He** 4.002602
2	3 **Li** 6.941	4 **Be** 9.012182											5 **B** 10.811	6 **C** 12.0107	7 **N** 14.0067	8 **O** 15.9994	9 **F** 18.9984032	10 **Ne** 20.1797
3	11 **Na** 22.98976928	12 **Mg** 24.3050											13 **Al** 26.9815386	14 **Si** 28.0855	15 **P** 30.973762	16 **S** 32.065	17 **Cl** 35.453	18 **Ar** 39.948
4	19 **K** 39.0983	20 **Ca** 40.078	21 **Sc** 44.955912	22 **Ti** 47.867	23 **V** 50.9415	24 **Cr** 51.9961	25 **Mn** 54.938045	26 **Fe** 55.845	27 **Co** 58.933195	28 **Ni** 58.6934	29 **Cu** 63.546	30 **Zn** 65.409	31 **Ga** 69.723	32 **Ge** 72.64	33 **As** 74.92160	34 **Se** 78.96	35 **Br** 79.904	36 **Kr** 83.798
5	37 **Rb** 85.4678	38 **Sr** 87.62	39 **Y** 88.90585	40 **Zr** 91.224	41 **Nb** 92.90638	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.07	45 **Rh** 102.90550	46 **Pd** 106.42	47 **Ag** 107.8682	48 **Cd** 112.411	49 **In** 114.818	50 **Sn** 118.710	51 **Sb** 121.760	52 **Te** 127.60	53 **I** 126.90447	54 **Xe** 131.293
6	55 **Cs** 132.9054519	56 **Ba** 137.327	71 **Lu** 174.967	72 **Hf** 178.49	73 **Ta** 180.94788	74 **W** 183.84	75 **Re** 186.207	76 **Os** 190.23	77 **Ir** 192.217	78 **Pt** 195.084	79 **Au** 196.966569	80 **Hg** 200.59	81 **Tl** 204.3833	82 **Pb** 207.2	83 **Bi** 208.98040	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
7	87 **Fr** (223)	88 **Ra** (226)	103 **Lr** (262)	104 **Rf** (261)	105 **Db** (262)	106 **Sg** (266)	107 **Bh** (264)	108 **Hs** (277)	109 **Mt** (268)	110 **Ds** (281)	111 **Rg** (272)	112 **Uub** (285)	113 **Uut** (284)	114 **Uuq** (289)	115 **Uup** (288)	116 **Uuh** (291)		118 **Uuo** (294)

Inner Transition Metals

Lanthanides	57 **La** 138.90547	58 **Ce** 140.116	59 **Pr** 140.90765	60 **Nd** 144.244	61 **Pm** (145)	62 **Sm** 150.36	63 **Eu** 151.964	64 **Gd** 157.25	65 **Tb** 158.92535	66 **Dy** 162.500	67 **Ho** 164.93032	68 **Er** 167.259	69 **Tm** 168.93421	70 **Yb** 173.04
Actinides	89 **Ac** (227)	90 **Th** 232.03806	91 **Pa** 231.03588	92 **U** 238.02891	93 **Np** (237)	94 **Pu** (244)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)

Metal

Metalloid

Nonmetal

though not complete, also has some of its elements placed as a row at the bottom of the table.

The groups are usually numbered. The numbering frequently seen in North America labels the groups with Roman numerals and A's and B's. In Europe a similar convention has been used, but some columns have the A's and B's interchanged. To eliminate this confusion, the International Union of Pure and Applied Chemistry (IUPAC) suggested a convention in which the columns are numbered 1 to 18. Figure 2.15 shows the traditional North American and the IUPAC conventions. When we refer to an element by its periodic group, we will use the traditional North American convention. The A groups are called *main-group* (or *representative*) *elements;* the B groups are called *transition elements.* The two rows of elements at the bottom of the table are called *inner transition elements* (the first row is referred to as the *lanthanides;* the second row, as the *actinides*).

As noted earlier, the elements in any one group have similar properties. For example, the elements in Group IA, often known as the *alkali metals,* are soft metals that react easily with water. (Hydrogen, a gas, is an exception and might better be put in a group by itself.) Sodium is an alkali metal. So is potassium. The Group VIIA elements, known as *halogens,* are also reactive elements. Chlorine is a halogen. We have already noted its vigorous reaction with sodium. Bromine, which is a red-brown liquid, is another halogen. It too reacts vigorously with sodium.

Metals, Nonmetals, and Metalloids

The elements of the periodic table in Figure 2.15 are divided by a heavy "staircase" line into metals on the left and nonmetals on the right. A **metal** is *a substance or mixture that has a characteristic luster, or shine, and is generally a good conductor of heat and electricity.* Except for mercury, the metallic elements are solids at room temperature (about 20°C). They are more or less *malleable* (can be hammered into sheets) and *ductile* (can be drawn into wire).

A **nonmetal** is *an element that does not exhibit the characteristics of a metal.* Most of the nonmetals are gases (for example, chlorine and oxygen) or solids (for example, phosphorus and sulfur). The solid nonmetals are usually hard, brittle substances. Bromine is the only liquid nonmetal.

Most of the elements bordering the staircase line in the periodic table (Figure 2.15) are metalloids, or semimetals. A **metalloid,** or **semimetal,** is *an element having both metallic and nonmetallic properties.* These elements, such as silicon (Si) and germanium (Ge), are usually good *semiconductors*—elements that, when pure, are poor conductors of electricity at room temperature but become moderately good conductors at higher temperatures. <

When these pure semiconductor elements have small amounts of certain other elements added to them (a process called doping*), they become very good conductors of electricity. Semiconductors are the critical materials in solid-state electronic devices.*

Exercise 2.3 By referring to the periodic table (Figure 2.15 or inside front cover), identify the group and period to which each of the following elements belongs. Then decide whether the element is a metal, nonmetal, or metalloid.

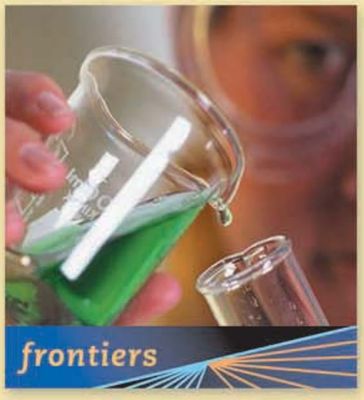
a. Se b. Cs c. Fe d. Cu e. Br

■ See Problems 2.57 and 2.58.

Concept Check 2.3

Consider the elements He, Ne, and Ar. Can you come up with a reason why they are in the same group in the periodic table?

A Chemist Looks at . . .



Thirty Seconds on the Island of Stability

It appears that scientists have now been on the "Island of Stability"—for at least 30 seconds. In recent years, the island of stability has attracted nuclear scientists the way Mount Everest has attracted mountain climbers. Like mountain climbers, scientists have wanted to surmount their own tallest peak—discovery of the relatively stable, superheavy nuclide at the peak of stability that is surrounded by foothills comprised of less stable nuclides. Discovering the superheavy nuclide at the peak would also mean discovering a new element—a task that has had its allure since Lavoisier gave currency to the concept of elements in the late eighteenth century.

Theoretical physicists predicted the existence of the island of stability, centered at element 114 with mass number 298, in the 1960s. The term *stability* refers here to that of the atomic nucleus. An unstable nucleus tends to fall apart by radioactive decay—a piece spontaneously flies off the nucleus, leaving a different one behind. Approximately 275 nuclides are completely stable, or nonradioactive. All of these nuclides have atomic numbers (or numbers of protons) no greater than 83 (the atomic number for the element bismuth). Beyond bismuth, all elements are radioactive and become increasingly unstable. In fact, none of the original, primeval elements past uranium (element 92)—the *transuranium* elements—exists any longer; they have long since vanished by radioactive decay. Scientists have made transuranium elements in the laboratory, however.

In February 1996, for example, scientists at Darmstadt, Germany, produced a few atoms of element 112 by bombarding a lead target with zinc atoms. Each atom of element 112 lasted only about 240 microseconds. We can represent the result by a nuclear equation, which is similar to a chemical equation (described at the end of the chapter).

$$^{70}_{30}\mathrm{Zn} + ^{208}_{82}\mathrm{Pb} \longrightarrow ^{277}_{112}\mathrm{Uub} + ^{1}_{0}\mathrm{n}$$

We introduced the nuclide symbols used here in Section 2.3. The equation is a summary of the following. A high-speed zinc-70 nucleus (as an ion, which is an electrically charged atom) smashes into a lead-208 target where, as the arrow denotes, it yields new products, a nucleus of element 112 (mass number 277) and a neutron, denoted $^{1}_{0}\mathrm{n}$. Element 112 has the provisional name of ununbium until several other laboratories can reproduce the work, after

Chemical Substances: Formulas and Names

Atomic theory has developed steadily since Dalton's time and has become the cornerstone of chemistry. It results in an enormous simplification: all of the millions of compounds we know today are composed of the atoms of just a few elements. Now we look more closely at how we describe the composition and structure of chemical substances in terms of atoms.

2.6 Chemical Formulas; Molecular and Ionic Substances

The **chemical formula** of a substance is *a notation that uses atomic symbols with numerical subscripts to convey the relative proportions of atoms of the different elements in the substance.* Consider the formula of aluminum oxide, Al_2O_3. This means that the compound is composed of aluminum atoms and oxygen atoms in the ratio 2 : 3. Consider the formula for sodium chloride, NaCl. When no subscript is written for a symbol, it is assumed to be 1. Therefore, the formula NaCl means which the discoverer may name it. Figure 2.16 shows the ion accelerator (UNILAC) that produced the zinc ions for this experiment.

Scientists in Dubna, north of Moscow in Russia, and their American collaborators from Lawrence Livermore Laboratory in California produced one atom of element 114 (provisional name ununquadium) in December 1998, placing scientists on the shore of the island of stability. In this experiment, scientists bombarded a plutonium-244 target with calcium-48 atoms.

$$^{48}_{20}\mathrm{Ca} + ^{244}_{94}\mathrm{Pu} \longrightarrow ^{289}_{114}\mathrm{Uuq} + 3^{1}_{0}\mathrm{n}$$

The atom of element 114 lasted about 30 seconds. While a 30-second lifetime may seem brief, it is many times longer than that seen in the previously discovered element 112. This same group of scientists has recently reported (in 2001) the discovery of element 116 (ununhexium).

As a rather bizarre footnote to this story, the "discovery" of element 118 was reported in 1999 by scientists at Lawrence Berkeley National Laboratory in California. Two years later, these scientists reported that they were unable to reproduce their results. The laboratory fired one of the scientists, who they alleged had faked the data. Fortunately, fraud is rather unusual in chemistry; when it does occur, it is discovered, as in this case, when the work cannot be reproduced. Notably, element 118 was rediscovered in 2006 by the same Russian and American team who discovered elements 114 and 116, bringing scientists one step closer to the island of stability.

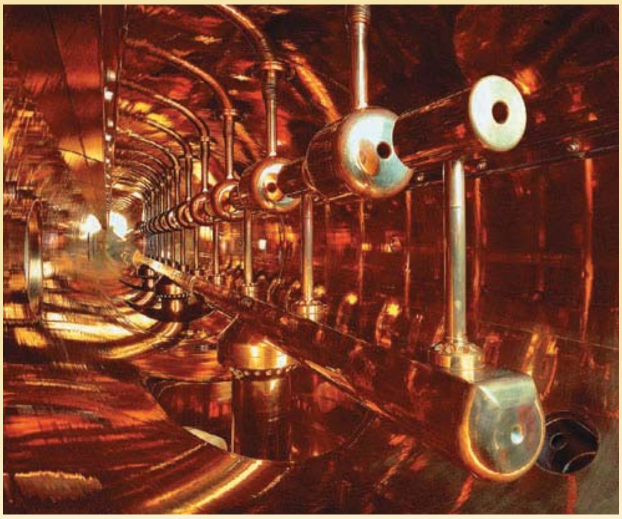


FIGURE 2.16 ▲

Within the UNILAC ion accelerator

The accelerator consists of a series of electrodes arranged one after the other in a copper-plated steel container. Each electrode accelerates the ion. Eventually, the ion gains sufficient speed, and therefore energy, to cause its nucleus to fuse on collision with a target atom nucleus giving a new nucleus.

■ See Problems 2.131 and 2.132.

that the compound is composed of sodium atoms and chlorine atoms in the ratio 1 : 1 (Figure 2.17).

Additional information may be conveyed by different kinds of chemical formulas. To understand this, we need to look briefly at two main types of substances: molecular and ionic.

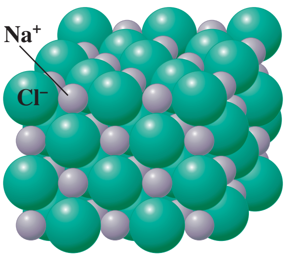


FIGURE 2.17 ▲

Sodium chloride model

This model of a sodium chloride crystal illustrates the 1 : 1 packing of the Na^+ and Cl^- ions.

Molecular Substances

A **molecule** is *a definite group of atoms that are chemically bonded together—that is, tightly connected by attractive forces.* The nature of these strong forces is discussed in Chapters 9 and 10. A *molecular substance* is a substance that is composed of molecules, all of which are alike. The molecules in such a substance are so small that even extremely minute samples contain tremendous numbers of them. One billionth (10^{-9}) of a drop of water, for example, contains about 2 trillion (2×10^{12}) water molecules. <

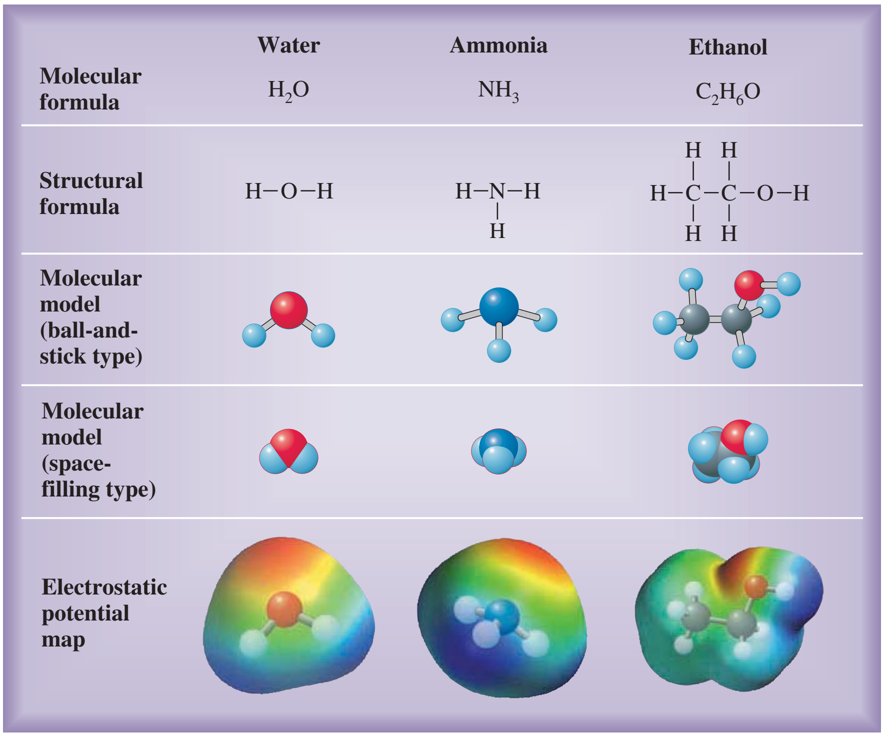
Another way to understand the large numbers of molecules involved in relatively small quantities of matter is to consider 1 g of water (about one-fifth teaspoon). It contains 3.3×10^{22} water molecules. If you had a penny for every molecule in this quantity of water, the height of your stack of pennies would be about 300 million times the distance from the earth to the sun.

A **molecular formula** *gives the exact number of different atoms of an element in a molecule.* The hydrogen peroxide molecule contains two hydrogen atoms and two oxygen atoms chemically bonded. Therefore, its molecular formula is H_2O_2. Other simple molecular substances are water, H_2O; ammonia, NH_3; carbon dioxide, CO_2; and ethanol (ethyl alcohol), C_2H_6O. In Chapter 3, you will see how we determine such formulas.

FIGURE 2.18

Examples of molecular and structural formulas, molecular models, and electrostatic potential maps

Three common molecules—water, ammonia, and ethanol—are shown. The electrostatic potential map representation at the bottom of the figure illustrates the distribution of electrons in the molecule using a color spectrum. Colors range from red (relatively high electron density) all the way to blue (low electron density).



The atoms in a molecule are not simply piled together randomly. Rather, the atoms are chemically bonded in a definite way. A *structural formula* is a chemical formula that shows how the atoms are bonded to one another in a molecule. For example, it is known that each of the hydrogen atoms in the water molecule is bonded to the oxygen atom. Thus, the structural formula of water is H—O—H. A line joining two atomic symbols in such a formula represents the chemical bond connecting the atoms. Figure 2.18 shows some structural formulas. Structural formulas are sometimes condensed in writing. For example, the structural formula of ethanol may be written CH_3CH_2OH or C_2H_5OH, depending on the detail you want to convey.

The atoms in a molecule are not only connected in definite ways but exhibit definite spatial arrangements as well. Chemists often construct molecular models as an aid in visualizing the shapes and sizes of molecules. Figure 2.18 shows molecular models for several compounds. While the ball-and-stick type of model shows the bonds and bond angles clearly, the space-filling type gives a more realistic feeling of the space occupied by the atoms. Chemists also use computer-generated models of molecules, which can be produced in a variety of forms.

Some elements are molecular substances and are represented by molecular formulas. Chlorine, for example, is a molecular substance and has the formula Cl_2, each molecule being composed of two chlorine atoms bonded together. Sulfur consists of molecules composed of eight atoms; its molecular formula is S_8. Helium and neon are composed of isolated atoms; their formulas are He and Ne, respectively. Other elements, such as carbon (in the form of graphite or diamond), do not have a simple molecular structure but consist of a very large, indefinite number of atoms bonded together. These elements are represented simply by their atomic symbols. (An important exception is a form of carbon called buckminsterfullerene, which was discovered in 1985 and has the molecular formula C_{60}.) Models of some elementary substances are shown in Figure 2.19.

An important class of molecular substances is the polymers. **Polymers** are *very large molecules that are made up of a number of smaller molecules repeatedly linked together.* **Monomers** are *the small molecules that are linked together to form the polymer.* A good analogy for the formation of polymers from monomers is that of making a chain of paper clips. Say you have boxes of red, blue, and yellow paper clips.

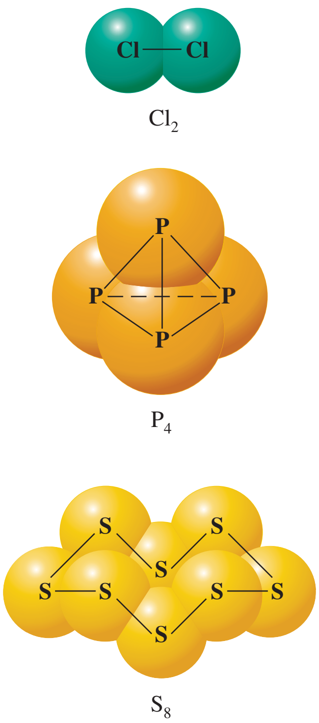


FIGURE 2.19

Molecular models of some elementary substances

Top to bottom: Chlorine, Cl_2; white phosphorus, P_4; and sulfur, S_8.

Each paper clip represents a monomer. You can make a large chain of paper clips, representing the polymer, with a variety of patterns and repeating units. One chain might involve a repeating pattern of two red, three yellow, and one blue. Another chain might consist of only blue paper clips. Just like the paper clips, the creation of a particular polymer in a laboratory involves controlling both the monomers that are being linked together and the linkage pattern.

Polymers are both natural and synthetic. Hard plastic soda bottles are made from chemically linking two different monomers in an alternating pattern. Wool and silk are natural polymers of amino acids linked by peptide bonds. Nylon® for fabrics, Kevlar® for bulletproof vests, and Nomex® for flame-retardant clothing all contain a CONH linkage. Plastics and rubbers are also polymers that are made from carbon- and hydrogen-containing monomers. Even the Teflon® coating on cookware is a polymer that is the result of linking CF_2CF_2 monomers (Figure 2.20). From these examples, it is obvious that polymers are very important molecular materials that we use every day in a wide variety of applications. For more on the chemistry of polymers, refer to Chapter 24.

FIGURE 2.20 ▲

$F_2C{=}CF_2$ monomer and Teflon®
Top: Model of the monomer used to make Teflon®, CF_2CF_2. *Center:* Model showing linkage of CF_2CF_2 monomers that make Teflon®. *Bottom:* Pan with Teflon® coating.

Ionic Substances

Although many substances are molecular, others are composed of ions (pronounced "eye′-ons"). An **ion** is *an electrically charged particle obtained from an atom or chemically bonded group of atoms by adding or removing electrons.* Sodium chloride is a substance made up of ions.

Although isolated atoms are normally electrically neutral and therefore contain equal numbers of positive and negative charges, during the formation of certain compounds atoms can become ions. Metal atoms tend to lose electrons, whereas nonmetals tend to gain electrons. When a metal atom such as sodium and a nonmetal atom such as chlorine approach one another, an electron can transfer from the metal atom to the nonmetal atom to produce ions.

An atom that picks up an extra electron becomes *a negatively charged ion,* called an **anion** (pronounced "an′-ion"). An atom that loses an electron becomes *a positively charged ion,* called a **cation** ("cat′-ion"). A sodium atom, for example, can lose an electron to form a sodium cation (denoted Na^+). A chlorine atom can gain an electron to form a chloride anion (denoted Cl^-). A calcium atom can lose two electrons to form a calcium cation, denoted Ca^{2+}. Note that the positive-two charge on the ion is indicated by a superscript 2+.

Some ions consist of two or more atoms chemically bonded but having an excess or deficiency of electrons so that the unit has an electric charge. An example is the sulfate ion, SO_4^{2-}. The superscript 2− indicates an excess of two electrons on the group of atoms.

An **ionic compound** is *a compound composed of cations and anions.* Sodium chloride consists of equal numbers of sodium ions, Na^+, and chloride ions, Cl^-. The strong attraction between positive and negative charges holds the ions together in a regular arrangement in space. For example, in sodium chloride, each Na^+ ion is surrounded by six Cl^- ions, and each Cl^- ion is surrounded by six Na^+ ions. The result is a *crystal,* which is a kind of solid having a regular three-dimensional arrangement of atoms, molecules, or (as in the case of sodium chloride) ions. Figure 2.21 shows sodium chloride crystals and two types of models used to depict the arrangement of the ions in the crystal. The number of ions in an individual sodium chloride crystal determines the size of the crystal.

The formula of an ionic compound is written by giving the smallest possible integer number of different ions in the substance, except that the charges on the ions are omitted so that the formulas merely indicate the atoms involved. For example, sodium chloride contains equal numbers of Na^+ and Cl^- ions. The formula is written NaCl (not Na^+Cl^-). Iron(III) sulfate is a compound consisting of iron(III) ions, Fe^{3+}, and sulfate ions, SO_4^{2-}, in the ratio 2 : 3. The formula is written $Fe_2(SO_4)_3$, in which parentheses enclose the formula of an ion composed of more than one atom (again, omitting ion charges); parentheses are used only when there are two or more such ions.

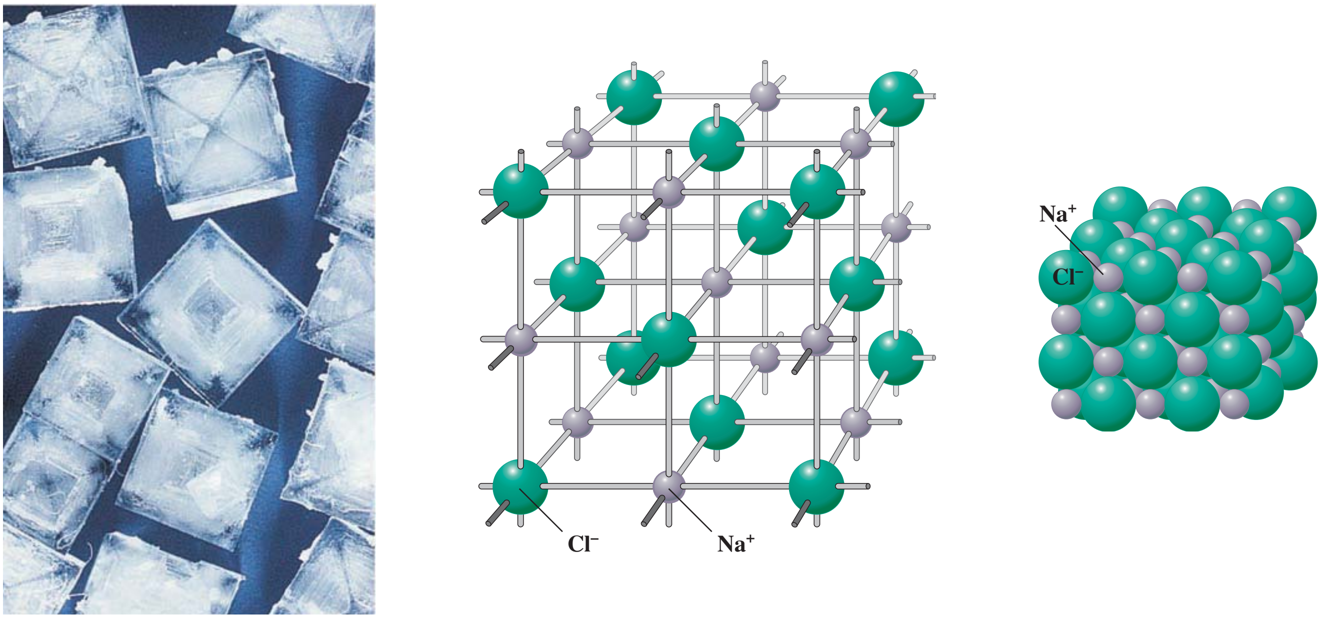


FIGURE 2.21

The sodium chloride crystal

Left: A photograph showing crystals of sodium chloride. *Center:* A model of a portion of a crystal detailing the regular arrangement of sodium ions and chloride ions. Each sodium ion is surrounded by six chloride ions, and each chloride ion is surrounded by six sodium ions. *Right:* A model highlighting the packing arrangement of the sodium and chloride ions in a solid sodium chloride crystal.

Although ionic substances do not contain molecules, we can speak of the smallest unit of such a substance. The **formula unit** of a substance is *the group of atoms or ions explicitly symbolized in the formula.* For example, the formula unit of water, H_2O, is the H_2O molecule. The formula unit of iron(III) sulfate, $Fe_2(SO_4)_3$, consists of two Fe^{3+} ions and three SO_4^{2-} ions. The formula unit is the smallest unit of such substances.

All substances, including ionic compounds, are electrically neutral. You can use this fact to obtain the formula of an ionic compound, given the formulas of the ions. This is illustrated in the following example.

Example 2.3 Writing an Ionic Formula, Given the Ions

a. Chromium(III) oxide is used as a green paint pigment (Figure 2.22). It is a compound composed of Cr^{3+} and O^{2-} ions. What is the formula of chromium(III) oxide?

b. Strontium oxide is a compound composed of Sr^{2+} and O^{2-} ions. Write the formula of this compound.

FIGURE 2.22

Chromium(III) oxide

The compound is used as a green paint pigment.

Problem Strategy Because a compound is neutral, the sum of positive and negative charges equals zero. Consider the ionic compound $CaCl_2$, which consists of one Ca^{2+} ion and two Cl^- ions. The sum of the charges is

$$\underbrace{1 \times (+2)}_{\text{Positive charge}} + \underbrace{2 \times (-1)}_{\text{Negative charge}} = 0$$

(continued)

(continued)

Note that the number of calcium ions in $CaCl_2$ equals the magnitude of charge on the chloride ion (1), whereas the number of chloride ions in $CaCl_2$ equals the magnitude of charge on the calcium ion (2). In general, you use the magnitude of charge on one ion to obtain the subscript for the other ion. You may need to simplify the formula you obtain this way so that it expresses the simplest ratio of ions.

Solution a. You can achieve electrical neutrality by taking as many cations as there are units of charge on the anion and as many anions as there are units of charge on the cation. Two Cr^{3+} ions have a total charge of 6+, and three O^{2-} ions have a total charge of 6−, giving the combination a net charge of zero. The simplest ratio of Cr^{3+} to O^{2-} is 2 : 3, and the formula is $\mathbf{Cr_2O_3}$**.** Note that the charge (without its sign) on one ion becomes the subscript of the other ion.

$$Cr^{3+} \quad O^{2-} \qquad \text{or} \qquad Cr_2O_3$$

b. You see that equal numbers of Sr^{2+} and O^{2-} ions will give a neutral compound. Thus, the formula is SrO. If you use the units of charge to find the subscripts, you get

$$Sr^{2+} \quad O^{2-} \qquad \text{or} \qquad Sr_2O_2$$

The final formula is **SrO,** because this gives the simplest ratio of ions.

Answer Check When writing ionic formulas, always make certain that the formula that you write reflects the smallest whole-number ratio of the ions. For example, using the technique presented in this example, Pb^{4+} and O^{2-} combine to give Pb_2O_4; however, the correct formula is PbO_2, which reflects the simplest whole-number ratio of ions.

Exercise 2.4 Potassium chromate is an important compound of chromium (Figure 2.23). It is composed of K^+ and CrO_4^{2-} ions. Write the formula of the compound.

FIGURE 2.23

Potassium chromate

Many compounds of chromium have bright colors, which is the origin of the name of the element. It comes from the Greek word *chroma,* meaning "color."

■ See Problems 2.75 and 2.76.

2.7 Organic Compounds

An important class of *molecular substances that contain carbon combined with other elements, such as hydrogen, oxygen, and nitrogen* is **organic compounds.** Organic chemistry is the area of chemistry that is concerned with these compounds (Chapter 23 is devoted to this topic). Historically, organic compounds were restricted to those that could be produced only from living entities and were thought to contain a "vital force" based on their natural origin. The concept of the vital force was disproved in 1828 when a German chemist, Friedrick Wöhler, synthesized urea (a molecular compound in human urine, CH_4N_2O, Figure 2.24) from the molecular compounds ammonia (NH_3) and cyanic acid (HNCO). His work clearly demonstrated that a given compound is exactly the same whether it comes from a living entity or is synthesized.

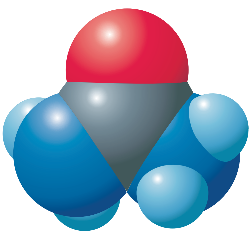
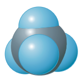


FIGURE 2.24

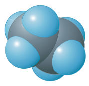
Molecular model of urea (CH_4N_2O)

Urea was the first organic molecule deliberately synthesized by a chemist from non-organic compounds.

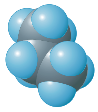
Organic compounds make up the majority of all known compounds. Since 1957, more than 13 million (60%) of the recorded substances in an international materials registry have been listed as organic. You encounter organic compounds in both living and nonliving materials every day. The proteins, amino acids, enzymes, and DNA that make up your body are all either individual organic molecules or contain organic



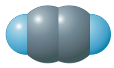
CH_4
Methane



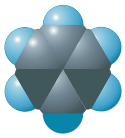
C_2H_6
Ethane



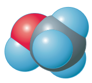
C_3H_8
Propane



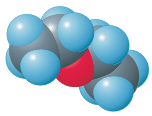
C_2H_2
Acetylene



C_6H_6
Benzene



CH_3OH
Methanol



$CH_3CH_2OCH_2CH_3$
Diethyl ether

molecules. Table sugar, peanut oil, antibiotic medicines, and methanol (windshield washer) are all examples of organic molecules as well. Organic chemistry and the compounds produced by the reactions of organic molecules are probably responsible for the majority of the materials that currently surround you as you read this page of text.

The simplest organic compounds are hydrocarbons. **Hydrocarbons** are *those compounds containing only hydrogen and carbon.* Common examples include methane (CH_4), ethane (C_2H_6), propane (C_3H_8), acetylene (C_2H_2), and benzene (C_6H_6). Hydrocarbons are used extensively as sources of energy for heating our homes, for powering internal combustion engines, and for generating electricity. They also are the starting materials for most plastics. Much of the mobility and comfort of our current civilization is built on the low cost and availability of hydrocarbons.

The chemistry of organic molecules is often determined by groups of atoms in the molecule that have characteristic chemical properties. A **functional group** is a *reactive portion of a molecule that undergoes predictable reactions.* When you use the term *alcohol* when referring to a molecular compound, you actually are indicating a molecule that contains an —OH functional group. Methyl alcohol has the chemical formula CH_3OH. The term *ether* indicates that an organic molecule contains an oxygen atom between two carbon atoms, as in diethyl ether ($CH_3CH_2OCH_2CH_3$). Table 2.2 contains a few examples of organic functional groups along with example compounds.

Chapter 23 contains a much more extensive treatment of organic chemistry. Due to its importance and its relation to living organisms, a substantial part of your future chemistry experience will likely be in the field of organic chemistry.

TABLE 2.2 Examples of Organic Functional Groups

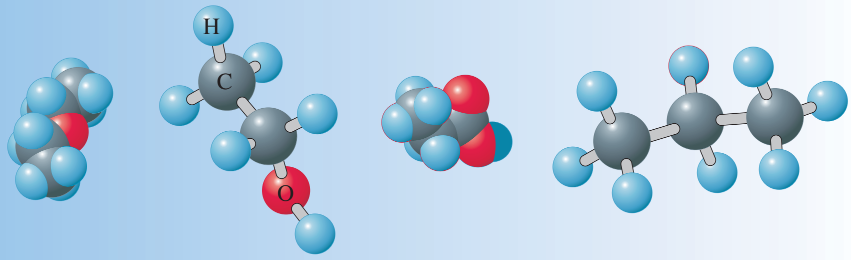
Functional Group	Name of Functional Group	Example Molecule	Common Use
—OH	Alcohol	Methyl alcohol (CH_3OH)	Windshield washer
—O—	Ether	Dimethyl ether (CH_3OCH_3)	Solvent
—COOH	Carboxylic acid	Acetic acid (CH_3COOH)	Acid in vinegar

2.8 Naming Simple Compounds

Before the structural basis of chemical substances became established, compounds were named after people, places, or particular characteristics. Examples are Glauber's salt (sodium sulfate, discovered by J. R. Glauber), sal ammoniac (ammonium chloride, named after the ancient Egyptian deity Ammon from the temple near which the substance was made), and washing soda (sodium carbonate, used for softening wash water). Today several million compounds are known and thousands of new ones are discovered every year. Without a system for naming compounds, coping with this multitude of substances would be a hopeless task. **Chemical nomenclature** is *the systematic naming of chemical compounds.*

Concept Check 2.4

Identify the following compounds as being a hydrocarbon, an alcohol, an ether, or a carboxylic acid.



(a) (b) (c) (d)

If a compound is not classified as organic, as discussed in Section 2.7, it must be inorganic. **Inorganic compounds** are *composed of elements other than carbon.* A few notable exceptions to this classification scheme include carbon monoxide, carbon dioxide, carbonates, and the cyanides; all contain carbon and yet are generally considered to be inorganic.

In this section, we discuss the nomenclature of some simple inorganic compounds. We first look at the naming of ionic compounds. Then, we look at the naming of some simple molecular compounds, including binary molecular compounds (molecular compounds of two elements) and acids. Finally, we look at hydrates of ionic compounds. These substances contain water molecules in loose association with ionic compounds.

Ionic Compounds

Ionic compounds, as we saw in the previous section, are substances composed of ions. Most ionic compounds contain metal and nonmetal atoms—for example, NaCl. (The ammonium salts, such as NH_4Cl, are a prominent exception.) You name an ionic compound by giving the name of the cation followed by the name of the anion. For example,

$$\underbrace{\text{potassium}}_{\text{cation name}}\ \underbrace{\text{sulfate}}_{\text{anion name}}$$

Before you can name ionic compounds, you need to be able to write and name ions.

The simplest ions are monatomic. A **monatomic ion** is *an ion formed from a single atom.* Table 2.3 lists common monatomic ions of the main-group elements. You

TABLE 2.3 Common Monatomic Ions of the Main-Group Elements*

	IA	IIA	IIIA	IVA	VA	VIA	VIIA
Period 1							H^-
Period 2	Li^+	Be^{2+}	B	C	N^{3-}	O^{2-}	F^-
Period 3	Na^+	Mg^{2+}	Al^{3+}	Si	P	S^{2-}	Cl^-
Period 4	K^+	Ca^{2+}	Ga^{3+}	Ge	As	Se^{2-}	Br^-
Period 5	Rb^+	Sr^{2+}	In^{3+}	Sn^{2+}	Sb	Te^{2-}	I^-
Period 6	Cs^+	Ba^{2+}	Tl^+, Tl^{3+}	Pb^{2+}	Bi^{3+}		

*Elements shown in color do not normally form compounds having monatomic ions.

may want to look at the table while you read first the rules for predicting the charges on such ions and then the rules for naming the monatomic ions.

Rules for Predicting the Charges on Monatomic Ions

1. Most of the main-group metallic elements have one monatomic cation with a charge equal to the group number in the periodic table (the Roman numeral). Example: aluminum, in Group IIIA, has a monatomic ion Al^{3+}.
2. Some metallic elements of high atomic number are exceptions to the previous rule; they have more than one cation. These elements have common cations with a charge equal to the group number minus 2, in addition to having a cation with a charge equal to the group number. Example: The common ion of lead is Pb^{2+}. (The group number is 4; the charge is $4 - 2$.) In addition to compounds containing Pb^{2+}, some lead compounds contain Pb^{4+}.
3. Most transition elements form more than one monatomic cation, each with a different charge. Most of these elements have one ion with a charge of $+2$. Example: Iron has common cations Fe^{2+} and Fe^{3+}. Copper has common cations Cu^{+} and Cu^{2+}.
4. The charge on a monatomic anion for a nonmetallic main-group element equals the group number minus 8. Example: Oxygen has the monatomic anion O^{2-}. (The group number is 6; the charge is $6 - 8$.)

Rules for Naming Monatomic Ions

1. Monatomic cations are named after the element if there is only one such ion. Example: Al^{3+} is called aluminum ion; Na^{+} is called sodium ion.
2. If there is more than one monatomic cation of an element, Rule 1 is not sufficient. The *Stock system* of nomenclature names the cations after the element, as in Rule 1, but follows this by a Roman numeral in parentheses denoting the charge on the ion. Example: Fe^{2+} is called iron(II) ion and Fe^{3+} is called iron(III) ion.

   The Roman numeral actually denotes the oxidation state, *or* oxidation number, *of the atom in the compound. For a monatomic ion, the oxidation state equals the charge. Otherwise, the oxidation state is a hypothetical charge assigned in accordance with certain rules; see Section 4.5.*

   In an older system of nomenclature, such ions are named by adding the suffixes *-ous* and *-ic* to a stem name of the element (which may be from the Latin) to indicate the ions of lower and higher charge, respectively. Example: Fe^{2+} is called ferrous ion; Fe^{3+}, ferric ion. Cu^{+} is called cuprous ion; Cu^{2+}, cupric ion.

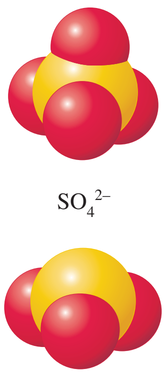
   Table 2.4 lists some common cations of the transition elements. Most of these elements have more than one ion, so require the Stock nomenclature system or the older suffix system. A few, such as zinc, have only a single ion that is normally encountered, and you usually name them by just the metal name. You would not be wrong, however, if, for example, you named Zn^{2+} as zinc(II) ion.
3. The names of the monatomic anions are obtained from a stem name of the element followed by the suffix *-ide*. Example: Br^{-} is called bromide ion, from the stem name *brom-* for bromine and the suffix *-ide*.

TABLE 2.4 Common Cations of the Transition Elements

Ion	Ion Name	Ion	Ion Name	Ion	Ion Name
Cr^{3+}	Chromium(III) or chromic	Co^{2+}	Cobalt(II) or cobaltous	Zn^{2+}	Zinc
Mn^{2+}	Manganese(II) or manganous	Ni^{2+}	Nickel(II) or nickel	Ag^{+}	Silver
Fe^{2+}	Iron(II) or ferrous	Cu^{+}	Copper(I) or cuprous	Cd^{2+}	Cadmium
Fe^{3+}	Iron(III) or ferric	Cu^{2+}	Copper(II) or cupric	Hg^{2+}	Mercury(II) or mercuric

TABLE 2.5 Some Common Polyatomic Ions

Name	Formula	Name	Formula
Mercury(I) or mercurous	Hg_2^{2+}	Permanganate	MnO_4^-
Ammonium	NH_4^+	Nitrite	NO_2^-
Cyanide	CN^-	Nitrate	NO_3^-
Carbonate	CO_3^{2-}	Hydroxide	OH^-
Hydrogen carbonate (or bicarbonate)	HCO_3^-	Peroxide	O_2^{2-}
Acetate	$C_2H_3O_2^-$	Phosphate	PO_4^{3-}
Oxalate	$C_2O_4^{2-}$	Monohydrogen phosphate	HPO_4^{2-}
Hypochlorite	ClO^-	Dihydrogen phosphate	$H_2PO_4^-$
Chlorite	ClO_2^-	Sulfite	SO_3^{2-}
Chlorate	ClO_3^-	Sulfate	SO_4^{2-}
Perchlorate	ClO_4^-	Hydrogen sulfite (or bisulfite)	HSO_3^-
Chromate	CrO_4^{2-}	Hydrogen sulfate (or bisulfate)	HSO_4^-
Dichromate	$Cr_2O_7^{2-}$	Thiosulfate	$S_2O_3^{2-}$



SO_4^{2-}

SO_3^{2-}

FIGURE 2.25

Sulfate and sulfite oxoanions

Molecular models of the sulfate *(top)* and sulfite *(bottom)* oxoanions.

A **polyatomic ion** is *an ion consisting of two or more atoms chemically bonded together and carrying a net electric charge.* Table 2.5 lists some common polyatomic ions. The first two are cations (Hg_2^{2+} and NH_4^+); the rest are anions. There are no simple rules for writing the formulas of such ions, although you will find it helpful to note a few points.

Most of the ions in Table 2.5 are *oxoanions* (also called *oxyanions*), which consist of oxygen with another element (called the characteristic or central element). Sulfur, for example, forms the oxoanions sulfate ion, SO_4^{2-}, and sulfite ion, SO_3^{2-} (Figure 2.25). The oxoanions in the table are grouped by the characteristic element. For instance, the sulfur ions occur as a group at the end of the table.

Note that the names of the oxoanions have a stem name from the characteristic element, plus a suffix *-ate* or *-ite.* These suffixes denote the relative number of oxygen atoms in the oxoanions of a given characteristic element. The name of the oxoanion with the greater number of oxygen atoms has the suffix *-ate;* the name of the oxoanion with the lesser number of oxygen atoms has the suffix *-ite.* The sulfite ion, SO_3^{2-}, and sulfate ion, SO_4^{2-}, are examples. Another example are the two oxoanions of nitrogen listed in Table 2.5:

NO_2^-	nitr<u>ite</u> ion
NO_3^-	nitr<u>ate</u> ion

Unfortunately, the suffixes do not tell you the actual number of oxygen atoms in the oxoanion, only the relative number. However, if you were given the formulas of the two anions, you could name them.

The two suffixes *-ite* and *-ate* are not enough when there are more than two oxoanions of a given characteristic element. The oxoanions of chlorine are an example. Table 2.5 lists four oxoanions of chlorine: ClO^-, ClO_2^-, ClO_3^-, and ClO_4^-. In such cases, two prefixes *hypo-* and *per-* are used in addition to the two suffixes. The two oxoanions with the least number of oxygen atoms (ClO^- and ClO_2^-) are named using the suffix *-ite,* and the prefix *hypo-* is added to the one of these two ions with the fewer oxygen atoms:

ClO^-	<u>hypo</u>chlor<u>ite</u> ion
ClO_2^-	chlor<u>ite</u> ion

The two oxoanions with the greatest number of oxygen atoms (ClO_3^- and ClO_4^-) are named using the suffix *-ate,* and the prefix *per-* is added to the one of these two ions with the greater number of oxygen atoms:

ClO_3^-	chlor<u>ate</u> ion
ClO_4^-	<u>per</u>chlor<u>ate</u> ion

Some of the polyatomic ions in Table 2.5 are oxoanions bonded to one or more hydrogen ions (H^+). They are sometimes referred to as *acid anions,* because acids are substances that provide H^+ ions. As an example, monohydrogen phosphate ion, HPO_4^{2-}, is essentially a phosphate ion (PO_4^{3-}) to which a hydrogen ion (H^+) has bonded. The prefix *mono-,* from the Greek, means "one." Similarly, *di-* is a prefix from the Greek meaning "two," so dihydrogen phosphate ion is a phosphate ion to which two hydrogen ions have bonded. In an older terminology, ions such as hydrogen carbonate and hydrogen sulfate were called bicarbonate and bisulfate, respectively.

The last anion in the table is thiosulfate ion, $S_2O_3^{2-}$. The prefix *thio-* means that an oxygen atom in the root ion name (sulfate, SO_4^{2-}) has been replaced by a sulfur atom.

There are only a few common polyatomic cations. Those listed in Table 2.5 are the mercury(I) ion (also called mercurous ion) and the ammonium ion. Mercury(I) ion is one of the few common metal ions that is not monatomic; its formula is Hg_2^{2+}. The ion charge indicated in parentheses is the charge per metal atom. The ammonium ion, NH_4^+, is one of the few common cations composed of only nonmetal atoms.

Now that we have discussed the naming of ions, let us look at the naming of ionic compounds. Example 2.4 illustrates how you name an ionic compound given its formula.

Example 2.4 Naming an Ionic Compound from Its Formula

Name the following: a. Mg_3N_2, b. $CrSO_4$.

Problem Strategy First, you need to identify the type of compound that you are dealing with. Because each of these compounds contains both metal and nonmetal atoms, we expect them to be ionic, so we need to apply the rules for naming ionic compounds. When naming ionic compounds, start by writing the formulas of the cations and anions in the compound, then name the ions. Look first for any monatomic cations and anions whose charges are predictable. Use the rules given in the text to name the ions. For those metals having more than one cation, you will need to deduce the charge on the metal ion (say the ion of Cr in $CrSO_4$) from the charge on the anion. You will need to memorize or have available the formulas and names of the polyatomic anions to be able to write them from the formula of the compound.

Solution a. Magnesium, a Group IIA metal, is expected to form only a 2+ ion (Mg^{2+}, the magnesium ion). Nitrogen (Group VA) is expected to form an anion of charge equal to the group number minus 8 (N^{3-}, the nitride ion). You can check that these ions would give the formula Mg_3N_2. The name of the compound is **magnesium nitride** (the name of the cation followed by the name of the anion).

b. Chromium is a transition element and, like most such elements, has more than one monatomic ion. You can find the charge on the Cr ion if you know the formula of the anion. From Table 2.5, you see that the SO_4 in $CrSO_4$ refers to the anion SO_4^{2-} (the sulfate ion). Therefore, the Cr cation must be Cr^{2+} to give electrical neutrality. The name of Cr^{2+} is chromium(II) ion, so the name of the compound is **chromium(II) sulfate.**

Answer Check Whenever you have to name a compound, you can check your answer to see if the name will lead you back to the correct formula.

Exercise 2.5 Write the names of the following compounds: a. CaO, b. $PbCrO_4$.

■ See Problems 2.77 and 2.78.

Example 2.5 Writing the Formula from the Name of an Ionic Compound

Write formulas for the following compounds: a. iron(II) phosphate, b. titanium(IV) oxide.

Problem Strategy As in all nomenclature problems, first determine the type of compound you are working with. The compounds presented in this problem contain both metal and nonmetal atoms, so they are ionic. Use the name of the compound, obtain the name of the ions, and then write the formulas of the ions. Finally, use the method of Example 2.3 to obtain the formula of the compound from the formulas of the ions.

Solution a. Iron(II) phosphate contains the iron(II) ion, Fe^{2+}, and the phosphate ion, $PO_4^{\ 3-}$. (You will need to have memorized the formula of the phosphate ion, or have access to Table 2.5.) Once you know the formulas of the ions, you use the method of Example 2.3 to obtain the formula. The formula is $\mathbf{Fe_3(PO_4)_2}$.

b. Titanium(IV) oxide is composed of titanium(IV) ions, Ti^{4+}, and oxide ions, O^{2-}. The formula of titanium(IV) oxide is $\mathbf{TiO_2}$.

Answer Check Whenever you have to write the formula of a compound from its name, make sure that the formula you have written will result in the correct name.

Exercise 2.6 A compound has the name thallium(III) nitrate. What is its formula? The symbol of thallium is Tl.

■ See Problems 2.79 and 2.80.

Binary Molecular Compounds

A **binary compound** is *a compound composed of only two elements.* Binary compounds composed of a metal and a nonmetal are usually ionic and are named as ionic compounds, as we have just discussed. (For example, NaCl, $MgBr_2$, and Al_2N_3 are all binary ionic compounds.) Binary compounds composed of two nonmetals or metalloids are usually molecular and are named using a prefix system. Examples of binary molecular compounds are H_2O, NH_3, and CCl_4. Using this prefix system, you name the two elements using the order given by the formula of the compound.

Order of Elements in the Formula The order of elements in the formula of a binary molecular compound is established by convention. By this convention, the nonmetal or metalloid occurring first in the following sequence is written first in the formula of the compound.

$$\begin{array}{ll} \text{Element} & \underbrace{\text{B}}_{\text{IIIA}}\ \underbrace{\text{Si C}}_{\text{IVA}}\ \underbrace{\text{Sb As P N}}_{\text{VA}}\ \text{H}\ \underbrace{\text{Te Se S}}_{\text{VIA}}\ \underbrace{\text{I Br Cl}}_{\text{VIIA}}\ \text{O F} \\ \text{Group} & \end{array}$$

You can reproduce this order easily if you have a periodic table available. You arrange the nonmetals and metalloids in the order of the groups of the periodic table, listing the elements from the bottom of the group upward. Then you place H between Groups VA and VIA, and move O so that it is just before F.

This order places the nonmetals and metalloids approximately in order of increasing nonmetallic character. Thus, in the formula, the first element is the more metallic and the second element is the more nonmetallic, similar to the situation with binary ionic compounds. For example, the compound whose molecule contains three fluor ine atoms and one nitrogen atom is written NF_3, not F_3N.

Rules for Naming Binary Molecular Compounds Now let us look at the rules for naming binary molecular compounds by the *prefix system.*

1. The name of the compound usually has the elements in the order given in the formula.

TABLE 2.6

Greek Prefixes for Naming Compounds

Number	Prefix
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

2. You name the first element using the exact element name.
3. You name the second element by writing the stem name of the element with the suffix *-ide* (as if the element occurred as the anion).
4. You add a prefix, derived from the Greek, to each element name to denote the subscript of the element in the formula. (The Greek prefixes are listed in Table 2.6.) Generally, the prefix *mono-* is not used, unless it is needed to distinguish two compounds of the same two elements.

Consider N_2O_3. This is a binary molecular compound, as you would predict because N and O are nonmetals. According to Rule 1, you name it after the elements, N before O, following the order of elements in the formula. By Rule 2, the N is named exactly as the element (nitrogen), and by Rule 3, the O is named as the anion (oxide). The compound is a nitrogen oxide. Finally, you add prefixes to denote the subscripts in the formula (Rule 4). As Table 2.6 shows, the prefix for two is *di-*, and the prefix for three is *tri-*. The name of the compound is <u>di</u>nitrogen <u>tri</u>oxide.

Here are some examples to illustrate how the prefix *mono-* is used. There is only one compound of hydrogen and chlorine: HCl. It is called hydrogen chloride, not monohydrogen monochloride, since the prefix *mono-* is not generally written. On the other hand, there are two common compounds of carbon and oxygen: CO and CO_2. They are both carbon oxides. To distinguish one from the other, we name them carbon monoxide and carbon dioxide, respectively.

Here are some other examples of prefix names for binary molecular compounds.

SF_4	sulfur tetrafluoride	SF_6	sulfur hexafluoride
ClO_2	chlorine dioxide	Cl_2O_7	dichlorine hept(a)oxide

The final vowel in a prefix is often dropped before a vowel in a stem name, for ease in pronunciation. Tetraoxide, for instance, becomes tetroxide.

A few compounds have older, well-established names. The compound H_2S would be named dihydrogen sulfide by the prefix system, but is commonly called hydrogen sulfide. NO is still called nitric oxide, although its prefix name is nitrogen monoxide. The names water and ammonia are used for H_2O and NH_3, respectively.

Example 2.6 Naming a Binary Compound from Its Formula

Name the following compounds: a. N_2O_4, b. P_4O_6.

Problem Strategy First, you need to determine the type of compound that you are dealing with: molecular or ionic. In this case the compound contains only nonmetals, so it is molecular. Molecular compounds require that you use the Greek prefixes (Table 2.6), naming the elements in the same order as in the formulas.

Solution a. N_2O_4 contains two nitrogen atoms and four oxygen atoms. Consulting Table 2.6 reveals that the corresponding prefixes are *di-* and *tetra-*, so the name is **<u>*di*</u>nitrogen <u>*tetr*</u>oxide.**

b. P_4O_6 contains four phosphorus atoms and six oxygen atoms. Consulting Table 2.6 reveals that the corresponding prefixes are *tetra-* and *hexa-*, so the name is **<u>*tetra*</u>phosphorus <u>*hex*</u>oxide.**

Answer Check Check to see whether the name leads back to the correct formula.

Exercise 2.7 Name the following compounds: a. Cl_2O_6, b. PCl_3, c. PCl_5.

■ See Problems 2.83 and 2.84.

Example 2.7 Writing the Formula from the Name of a Binary Compound

Give the formulas of the following compounds: a. disulfur dichloride, b. tetraphosphorus trisulfide.

Problem Strategy As always, first determine whether you are working with a molecular or ionic compound. Because these compounds contain only nonmetals, they must be molecular. Given that they are molecular, you will need to use the Greek prefixes in Table 2.6 to give you the corresponding names for each subscript.

Solution You change the names of the elements to symbols and translate the prefixes to subscripts. The formulas are a. $\mathbf{S_2Cl_2}$ and b. $\mathbf{P_4S_3}$.

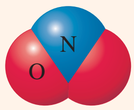
Answer Check Check to see whether the formula you have written will result in the correct name.

Exercise 2.8 Give formulas for the following compounds: a. carbon disulfide, b. sulfur trioxide.

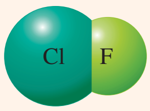
■ See Problems 2.85 and 2.86.

Example 2.8 Naming a Binary Chemical Compound from Its Molecular Model

Name the chemical compounds shown here.



(a)



(b)

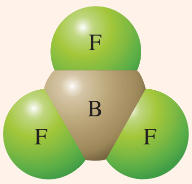
Problem Strategy After recognizing that these are molecular compounds, start by writing down the elements using subscripts to denote how many atoms of each element occur in the compound. Next, determine which element should be written first in the chemical formula. (The more metallic element is written first.)

Solution a. The elements in the compound are O and N. There are two O atoms for each N atom, so the chemical formula could be O_2N. Applying the rule for writing the more metallic element first, the formula is rearranged to yield the correct formula, NO_2. Using Greek prefixes and naming the elements in the order in which they appear, we get the correct chemical formula: **nitrogen dioxide.**

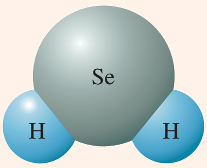
b. Following the same procedure as part a, we get **chlorine monofluoride.**

Answer Check Make sure that the answer reflects the correct number of each of the elements given in the molecular structure.

Exercise 2.9
Using the molecular models, name the following chemical compounds:



(a)



(b)

■ See Problems 2.87 and 2.88.

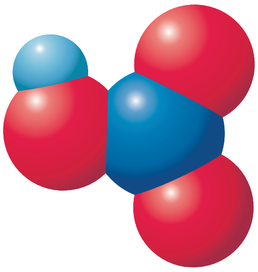


FIGURE 2.26 ▲
Molecular model of nitric acid

Acids and Corresponding Anions

Acids are an important class of compounds, and we will look closely at them in the next chapter. Here we want merely to see how we name these compounds and how they are related to the anions that we encounter in some ionic compounds. For our present purposes, an *acid* is a molecular compound that yields hydrogen ions, H^+, and an anion for each acid molecule when the acid dissolves in water. An example is nitric acid, HNO_3. The HNO_3 molecule yields one H^+ ion and one nitrate ion, NO_3^-, in aqueous (water) solution.

Nitric acid, HNO_3, is an oxoacid, or oxyacid (Figure 2.26). An **oxoacid** is *an acid containing hydrogen, oxygen, and another element (often called the central atom).* In water the oxoacid molecule yields one or more hydrogen ions, H^+, and an oxoanion. The names of the oxoacids are related to the names of the corresponding oxoanions. If you know the name of the oxoanion, you can obtain the name of the corresponding acid by replacing the suffix as follows:

Anion Suffix	*Acid Suffix*
-ate	*-ic*
-ite	*-ous*

The following diagram illustrates how to apply this information to naming oxoacids.

Acid	Contains	Name
HNO_3 →	nitrate anion therefore	nitric acid
	ate to ic	
HNO_2 →	nitrite anion therefore	nitrous acid
	ite to ous	

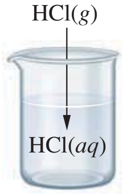
Table 2.7 lists some oxoanions and their corresponding oxoacids.

Some binary compounds of hydrogen and nonmetals yield acidic solutions when dissolved in water. These *solutions* are named like compounds by using the prefix *hydro-* and the suffix *-ic* with the stem name of the nonmetal, followed by the word *acid*. We denote the solution by the formula of the binary compound followed by (*aq*)

TABLE 2.7 Some Oxoanions and Their Corresponding Oxoacids

Oxoanion		Oxoacid	
CO_3^{2-}	Carbon*ate ion*	H_2CO_3	Carbon*ic acid*
NO_2^-	Nitr*ite ion*	HNO_2	Nitr*ous acid*
NO_3^-	Nitr*ate ion*	HNO_3	Nitr*ic acid*
PO_4^{3-}	Phosph*ate ion*	H_3PO_4	Phosphor*ic acid*
SO_3^{2-}	Sulf*ite ion*	H_2SO_3	Sulfur*ous acid*
SO_4^{2-}	Sulf*ate ion*	H_2SO_4	Sulfur*ic acid*
ClO^-	*Hypo*chlor*ite ion*	$HClO$	*Hypo*chlor*ous acid*
ClO_2^-	Chlor*ite ion*	$HClO_2$	Chlor*ous acid*
ClO_3^-	Chlor*ate ion*	$HClO_3$	Chlor*ic acid*
ClO_4^-	*Per*chlor*ate ion*	$HClO_4$	*Per*chlor*ic acid*

for aqueous (water) solution. The corresponding binary compound can be distinguished from the solution by appending the state of the compound to the formula. Thus, when hydrogen chloride gas, HCl(*g*), is dissolved in water, it forms hydrochloric acid, HCl(*aq*).



Here are some other examples:

Binary Compound	*Acid Solution*
HBr(*g*), hydrogen bromide	*hydro*bromic *acid,* HBr(*aq*)
HF(*g*), hydrogen fluoride	*hydro*fluoric *acid,* HF(*aq*)

Example 2.9 Writing the Name and Formula of an Anion from the Acid

Selenium has an oxoacid, H_2SeO_4, called selenic acid. What is the formula and name of the corresponding anion?

Problem Strategy Knowing that H_2SeO_4 is an oxyacid, we must first determine the oxoanion name from the examples provided in Table 2.7. From the oxoanion name you determine whether a prefix (*per-* or *hypo-*) is necessary and whether the proper suffix is *-ite* or *-ate*. You then name the anion.

Solution When you remove two H^+ ions from H_2SeO_4, you obtain the $\mathbf{SeO_4^{2-}}$ ion. You name the ion from the acid by replacing *-ic* with *-ate*. The anion is called the **selenate ion.**

Answer Check Check to see whether the name you have written leads back to the correct formula.

Exercise 2.10 What are the name and formula of the anion corresponding to perbromic acid, $HBrO_4$?

■ See Problems 2.89 and 2.90.

Hydrates

A **hydrate** is *a compound that contains water molecules weakly bound in its crystals.* These substances are often obtained by evaporating an aqueous solution of the compound. Consider copper(II) sulfate. When an aqueous solution of this substance is evaporated, blue crystals form in which each formula unit of copper(II) sulfate, $CuSO_4$, is associated with five molecules of water. The formula of the hydrate is written $CuSO_4{\cdot}5H_2O$, where a centered dot separates $CuSO_4$ and $5H_2O$. When the blue crystals of the hydrate are heated, the water is driven off, leaving behind white crystals of copper(II) sulfate without associated water, a substance called *anhydrous* copper(II) sulfate (see Figure 2.27).

Hydrates are named from the anhydrous compound, followed by the word *hydrate* with a prefix to indicate the number of water molecules per formula unit of the compound. For example, $CuSO_4{\cdot}5H_2O$ is known as copper(II) sulfate pentahydrate.

FIGURE 2.27 ▲
Copper(II) sulfate
The hydrate $CuSO_4{\cdot}5H_2O$ is blue; the anhydrous compound, $CuSO_4$, is white.

Example 2.10 Naming a Hydrate from Its Formula

Epsom salts has the formula $MgSO_4 \cdot 7H_2O$. What is the chemical name of the substance?

Problem Strategy You need to identify the type of compound that you are dealing with, ionic or molecular. Because this compound contains both metals and nonmetals, it is ionic. However, this compound has the additional feature that it contains water molecules. Therefore, you need to apply ionic naming rules and then use Greek prefixes (Table 2.6) to indicate the number of water molecules that are part of the formula. The word *hydrate* will be used to represent water in the name.

Solution $MgSO_4$ is magnesium sulfate. $MgSO_4 \cdot 7H_2O$ is **magnesium sulfate heptahydrate.**

Answer Check Make sure that the name leads back to the formula that you started with.

Exercise 2.11 Washing soda has the formula $Na_2CO_3 \cdot 10H_2O$. What is the chemical name of this substance?

■ See Problems 2.91 and 2.92.

Example 2.11 Writing the Formula from the Name of a Hydrate

The mineral gypsum has the chemical name calcium sulfate dihydrate. What is the chemical formula of this substance?

Problem Strategy You need to identify the type of compound that you are dealing with, ionic or molecular. Because this compound contains both metals and nonmetals, it is ionic. However, this compound has the additional feature that it contains water molecules. Therefore, we need to apply ionic naming rules and then use Greek prefixes (Table 2.6) to indicate the number of water molecules that are part of the formula. The word *hydrate* will be used to represent water in the name.

Solution Calcium sulfate is composed of calcium ions (Ca^{2+}) and sulfate ions (SO_4^{2-}), so the formula of the anhydrous compound is $CaSO_4$. Since the mineral is a *di*hydrate, the formula of the compound is $\mathbf{CaSO_4 \cdot 2H_2O}$.

Answer Check Make sure that the formula will result in the correct name that you started with.

Exercise 2.12 Photographers' hypo, used to fix negatives during the development process, is sodium thiosulfate pentahydrate. What is the chemical formula of this compound?

■ See Problems 2.93 and 2.94.

Concept Check 2.5

You take a job with the U.S. Environmental Protection Agency (EPA) inspecting college chemistry laboratories. On the first day of the job while inspecting Generic University you encounter a bottle with the formula Al_2Q_3 that was used as an unknown compound in an experiment. Before you send the compound off to the EPA lab for analysis, you want to narrow down the possibilities for element Q. What are the likely (real element) candidates for element Q?

Chemical Reactions: Equations

An important feature of atomic theory is its explanation of a chemical reaction as a rearrangement of the atoms of substances. The reaction of sodium metal with chlorine gas described in the chapter opening involves the rearrangement of the atoms of sodium and chlorine to give the new combination of atoms in sodium chloride. Such a rearrangement of atoms is conveniently represented by a chemical equation, which uses chemical formulas.

2.9 Writing Chemical Equations

A **chemical equation** is *the symbolic representation of a chemical reaction in terms of chemical formulas.* For example, the burning of sodium in chlorine to produce sodium chloride is written

$$2\text{Na} + \text{Cl}_2 \longrightarrow 2\text{NaCl}$$

The formulas on the left side of an equation (before the arrow) represent the reactants; a **reactant** is *a starting substance in a chemical reaction.* The arrow means "react to form" or "yield." The formulas on the right side represent the products; a **product** is *a substance that results from a reaction.* Note the *coefficient* of 2 in front of the formula NaCl. Coefficients in front of the formula give the relative number of molecules or formula units involved in the reaction. When no coefficient is written, it is understood to be 1.

In many cases, it is useful to indicate the states or phases of the substances in an equation. You do this by placing appropriate labels indicating the phases within parentheses following the formulas of the substances. You use the following phase labels:

$$(g) = \text{gas},\ (l) = \text{liquid},\ (s) = \text{solid},\ (aq) = \text{aqueous (water) solution}$$

When you use these labels, the previous equation becomes

$$2\text{Na}(s) + \text{Cl}_2(g) \longrightarrow 2\text{NaCl}(s)$$

You can also indicate in an equation the conditions under which a reaction takes place. If the reactants are heated to make the reaction go, you can indicate this by putting the symbol Δ (capital Greek delta) over the arrow. For example, the equation

$$2\text{NaNO}_3(s) \xrightarrow{\Delta} 2\text{NaNO}_2(s) + \text{O}_2(g)$$

indicates that solid sodium nitrate, $NaNO_3$, decomposes when heated to give solid sodium nitrite, $NaNO_2$, and oxygen gas, O_2.

When an aqueous solution of hydrogen peroxide, H_2O_2, comes in contact with platinum metal, Pt, the hydrogen peroxide decomposes into water and oxygen gas. The platinum acts as a *catalyst,* a substance that speeds up a reaction without undergoing any net change itself. You write the equation for this reaction as follows. The catalyst, Pt, is written over the arrow. <

Platinum is a silvery-white metal used for jewelry. It is also valuable as a catalyst for many reactions, including those that occur in the catalytic converters of automobiles.

$$2\text{H}_2\text{O}_2(aq) \xrightarrow{\text{Pt}} 2\text{H}_2\text{O}(l) + \text{O}_2(g)$$

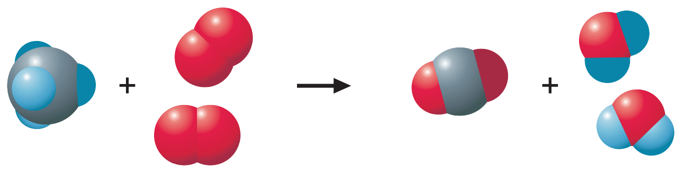
You also can combine the symbol Δ with the formula of a catalyst, placing one above the arrow and the other below the arrow.

2.10 Balancing Chemical Equations

When the coefficients in a chemical equation are correctly given, the numbers of atoms of each element are equal on both sides of the arrow. The equation is then said to be *balanced.* That a chemical equation should be balanced follows from atomic theory. A chemical reaction involves simply a recombination of the atoms; none are destroyed and none are created. Consider the burning of natural gas, which is composed mostly of methane, CH_4. Using atomic theory, you describe this as the chemical reaction of one molecule of methane, CH_4, with two molecules of oxygen, O_2, to form one molecule of carbon dioxide, CO_2, and two molecules of water, H_2O.

$$\underset{\substack{\text{one molecule}\\ \text{of methane}}}{\text{CH}_4} \underset{+}{+} \underset{\substack{\text{two molecules}\\ \text{of oxygen}}}{2\text{O}_2} \underset{\substack{\text{react to}\\ \text{form}}}{\longrightarrow} \underset{\substack{\text{one molecule}\\ \text{of carbon dioxide}}}{\text{CO}_2} \underset{+}{+} \underset{\substack{\text{two molecules}\\ \text{of water}}}{2\text{H}_2\text{O}}$$

FIGURE 2.28 ▶
Representation of the reaction of methane with oxygen
Molecular models represent the reaction of CH_4 with O_2 to give CO_2 and H_2O.



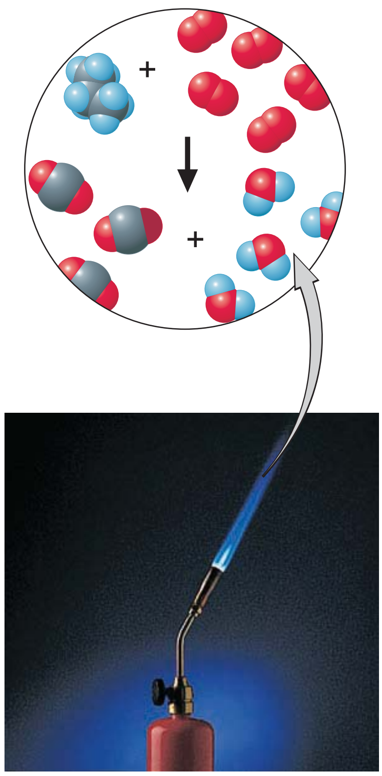


FIGURE 2.29 ▲
The burning of propane gas
Propane gas, C_3H_8, from the tank burns by reacting with oxygen, O_2, in air to give carbon dioxide, CO_2, and water, H_2O.

Figure 2.28 represents the reaction in terms of molecular models.

Before you can write a balanced chemical equation for a reaction, you must determine *by experiment* those substances that are reactants and those that are products. You must also determine the formulas of each substance. Once you know these things, you can write the balanced chemical equation.

As an example, consider the burning of propane gas (Figure 2.29). By experiment, you determine that propane reacts with oxygen in air to give carbon dioxide and water. You also determine from experiment that the formulas of propane, oxygen, carbon dioxide, and water are C_3H_8, O_2, CO_2, and H_2O, respectively. Then you can write

$$C_3H_8 + O_2 \longrightarrow CO_2 + H_2O$$

This equation is not balanced because the coefficients that give the relative number of molecules involved have not yet been determined. *To balance the equation, you select coefficients that will make the numbers of atoms of each element equal on both sides of the equation.*

Because there are three carbon atoms on the left side of the equation (C_3H_8), you must have three carbon atoms on the right. This is achieved by writing 3 for the coefficient of CO_2.

$$\underline{1}C_3H_8 + O_2 \longrightarrow \underline{3}CO_2 + H_2O$$

(We have written the coefficient for C_3H_8 to show that it is now determined; normally when the coefficient is 1, it is omitted.) Similarly, there are eight hydrogen atoms on the left (in C_3H_8), so you write 4 for the coefficient of H_2O.

$$\underline{1}C_3H_8 + O_2 \longrightarrow \underline{3}CO_2 + \underline{4}H_2O$$

The coefficients on the right are now determined, and there are ten oxygen atoms on this side of the equation: 6 from the three CO_2 molecules and 4 from the four H_2O molecules. You now write 5 for the coefficient of O_2. We drop the coefficient 1 from C_3H_8 and have the balanced equation.

$$C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$$

It is a good idea to check your work by counting the number of atoms of each element on the left side and then on the right side of the equation.

The combustion of propane can be equally well represented by

$$2C_3H_8 + 10O_2 \longrightarrow 6CO_2 + 8H_2O$$

in which the coefficients of the previous equation have been doubled. Usually, however, it is preferable to *write the coefficients so that they are the smallest whole numbers possible.* When we refer to a balanced equation, we will assume that this is the case unless otherwise specified.

The method we have outlined—called *balancing by inspection*—is essentially a trial-and-error method. < It can be made easier, however, by observing the following rule:

Balancing equations this way is relatively quick for all but the most complicated equations. A special method for such equations is described in Chapter 19.

> Balance first the atoms for elements that occur in only one substance on each side of the equation.

The usefulness of this rule is illustrated in the next example. Remember that in any formula, such as $Fe_2(SO_4)_3$, a subscript to the right of the parentheses multiplies *each* subscript within the parentheses. Thus, $Fe_2(SO_4)_3$ represents 4×3 oxygen atoms. Remember also that you cannot change a subscript in any formula; only the coefficients can be altered to balance an equation.

Example 2.12 Balancing Simple Equations

Balance the following equations.

a. $H_3PO_3 \longrightarrow H_3PO_4 + PH_3$ b. $Ca + H_2O \longrightarrow Ca(OH)_2 + H_2$

c. $Fe_2(SO_4)_3 + NH_3 + H_2O \longrightarrow Fe(OH)_3 + (NH_4)_2SO_4$

Problem Strategy Look at the chemical equation for atoms of elements that occur in only one substance on each side of the equation. Begin by balancing the equation in one of these atoms. Use the number of atoms on the left side of the arrow as the coefficient of the substance containing that element on the right side, and vice versa. After balancing one element in an equation, the rest become easier.

Solution a. Oxygen occurs in just one of the products (H_3PO_4). It is therefore easiest to balance O atoms first. To do this, note that H_3PO_3 has three O atoms; use 3 as the coefficient of H_3PO_4 on the right side. There are four O atoms in H_3PO_4 on the right side; use 4 as the coefficient of H_3PO_3 on the left side.

$$\underline{4}H_3PO_3 \longrightarrow \underline{3}H_3PO_4 + PH_3$$

This equation is now also balanced in P and H atoms. Thus, the balanced equation is

$$\mathbf{4H_3PO_3 \longrightarrow 3H_3PO_4 + PH_3}$$

b. The equation is balanced in Ca atoms as it stands.

$$\underline{1}Ca + H_2O \longrightarrow \underline{1}Ca(OH)_2 + H_2$$

O atoms occur in only one reactant and in only one product, so they are balanced next.

$$\underline{1}Ca + \underline{2}H_2O \longrightarrow \underline{1}Ca(OH)_2 + H_2$$

The equation is now also balanced in H atoms. Thus, the answer is

$$\mathbf{Ca + 2H_2O \longrightarrow Ca(OH)_2 + H_2}$$

c. To balance the equation in Fe atoms, you write

$$\underline{1}Fe_2(SO_4)_3 + NH_3 + H_2O \longrightarrow \underline{2}Fe(OH)_3 + (NH_4)_2SO_4$$

You balance the S atoms by placing the coefficient 3 for $(NH_4)_2SO_4$.

$$\underline{1}Fe_2(SO_4)_3 + \underline{6}NH_3 + H_2O \longrightarrow \underline{2}Fe(OH)_3 + \underline{3}(NH_4)_2SO_4$$

Now you balance the N atoms.

$$\underline{1}Fe_2(SO_4)_3 + \underline{6}NH_3 + H_2O \longrightarrow \underline{2}Fe(OH)_3 + \underline{3}(NH_4)_2SO_4$$

To balance the O atoms, you first count the number of O atoms on the right (18). Then you count the number of O atoms in substances on the left with known coefficients. There are 12 O's in $Fe_2(SO_4)_3$; hence, the number of remaining O's (in H_2O) must be $18 - 12 = 6$.

$$\underline{1}Fe_2(SO_4)_3 + \underline{6}NH_3 + \underline{6}H_2O \longrightarrow \underline{2}Fe(OH)_3 + \underline{3}(NH_4)_2SO_4$$

(continued)

(continued)

Finally, note that the equation is now balanced in H atoms. The answer is

$$\mathbf{Fe_2(SO_4)_3 + 6NH_3 + 6H_2O \longrightarrow 2Fe(OH)_3 + 3(NH_4)_2SO_4}$$

Answer Check As a final step when balancing chemical equations, always count the number of atoms of each element on both sides of the equation to make certain that they are equal.

Exercise 2.13 Find the coefficients that balance the following equations.

a. $O_2 + PCl_3 \longrightarrow POCl_3$

b. $P_4 + N_2O \longrightarrow P_4O_6 + N_2$

c. $As_2S_3 + O_2 \longrightarrow As_2O_3 + SO_2$

d. $Ca_3(PO_4)_2 + H_3PO_4 \longrightarrow Ca(H_2PO_4)_2$

■ See Problems 2.97 and 2.98.

A Checklist for Review

Important Terms

atomic theory (2.1)
atom (2.1)
element (2.1, 2.3)
compound (2.1)
chemical reaction (2.1)
atomic symbol (2.1)
law of multiple proportions (2.1)
nucleus (2.2)
electron (2.2)
proton (2.3)
atomic number (*Z*) (2.3)
neutron (2.3)
mass number (*A*) (2.3)
nuclide (2.3)
isotope (2.3)
atomic mass unit (amu) (2.4)
atomic mass (2.4)
fractional abundance (2.4)
periodic table (2.5)
period (of periodic table) (2.5)
group (of periodic table) (2.5)
metal (2.5)
nonmetal (2.5)
metalloid (semimetal) (2.5)
chemical formula (2.6)
molecule (2.6)
molecular formula (2.6)
polymer (2.6)
monomer (2.6)
ion (2.6)
anion (2.6)
cation (2.6)
ionic compound (2.6)
formula unit (2.6)
organic compound (2.7)
hydrocarbon (2.7)
functional group (2.7)
chemical nomenclature (2.8)
inorganic compound (2.8)
monatomic ion (2.8)
polyatomic ion (2.8)
binary compound (2.8)
oxoacid (2.8)
hydrate (2.8)
chemical equation (2.9)
reactant (2.9)
product (2.9)

Summary of Facts and Concepts

Atomic theory is central to chemistry. According to this theory, all matter is composed of small particles, or *atoms*. Each *element* is composed of the same kind of atom, and a *compound* is composed of two or more elements chemically combined in fixed proportions. A *chemical reaction* consists of the rearrangement of the atoms present in the reacting substances to give new chemical combinations present in the substances formed by the reaction. Although Dalton considered atoms to be the ultimate particles of matter, we now know that atoms themselves have structure. An atom has a *nucleus* and *electrons*. These atomic particles were discovered by J. J. Thomson and Ernest Rutherford.

Thomson established that cathode rays consist of negatively charged particles, or electrons, that are constituents of all atoms. Thomson measured the mass-to-charge ratio of the electron, and later Millikan measured its charge. From these measurements, the electron was found to be more than 1800 times lighter than the lightest atom.

Rutherford proposed the nuclear model of the atom to account for the results of experiments in which alpha particles were scattered from metal foils. According to this model, the atom consists of a central core, or nucleus, around which the electrons exist. The nucleus has most of the mass of the atom and consists of *protons* (with a positive charge) and *neutrons* (with no charge). Each chemically distinct atom has a nucleus with a specific number of protons (*atomic number*), and around the nucleus in the neutral atom are an equal number of electrons.

The number of protons plus neutrons in a nucleus equals the *mass number.* Atoms whose nuclei have the same number of protons but different number of neutrons are called *isotopes.*

The *atomic mass* can be calculated from the isotopic masses and *fractional abundances* of the isotopes in a naturally occurring element. These data can be determined by the use of a mass spectrometer, which separates ions according to their mass-to-charge ratios.

The elements can be arranged in rows and columns by atomic number to form the *periodic table.* Elements in a given *group* (column) have similar properties. (A *period* is a row in the periodic table.) Elements on the left and at the center of the table are *metals;* those on the right are *nonmetals.*

A *chemical formula* is a notation used to convey the relative proportions of the atoms of the different elements in a substance. If the substance is molecular, the formula gives the precise number of each kind of atom in the molecule. If the substance is ionic, the formula gives the relative number of different ions in the compound.

Chemical nomenclature is the systematic naming of compounds based on their formulas or structures. Rules are given for naming *ionic compounds, binary molecular compounds, acids,* and *hydrates.*

A chemical reaction occurs when the atoms in substances rearrange and combine into new substances. We represent a reaction by a *chemical equation,* writing a chemical formula for each reactant and product. The coefficients in the equation indicate the relative numbers of reactant and product molecules or formula units. Once the reactants and products and their formulas have been determined by experiment, we determine the coefficients by *balancing* the numbers of each kind of atom on both sides of the equation.

Media Summary

Visit the **student website at college.hmco.com/pic/ebbing9e** to help prepare for class, study for quizzes and exams, understand core concepts, and visualize molecular-level interactions. The following media activities are available for this chapter:

Prepare for Class

- **Video Lessons** *Mini lectures from chemistry experts*

  Early Discoveries and the Atom
  Understanding the Nucleus
  Understanding Electrons
  Modern Atomic Structure
  Mass Spectrometry: Determining Atomic Masses
  Creating the Periodic Table
  Describing Chemical Formulas
  Naming Chemical Compounds
  An Introduction to Chemical Reactions and Equations
  Balancing Chemical Equations

Improve Your Grade

- **Visualizations** *Molecular-level animations and lab demonstration videos*

  Cathode-Ray Tube
  Millikan's Oil Drop Experiment
  Gold Foil Experiment
  Determining Formulas for Ionic Compounds
  Conservation of Mass and Balancing Equations

- **Tutorials** *Animated examples and interactive activities*

  Isotopes
  Determining Formulas for Ionic Compounds
  Balancing Chemical Equations

- **Flashcards** *Key terms and definitions*

  Online Flashcards

- **Self-Assessment Questions** *Additional questions with full worked-out solutions*

  6 Self-Assessment Questions

ACE the Test

Multiple-choice quizzes
3 ACE Practice Tests

Access these resources using the passkey available free with new texts or for purchase separately.

Learning Objectives

2.1 Atomic Theory of Matter

- List the postulates of atomic theory.
- Define *element, compound,* and *chemical reaction* in the context of these postulates.
- Recognize the atomic symbols of the elements.
- Explain the significance of the law of multiple proportions.

2.2 The Structure of the Atom

- Describe Thomson's experiment in which he discovered the electron.
- Describe Rutherford's experiment that led to the nuclear model of the atom.

2.3 Nuclear Structure; Isotopes

- Name and describe the nuclear particles making up the nucleus of the atom.
- Define *atomic number, mass number,* and *nuclide.*
- Write the nuclide symbol for a given nuclide.
- Define and provide examples of *isotopes* of an element.
- Write the nuclide symbol of an element. Example 2.1

2.4 Atomic Masses

- Define *atomic mass unit* and *atomic mass.*
- Describe how a *mass spectrometer* can be used to determine the *fractional abundance* of the isotopes of an element.
- Determine the atomic mass of an element from the isotopic masses and fractional abundances. Example 2.2

2.5 Periodic Table of the Elements

- Identify periods and groups on the periodic table.
- Find the *main-group* and *transition* elements on the periodic table.
- Locate the *alkali metal* and *halogen* groups on the periodic table.
- Recognize the portions of the periodic table that contain the *metals, nonmetals,* and *metalloids* (*semimetals*).

2.6 Chemical Formulas; Molecular and Ionic Substances

- Determine when the *chemical formula* of a compound represents a *molecule.*
- Determine whether a chemical formula is also a *molecular formula.*
- Define *ion, cation,* and *anion.*
- Classify compounds as *ionic* or *molecular.*
- Define and provide examples for the term *formula unit.*
- Specify the charge on all substances, ionic and molecular.
- Write an ionic formula, given the ions. Example 2.3

2.7 Organic Compounds

- List the attributes of molecular substances that make them *organic compounds.*
- Explain what makes a molecule a *hydrocarbon.*
- Recognize some *functional groups* of organic molecules.

2.8 Naming Simple Compounds

- Recognize *inorganic compounds.*
- Learn the rules for predicting the charges of *monatomic ions* in ionic compounds.
- Apply the rules for naming monatomic ions.
- Learn the names and charges of common *polyatomic ions.*
- Name an ionic compound from its formula. Example 2.4
- Write the formula of an ionic compound from its name. Example 2.5
- Determine the order of elements in a *binary (molecular) compound.*
- Learn the rules for naming binary molecular compounds, including the Greek prefixes.
- Name a binary compound from its formula. Example 2.6
- Write the formula of a binary compound from its name. Example 2.7
- Name a binary molecular compound from its molecular model. Example 2.8
- Recognize molecular compounds that are acids.
- Determine whether an acid is an *oxoacid.*
- Learn the approach for naming binary acids and oxoacids.
- Write the name and formula of an anion from the acid. Example 2.9
- Recognize compounds that are *hydrates.*
- Learn the rules for naming hydrates.
- Name a hydrate from its formula. Example 2.10
- Write the formula of a hydrate from its name. Example 2.11

2.9 Writing Chemical Equations

- Identify the *reactants* and *products* in a chemical equation.
- Write chemical equations using appropriate phase labels, symbols of reaction conditions, and the presence of a catalyst.

2.10 Balancing Chemical Equations

- Determine if a chemical reaction is balanced.
- Master the techniques for balancing chemical equations. Example 2.12

Self-Assessment and Review Questions

2.1 Describe atomic theory and discuss how it explains the great variety of different substances. How does it explain chemical reactions?

2.2 Two compounds of iron and chlorine, A and B, contain 1.270 g and 1.904 g of chlorine, respectively, for each gram of iron. Show that these amounts are in the ratio 2:3. Is this consistent with the law of multiple proportions? Explain.

2.3 Explain the operation of a cathode-ray tube. Describe the deflection of cathode rays by electrically charged plates placed within the cathode-ray tube. What does this imply about cathode rays?

2.4 Explain Millikan's oil-drop experiment.

2.5 Describe the nuclear model of the atom. How does this model explain the results of alpha-particle scattering from metal foils?

2.6 What are the different kinds of particles in the atom's nucleus? Compare their properties with each other and with those of an electron.

2.7 Describe how protons and neutrons were discovered to be constituents of nuclei.

2.8 Oxygen consists of three different __________, each having eight protons but different numbers of neutrons.

2.9 Describe how Dalton obtained relative atomic masses.

2.10 Briefly explain how a mass spectrometer works. What kinds of information does one obtain from the instrument?

2.11 Define the term *atomic mass.* Why might the values of atomic masses on a planet elsewhere in the universe be different from those on earth?

2.12 What is the name of the element in Group IVA and Period 5?

2.13 Cite some properties that are characteristic of a metal.

2.14 Ethane consists of molecules with two atoms of carbon and six atoms of hydrogen. Write the molecular formula for ethane.

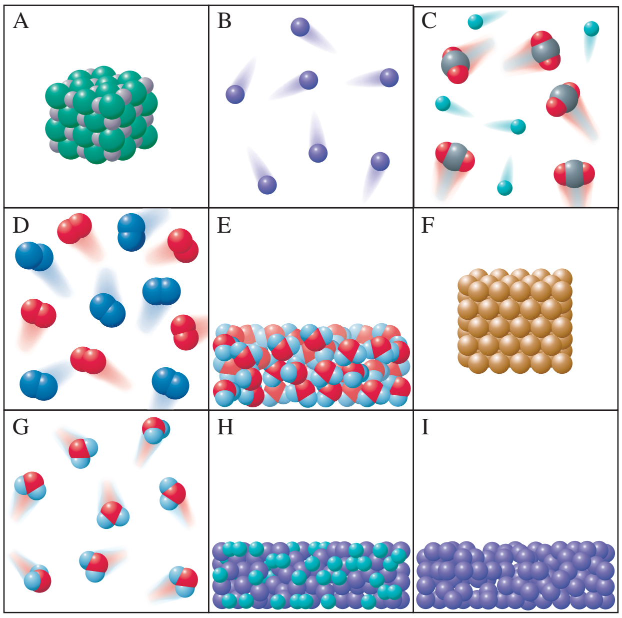
2.15 What is the difference between a molecular formula and a structural formula?

2.16 What is the fundamental difference between an organic substance and an inorganic substance? Write chemical formulas of three inorganic molecules that contain carbon.

2.17 Give an example of a binary compound that is ionic. Give an example of a binary compound that is molecular.

2.18 Which of the following models represent a(n):

- **a.** element
- **b.** compound
- **c.** mixture
- **d.** ionic solid
- **e.** gas made up of an element and a compound
- **f.** mixture of elements
- **g.** solid element
- **h.** solid
- **i.** liquid



2.19 The compounds CuCl and $CuCl_2$ were formerly called cuprous chloride and cupric chloride, respectively. What are their names using the Stock system of nomenclature? What are the advantages of the Stock system of nomenclature over the former one?

2.20 Explain what is meant by the term *balanced chemical equation.*

2.21 How many protons, neutrons, and electrons are in $^{119}Sn^{2+}$?

- **a.** 50 p, 69 n, 48 e^-
- **b.** 50 p, 69 n, 50 e^-
- **c.** 119 p, 50 n, 69 e^-
- **d.** 69 p, 50 n, 69 e
- **e.** 50 p, 119 n, 52 e^-

2.22 The atomic mass of Ga is 69.72 amu. There are only two naturally occurring isotopes of gallium: ^{69}Ga, with a mass of 69.0 amu, and ^{71}Ga, with a mass of 71.0 amu. The natural abundance of the ^{69}Ga isotope is approximately:

- **a.** 15%
- **b.** 30%
- **c.** 50%
- **d.** 65%
- **e.** 80%

2.23 In which of the following are the name and formula correctly paired?

- **a.** sodium sulfite: Na_2S
- **b.** calcium carbonate: $Ca(CO_3)_2$
- **c.** magnesium hydroxide: $Mg(OH)_2$
- **d.** nitrite: NO_2
- **e.** iron (III) oxide: FeO

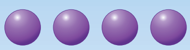
2.24 A chunk of an unidentified element (let's call it element "X") is reacted with sulfur to form an ionic compound with the chemical formula X_2S. Which of the following elements is the most likely identity of X?

- **a.** Mg
- **b.** Li
- **c.** Al
- **d.** C
- **e.** Cl

Concept Explorations

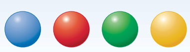
2.25 Average Atomic Mass

Part 1: Consider the four identical spheres below, each with a mass of 2.00 g.



Calculate the average mass of a sphere in this sample.

Part 2: Now consider a sample that consists of four spheres, each with a different mass: blue mass is 2.00 g, red mass is 1.75 g, green mass is 3.00 g, and yellow mass is 1.25 g.



- **a.** Calculate the average mass of a sphere in this sample.
- **b.** How does the average mass for a sphere in this sample compare with the average mass of the sample that consisted just of the blue spheres? How can such different samples have their averages turn out the way they did?

Part 3: Consider two jars. One jar contains 100 blue spheres, and the other jar contains 25 each of red, blue, green, and yellow colors mixed together.

- **a.** If you were to remove 50 blue spheres from the jar containing just the blue spheres, what would be total mass of spheres left in the jar? (Note that the masses of the spheres are given in Part 2.)

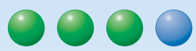
b. If you were to remove 50 spheres from the jar containing the mixture (assume you get a representative distribution of colors), what would be the total mass of spheres left in the jar?

c. In the case of the mixture of spheres, does the average mass of the spheres *necessarily* represent the mass of an individual sphere in the sample?

d. If you had 80.0 grams of spheres from the blue sample, how many spheres would you have?

e. If you had 60.0 grams of spheres from the mixed-color sample, how many spheres would you have? What assumption did you make about your sample when performing this calculation?

Part 4: Consider a sample that consists of three green spheres and one blue sphere. The green mass is 3.00 g, and the blue mass is 1.00 g.



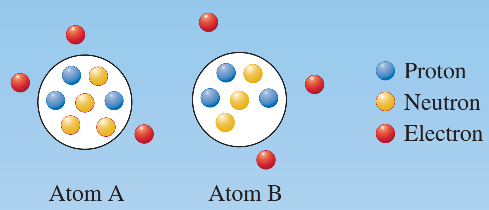
a. Calculate the *fractional abundance* of each sphere in the sample.

b. Use the fractional abundance to calculate the average mass of the spheres in this sample.

c. How are the ideas developed in this Concept Exploration related to the atomic masses of the elements?

2.26 Model of the Atom

Consider the following depictions of two atoms, which have been greatly enlarged so you can see the subatomic particles.



a. How many protons are present in atom A?

b. What is the significance of the number of protons depicted in atom A or any atom?

c. Can you identify the real element represented by the drawing of atom A? If so, what element does it represent?

d. What is the charge on element A? Explain how you arrived at your answer.

e. Write the nuclide symbol of atom A.

f. Write the atomic symbol and the atomic number of atom B.

g. What is the mass number of atom B? How does this mass number compare with that of atom A?

h. What is the charge on atom B?

i. Write the nuclide symbol of element B.

j. Draw pictures like those above of ${}^{6}_{3}Li^{+}$ and ${}^{6}_{3}Li^{-}$ atoms. What are the mass number and atomic number of each of these atoms?

k. Consider the two atoms depicted in this problem and the two that you just drew. What is the total number of lithium isotopes depicted? How did you make your decision?

l. Is the mass number of an isotope of an atom equal to the mass of the isotope of the atom? Be sure to explain your answer.

Conceptual Problems

2.27 One of the early models of the atom proposed that atoms were wispy balls of positive charge with the electrons evenly distributed throughout. What would you expect to observe if you conducted Rutherford's experiment and the atom had this structure?

2.28 A friend is trying to balance the following equation:

$$N_2 + H_2 \longrightarrow NH_3$$

He presents you with his version of the "balanced" equation:

$$N + H_3 \longrightarrow NH_3$$

You immediately recognize that he has committed a serious error; however, he argues that there is nothing wrong, since the equation is balanced. What reason can you give to convince him that his "method" of balancing the equation is flawed?

2.29 Given that the periodic table is an organizational scheme for the elements, what might be some other logical ways in which to group the elements that would provide meaningful chemical information in a periodic table of your own devising?

2.30 You discover a new set of polyatomic anions that has the newly discovered element "X" combined with oxygen. Since you made the discovery, you get to choose the names for these new polyatomic ions. In developing the names, you want to pay attention to convention. A colleague in your lab has come up with a name that she really likes for one of the ions. How would you name the following in a way consistent with her name?

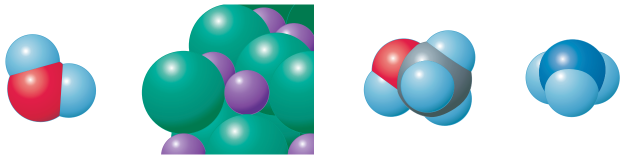
Formula	*Name*
XO_4^{2-}	
XO_3^{2-}	
XO_2^{2-}	excite
XO^{2-}	

2.31 You have the mythical metal element "X" that can exist as X^+, X^{2+}, and X^{5+} ions.

a. What would be the chemical formula for compounds formed from the combination of each of the X ions and SO_4^{2-}?

b. If the name of the element X is exy, what would be the names of each of the compounds from part a of this problem?

2.32 Match the molecular model with the correct chemical formula: CH_3OH, NH_3, KCl, H_2O.



(a) (b) (c) (d)

2.33 Consider a hypothetical case in which the charge on a proton is twice that of an electron. Using this hypothetical case, and the fact that atoms maintain a charge of 0, how many protons, neutrons, and electrons would a potassium-39 atom contain?

2.34 Currently, the atomic mass unit (amu) is based on being exactly one-twelfth the mass of a carbon-12 atom and is equal to 1.66×10^{-27} kg.

a. If the amu were based on sodium-23 with a mass equal to exactly 1/23 of the mass of a sodium-23 atom, would the mass of the amu be different?

b. If the new mass of the amu based on sodium-23 is 1.67×10^{-27} kg, how would the mass of a hydrogen atom, in amu, compare with the current mass of a hydrogen atom in amu?

2.35 For each of the following chemical reactions, write the correct element and/or compound symbols, formulas, and coefficients needed to produce complete, balanced equations. In all cases, the reactants are elements and the products are binary compounds.

a. ______ + ______ $\longrightarrow$ LiCl

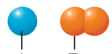
b. ______ + ______ $\longrightarrow$ Na_2S

c. Al + Br_2 $\longrightarrow$

d. Mg + N_2 $\longrightarrow$

e. ______ + ______ $\longrightarrow$ Ca_3P_2

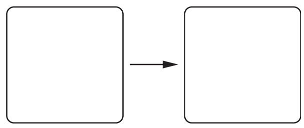
2.36 You perform a chemical reaction using the hypothetical elements A and B. These elements are represented by their molecular models shown below:



A B_2

The product of the reaction represented by molecular models is

a. Using the molecular models and the boxes, present a balanced chemical equation for the reaction of elements A and B.



b. Using the symbols A and B_2 for the chemical reaction, write a balanced chemical equation.

c. What are some real-element possibilities for element B?

Practice Problems

2.37 What is the name of the element represented by each of the following atomic symbols?

a. Ar **b.** Zn **c.** Ag **d.** Mg

2.38 For each atomic symbol, give the name of the element.

a. Ca **b.** Cu **c.** Hg **d.** Sn

2.39 Give the atomic symbol for each of the following elements.

a. potassium **b.** sulfur **c.** iron **d.** manganese

2.40 Give the atomic symbol for each of the following elements.

a. carbon **b.** sodium **c.** nickel **d.** lead

Electrons, Protons, and Neutrons

2.41 A student has determined the mass-to-charge ratio for an electron to be 5.64×10^{-12} kg/C. In another experiment, using Millikan's oil-drop apparatus, he found the charge on the electron to be 1.605×10^{-19} C. What would be the mass of the electron, according to these data?

2.42 The mass-to-charge ratio for the positive ion F^+ is 1.97×10^{-7} kg/C. Using the value of 1.602×10^{-19} C for the charge on the ion, calculate the mass of the fluorine atom. (The mass of the electron is negligible compared with that of the ion, so the ion mass is essentially the atomic mass.)

2.43 The following table gives the number of protons and neutrons in the nuclei of various atoms. Which atom is the isotope of atom A? Which atom has the same mass number as atom A?

	Protons	*Neutrons*
Atom A	18	19
Atom B	16	19
Atom C	18	18
Atom D	17	20

2.44 The following table gives the number of protons and neutrons in the nuclei of various atoms. Which atom is the isotope of atom A? Which atom has the same mass number as atom A?

	Protons	*Neutrons*
Atom A	32	39
Atom B	32	38
Atom C	38	50
Atom D	33	38

2.45 Naturally occurring chlorine is a mixture of the isotopes Cl-35 and Cl-37. How many protons and how many neutrons are there in each isotope? How many electrons are there in the neutral atoms?

2.46 Naturally occurring lithium is a mixture of ${}^{6}_{3}Li$ and ${}^{7}_{3}Li$. Give the number of protons, neutrons, and electrons in the neutral atom of each isotope.

2.47 What is the nuclide symbol for the nucleus that contains 14 protons and 14 neutrons?

2.48 An atom contains 11 protons and 11 neutrons. What is the nuclide symbol for the nucleus?

Atomic Masses

2.49 Ammonia is a gas with a characteristic pungent odor. It is sold as a water solution for use in household cleaning. The gas is a compound of nitrogen and hydrogen in the atomic ratio 1 : 3. A sample of ammonia contains 7.933 g N and 1.712 g H. What is the atomic mass of N relative to H?

2.50 Hydrogen sulfide is a gas with the odor of rotten eggs. The gas can sometimes be detected in automobile exhaust. It is a compound of hydrogen and sulfur in the atomic ratio 2 : 1. A sample of hydrogen sulfide contains 0.587 g H and 9.330 g S. What is the atomic mass of S relative to H?

2.51 Calculate the atomic mass of an element with two naturally occurring isotopes, from the following data:

Isotope	*Isotopic Mass (amu)*	*Fractional Abundance*
X-63	62.930	0.6909
X-65	64.928	0.3091

What is the identity of element X?

2.52 An element has two naturally occurring isotopes with the following masses and abundances:

Isotopic Mass (amu)	*Fractional Abundance*
49.9472	2.500×10^{-3}
50.9440	0.9975

What is the atomic mass of this element? What is the identity of the element?

2.53 An element has three naturally occurring isotopes with the following masses and abundances:

Isotopic Mass (amu)	*Fractional Abundance*
38.964	0.9326
39.964	1.000×10^{-4}
40.962	0.0673

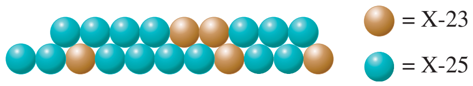
Calculate the atomic mass of this element. What is the identity of the element?

2.54 An element has three naturally occurring isotopes with the following masses and abundances:

Isotopic Mass (amu)	*Fractional Abundance*
27.977	0.9221
28.976	0.0470
29.974	0.0309

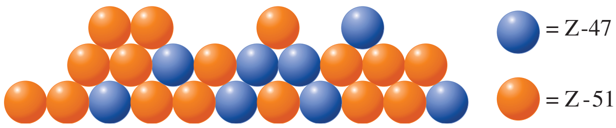
Calculate the atomic mass of this element. What is the identity of the element?

2.55 While traveling to a distant universe, you discover the hypothetical element "X." You obtain a representative sample of the element and discover that it is made up of two isotopes, X-23 and X-25. To help your science team calculate the atomic mass of the substance, you send the following drawing of your sample with your report.



In the report, you also inform the science team that the gold atoms are X-23, which have an isotopic mass of 23.02 amu, and the green atoms are X-25, which have an isotopic mass of 25.147 amu. What is the atomic mass of element X?

2.56 While roaming a parallel universe, you discover the hypothetical element "Z." You obtain a representative sample of the element and discover that it is made up of two isotopes, Z-47 and Z-51. To help your science team calculate the atomic mass of the substance, you send the following drawing of your sample with your report.



In the report, you also inform the science team that the blue atoms are Z-47, which have an isotopic mass of 47.510 amu, and the orange atoms are Z-51, which have an isotopic mass of 51.126 amu. What is the atomic mass of element Z?

Periodic Table

2.57 Identify the group and period for each of the following. Refer to the periodic table (Figure 2.15 or inside front cover). Label each as a metal, nonmetal, or metalloid.

a. C **b.** Po **c.** Cr **d.** Mg **e.** B

2.58 Refer to the periodic table (Figure 2.15 or inside front cover) and obtain the group and period for each of the following elements. Also determine whether the element is a metal, nonmetal, or metalloid.

a. S **b.** Fe **c.** Ba **d.** Cu **e.** Ne

2.59 Refer to the periodic table (Figure 2.15 or inside front cover) and answer the following questions.

a. What Group VIA element is a metalloid?
b. What is the Group IIIA element in Period 3?

2.60 Refer to the periodic table (Figure 2.15 or inside front cover) and answer the following questions.

a. What Group VA element is a metal?
b. What is the Group IIA element in Period 3?

2.61 Give one example (atomic symbol and name) for each of the following.

a. a main-group (representative) element in the second period
b. an alkali metal
c. a transition element in the fourth period
d. a lanthanide element

2.62 Give one example (atomic symbol and name) for each of the following.

a. a transition element in the fifth period
b. a halogen
c. a main-group (representative) element in the second period
d. an actinide element

Molecular and Ionic Substances

2.63 The normal form of the element sulfur is a brittle, yellow solid. This is a molecular substance, S_8. If this solid is vaporized, it first forms S_8 molecules; but at high temperature, S_2 molecules are formed. How do the molecules of the solid sulfur and of the hot vapor differ? How are the molecules alike?

2.64 White phosphorus is available in sticks, which have a waxy appearance. This is a molecular substance, P_4. When this solid is vaporized, it first forms P_4 molecules; but at high temperature, P_2 molecules are formed. How do the molecules of white phosphorus and those of the hot vapor differ? How are the molecules alike?

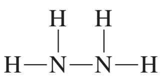
2.65 A 1.50-g sample of nitrous oxide (an anesthetic, sometimes called laughing gas) contains 2.05×10^{22} N_2O molecules. How many nitrogen atoms are in this sample? How many nitrogen atoms are in 1.00 g of nitrous oxide?

2.66 Nitric acid is composed of HNO_3 molecules. A sample weighing 4.50 g contains 4.30×10^{22} HNO_3 molecules. How many nitrogen atoms are in this sample? How many oxygen atoms are in 2.81 g of nitric acid?

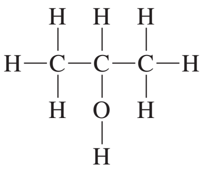
2.67 A sample of ammonia, NH_3, contains 3.3×10^{21} hydrogen atoms. How many NH_3 molecules are in this sample?

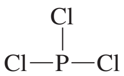
2.68 A sample of ethanol (ethyl alcohol), C_2H_5OH, contains 4.2×10^{23} hydrogen atoms. How many C_2H_5OH molecules are in this sample?

2.69 Give the molecular formula for each of the following structural formulas.

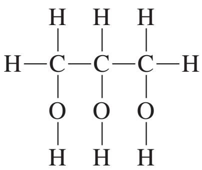
a. 
hydrazine

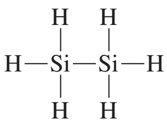
b. H—O—O—H
hydrogen peroxide

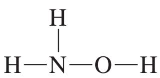
c. 
isopropyl alcohol

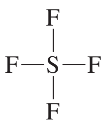
d. 
phosphorus trichloride

2.70 What molecular formula corresponds to each of the following structural formulas?

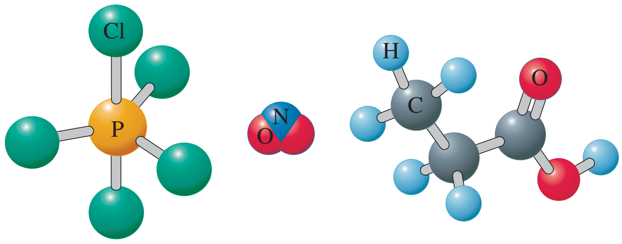
a. 
glycerol

b. 
disilane

c. 
hydroxylamine

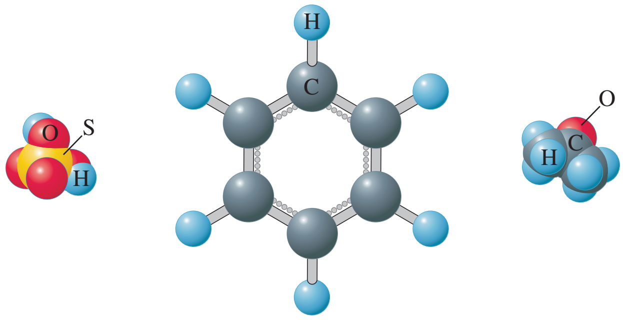
d. 
sulfur tetrafluoride

2.71 Write the molecular formula for each of the following compounds represented by molecular models.



(a) (b) (c)

2.72 Write the molecular formula for each of the following compounds represented by molecular models.



(a) (b) (c)

2.73 Iron(II) nitrate has the formula $Fe(NO_3)_2$. What is the ratio of iron atoms to oxygen atoms in this compound?

2.74 Ammonium phosphate, $(NH_4)_3PO_4$, has how many hydrogen atoms for each oxygen atom?

2.75 Write the formula for the compound of each of the following pairs of ions.

a. Fe^{3+} and CN^-
b. K^+ and SO_4^{2-}
c. Li^+ and N^{3-}
d. Ca^{2+} and P^{3-}

2.76 For each of the following pairs of ions, write the formula of the corresponding compound.

a. Co^{2+} and N^{3-}
b. NH_4^+ and PO_4^{3-}
c. Na^+ and Co_3^{2-}
d. Fe^{3+} and OH^-

Chemical Nomenclature

2.77 Name the following compounds.

a. Na_2SO_4 **b.** CaS **c.** CuCl **d.** Cr_2O_3

2.78 Name the following compounds.

a. Na_2O
b. Mn_2O_3
c. NH_4HCO_3
d. $Cu(NO_3)_2$

2.79 Write the formulas of:

a. lead(II) permanganate
b. barium hydrogen carbonate
c. cesium sulfide
d. iron(II) acetate

2.80 Write the formulas of:
a. sodium thiosulfate
b. copper(I) hydroxide
c. calcium hydrogen carbonate
d. nickel(II) phosphide

2.81 For each of the following binary compounds, decide whether the compound is expected to be ionic or molecular.
a. SeF_4 **b.** LiBr **c.** SiF_4 **d.** Cs_2O

2.82 For each of the following binary compounds, decide whether the compound is expected to be ionic or molecular.
a. AlN **b.** As_4O_6 **c.** ClF_3 **d.** Fe_2O_3

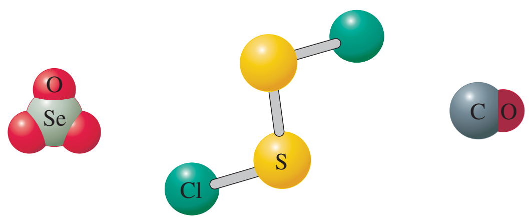
2.83 Give systematic names to the following binary compounds.
a. N_2O **b.** P_4O_{10} **c.** $AsCl_3$ **d.** Cl_2O_7

2.84 Give systematic names to the following binary compounds.
a. N_2F_2 **b.** CCl_4 **c.** N_2O_5 **d.** As_4O_6

2.85 Write the formulas of the following compounds.
a. nitrogen tribromide
b. xenon tetroxide
c. oxygen difluoride
d. dichlorine pentoxide

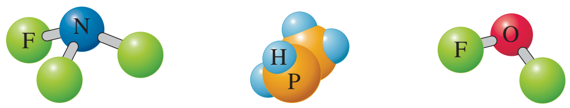
2.86 Write the formulas of the following compounds.
a. chlorine trifluoride
b. nitrogen dioxide
c. dinitrogen tetrafluoride
d. phosphorus pentafluoride

2.87 Write the systematic name for each of the following compounds represented by a molecular model.



(a) (b) (c)

2.88 Write the systematic name for each of the following molecules represented by a molecular model.



(a) (b) (c)

2.89 Give the name and formula of the acid corresponding to each of the following oxoanions.
a. bromate ion, BrO_3^-
b. hyponitrite ion, $N_2O_2^{2-}$
c. disulfite ion, $S_2O_5^{2-}$
d. arsenate ion, AsO_4^{3-}

2.90 Give the name and formula of the acid corresponding to each of the following oxoanions.
a. selenite ion, SeO_3^{2-}
b. sulfite ion, SO_3^{2-}
c. hypoiodite ion, IO^-
d. nitrite ion, NO_2^-

2.91 Glauber's salt has the formula $Na_2SO_4{\cdot}10H_2O$. What is the chemical name of this substance?

2.92 Emerald-green crystals of the substance $NiSO_4{\cdot}6H_2O$ are used in nickel plating. What is the chemical name of this compound?

2.93 Iron(II) sulfate heptahydrate is a blue-green, crystalline compound used to prepare other iron compounds. What is the formula of iron(II) sulfate heptahydrate?

2.94 Cobalt(II) chloride hexahydrate has a pink color. It loses water on heating and changes to a blue-colored compound. What is the formula of cobalt(II) chloride hexahydrate?

Chemical Equations

2.95 For the balanced chemical equation $Pb(NO_3)_2 + K_2CO_3 \longrightarrow PbCO_3 + 2KNO_3$, how many oxygen atoms are on the left side?

2.96 In the equation $2PbS + O_2 \longrightarrow 2PbO + 2SO_2$, how many oxygen atoms are there on the right side? Is the equation balanced as written?

2.97 Balance the following equations.
a. $Sn + NaOH \longrightarrow Na_2SnO_2 + H_2$
b. $Al + Fe_3O_4 \longrightarrow Al_2O_3 + Fe$
c. $CH_3OH + O_2 \longrightarrow CO_2 + H_2O$
d. $P_4O_{10} + H_2O \longrightarrow H_3PO_4$
e. $PCl_5 + H_2O \longrightarrow H_3PO_4 + HCl$

2.98 Balance the following equations.
a. $Cl_2O_7 + H_2O \longrightarrow HClO_4$
b. $MnO_2 + HCl \longrightarrow MnCl_2 + Cl_2 + H_2O$
c. $Na_2S_2O_3 + I_2 \longrightarrow NaI + Na_2S_4O_6$
d. $Al_4C_3 + H_2O \longrightarrow Al(OH)_3 + CH_4$
e. $NO_2 + H_2O \longrightarrow HNO_3 + NO$

2.99 Solid calcium phosphate and aqueous sulfuric acid solution react to give calcium sulfate, which comes out of the solution as a solid. The other product is phosphoric acid, which remains in solution. Write a balanced equation for the reaction using complete formulas for the compounds with phase labels.

2.100 Solid sodium metal reacts with water, giving a solution of sodium hydroxide and releasing hydrogen gas. Write a balanced equation for the reaction using complete formulas for the compounds with phase labels.

2.101 An aqueous solution of ammonium chloride and barium hydroxide is heated, and the compounds react to give off ammonia gas. Barium chloride solution and water are also products. Write a balanced equation for the reaction using complete formulas for the compounds with phase labels; indicate that the reactants are heated.

2.102 Lead metal is produced by heating solid lead(II) sulfide with solid lead(II) sulfate, resulting in liquid lead and sulfur dioxide gas. Write a balanced equation for the reaction using complete formulas for the compounds with phase labels; indicate that the reactants are heated.

General Problems

2.103 Two samples of different compounds of nitrogen and oxygen have the following compositions. Show that the compounds follow the law of multiple proportions. What is the ratio of oxygen in the two compounds for a fixed amount of nitrogen?

	Amount N	*Amount O*
Compound A	1.206 g	2.755 g
Compound B	1.651 g	4.714 g

2.104 Two samples of different compounds of sulfur and oxygen have the following compositions. Show that the compounds follow the law of multiple proportions. What is the ratio of oxygen in the two compounds for a fixed amount of sulfur?

	Amount S	*Amount O*
Compound A	1.210 g	1.811 g
Compound B	1.783 g	1.779 g

2.105 In a series of oil-drop experiments, the charges measured on the oil drops were -3.20×10^{-19} C, -6.40×10^{-19} C, -9.60×10^{-19} C, and -1.12×10^{-18} C. What is the smallest difference in charge between any two drops? If this is assumed to be the charge on the electron, how many excess electrons are there on each drop?

2.106 In a hypothetical universe, an oil-drop experiment gave the following measurements of charges on oil drops: -5.55×10^{-19} C, -9.25×10^{-19} C, -1.11×10^{-18} C, and -1.48×10^{-18} C. Assume that the smallest difference in charge equals the unit of negative charge in this universe. What is the value of this unit of charge? How many units of excess negative charge are there on each oil drop?

2.107 Compounds of europium, Eu, are used to make color television screens. The europium nucleus has a charge of +63. How many electrons are there in the neutral atom? in the Eu^{3+} ion?

2.108 Cesium, Cs, is used in photoelectric cells ("electric eyes"). The cesium nucleus has a charge of +55. What is the number of electrons in the neutral atom? in the Cs^{+} ion?

2.109 A nucleus of mass number 81 contains 46 neutrons. An atomic ion of this element has 36 electrons in it. Write the symbol for this atomic ion (give the symbol for the nucleus and give the ionic charge as a right superscript).

2.110 One isotope of a metallic element has mass number 80 and has 55 neutrons in the nucleus. An atomic ion has 23 electrons. Write the symbol for this ion (give the symbol for the nucleus and give the ionic charge as a right superscript).

2.111 Obtain the fractional abundances for the two naturally occurring isotopes of copper. The masses of the isotopes are $^{63}_{29}Cu$, 62.9298 amu; $^{65}_{29}Cu$, 64.9278 amu. The atomic mass is 63.546 amu.

2.112 Silver has two naturally occurring isotopes, one of mass 106.91 amu and the other of mass 108.90 amu. Find the fractional abundances for these two isotopes. The atomic mass is 107.87 amu.

2.113 Identify the following elements, giving their name and atomic symbol.
- **a.** a nonmetal that is normally a liquid
- **b.** a normally gaseous element in Group IA
- **c.** a transition element in Group VB, Period 5
- **d.** the halogen in Period 2

2.114 Identify the following elements, giving their name and atomic symbol.
- **a.** a normally liquid element in Group VIIA
- **b.** a metal that is normally a liquid
- **c.** a main-group element in Group IIIA, Period 4
- **d.** the alkali metal in Period 5

2.115 Give the names of the following ions.
a. Cr^{3+} **b.** Pb^{4+} **c.** Cu^{+} **d.** Cu^{2+}

2.116 Give the names of the following ions.
a. Mn^{2+} **b.** Ni^{2+} **c.** Co^{2+} **d.** Co^{3+}

2.117 Write formulas for all the ionic compounds that can be formed by combinations of these ions: Na^{+}, Ni^{2+}, $SO_4{}^{2-}$, and Cl^{-}.

2.118 Write formulas for all the ionic compounds that can be formed by combinations of these ions: Ca^{2+}, Cr^{3+}, O^{2-}, and $NO_3{}^{-}$.

2.119 Name the following compounds.
- **a.** $Sn_3(PO_4)_2$
- **b.** NH_4NO_2
- **c.** $Mg(OH)_2$
- **d.** $CrSO_4$

2.120 Name the following compounds.
a. $Cu(NO_2)_2$ **b.** $(NH_4)_3P$
c. Na_2SO_4 **d.** $HgCl_2$

2.121 Give the formulas for the following compounds.
- **a.** mercury(I) sulfide
- **b.** cobalt(III) sulfite
- **c.** ammonium dichromate
- **d.** aluminum nitride

2.122 Give the formulas for the following compounds.
- **a.** hydrogen peroxide
- **b.** aluminum phosphate
- **c.** lead(IV) phosphide
- **d.** boron trifluoride

2.123 Name the following molecular compounds.
a. $AsBr_3$ **b.** H_2Se **c.** P_2O_5 **d.** SiO_2

2.124 Name the following molecular compounds.
a. ClF_4 **b.** CS_2 **c.** NF_3 **d.** SF_6

2.125 Balance the following equations.
- **a.** $C_2H_6 + O_2 \longrightarrow CO_2 + H_2O$
- **b.** $P_4O_6 + H_2O \longrightarrow H_3PO_3$
- **c.** $KClO_3 \longrightarrow KCl + KClO_4$
- **d.** $(NH_4)_2SO_4 + NaOH \longrightarrow NH_3 + H_2O + Na_2SO_4$
- **e.** $NBr_3 + NaOH \longrightarrow N_2 + NaBr + HOBr$

2.126 Balance the following equations.

a. $NaOH + H_3PO_4 \longrightarrow Na_3PO_4 + H_2O$
b. $SiCl_4 + H_2O \longrightarrow SiO_2 + HCl$
c. $Ca_3(PO_4)_2 + C \longrightarrow Ca_3P_2 + CO$
d. $H_2S + O_2 \longrightarrow SO_2 + H_2O$
e. $N_2O_5 \longrightarrow NO_2 + O_2$

2.127 A monatomic ion has a charge of +2. The nucleus of the ion has a mass number of 62. The number of neutrons in the nucleus is 1.21 times that of the number of protons. How many electrons are in the ion? What is the name of the element?

2.128 A monatomic ion has a charge of +1. The nucleus of the ion has a mass number of 85. The number of neutrons in the nucleus is 1.30 times that of the number of protons. How many electrons are in the ion? What is the name of the element?

2.129 Natural carbon, which has an atomic mass of 12.011 amu, consists of carbon-12 and carbon-13 isotopes. Given that the mass of carbon-13 is 13.00335 amu, what would be the average atomic mass (in amu) of a carbon sample prepared by mixing equal numbers of carbon atoms from a sample of natural carbon and a sample of pure carbon-13?

2.130 Natural chlorine, which has an atomic mass of 35.4527 amu, consists of chlorine-35 and chlorine-37 isotopes. Given that the mass of chlorine-35 is 34.96885 amu, what is the average atomic mass (in amu) of a chlorine sample prepared by mixing equal numbers of chlorine atoms from a sample of natural chlorine and a sample of pure chlorine-35?

2.131 Describe the island of stability. What nuclide is predicted to be most stable?

2.132 Write the equation for the nuclear reaction in which element 112 was first produced.

Strategy Problems

2.133 Correct any mistakes in the naming of the following compounds or ions.

SO_3: sulfite
NO_2: nitrite
$PO_4{}^{3-}$: phosphite ion
N_2: nitride
$Mg(OH)_2$: maganese dihydroxide

2.134 An unknown metal (let's call it "M") is reacted with sulfur to produce a compound with the chemical formula M_2S_3. What is the charge on the metal in the compound M_2S_3? Name a metal that could be metal M.

2.135 Aluminum is reacted with sulfur to form a binary compound. Write the balanced chemical reaction and write the name of the compound formed by the reaction.

2.136 Ammonia gas reacts with molecular oxygen gas to form nitrogen monoxide gas and liquid water. Write the complete balanced reaction with all proper state symbols.

2.137 A hypothetical element X is found to have an atomic mass of 37.45 amu. Element X has only two isotopes, X-37 and X-38. The X-37 isotope has a fractional abundance of 0.7721 and an isotopic mass of 37.24. What is the isotopic mass of the other isotope?

2.138 A monotomic ion has a charge of +3. The nucleus of the ion has a mass number of 27. The number of neutrons in the nucleus is equal to the number of electrons in a S^{2+} ion. Identify the element and indicate the number of protons, neutrons, and electrons.

2.139 A small crystal of $CaCl_2$ that weighs 0.12 g contains 6.5×10^{20} formula units of $CaCl_2$. What is the total number of ions (cations and anions) that make up this crystal?

2.140 Write the formulas and names for all the ionic compounds that can form by combinations of the following ions: Mg^{2+}, Pb^{4+}, the carbonate anion, and the phosphide anion.

2.141 Name the following compounds:

SO_3
HNO_2
Mg_3N_3
HBr
$Cu_3(PO_4)_2$
$CuSO_4{\cdot}5H_2O$

2.142 The $IO_3{}^-$ anion is called iodate. There are three related ions: IO^-, $IO_2{}^-$, and $IO_4{}^-$. Using what you have learned about similar groups of anions, write the name for each of the following compounds:

$Pb(IO_3)_2$
KIO_4
$Zn(IO)_2$
$Al(IO_2)_3$

Cumulative-Skills Problems

2.143 There are 2.619×10^{22} atoms in 1.000 g of sodium. Assume that sodium atoms are spheres of radius 1.86 Å and that they are lined up side by side. How many miles in length is the line of sodium atoms?

2.144 There are 1.699×10^{22} atoms in 1.000 g of chlorine. Assume that chlorine atoms are spheres of radius 0.99 Å and that they are lined up side by side. How many miles in length is the line of chlorine atoms?

2.145 A sample of green crystals of nickel(II) sulfate heptahydrate was heated carefully to produce the bluish green nickel(II) sulfate hexahydrate. What are the formulas of the hydrates? If 8.753 g of the heptahydrate produces 8.192 g of the hexahydrate, how many grams of anhydrous nickel(II) sulfate could be obtained?

2.146 Cobalt(II) sulfate heptahydrate has pink-colored crystals. When heated carefully, it produces cobalt(II) sulfate monohydrate, which has red crystals. What are the formulas of these hydrates? If 3.548 g of the heptahydrate yields 2.184 g of the monohydrate, how many grams of the anhydrous cobalt(II) sulfate could be obtained?

2.147 A sample of metallic element X, weighing 3.177 g, combines with 0.6015 L of O_2 gas (at normal pressure and 20.0°C) to form the metal oxide with the formula XO. If the density of O_2 gas under these conditions is 1.330 g/L, what is the mass of this oxygen? The atomic mass of oxygen is 15.9994 amu. What is the atomic mass of X? What is the identity of X?

2.148 A sample of metallic element X, weighing 4.315 g, combines with 0.4810 L of Cl_2 gas (at normal pressure and 20.0°C) to form the metal chloride with the formula XCl. If the density of Cl_2 gas under these conditions is 2.948 g/L, what is the mass of the chlorine? The atomic mass of chlorine is 35.453 amu. What is the atomic mass of X? What is the identity of X?