

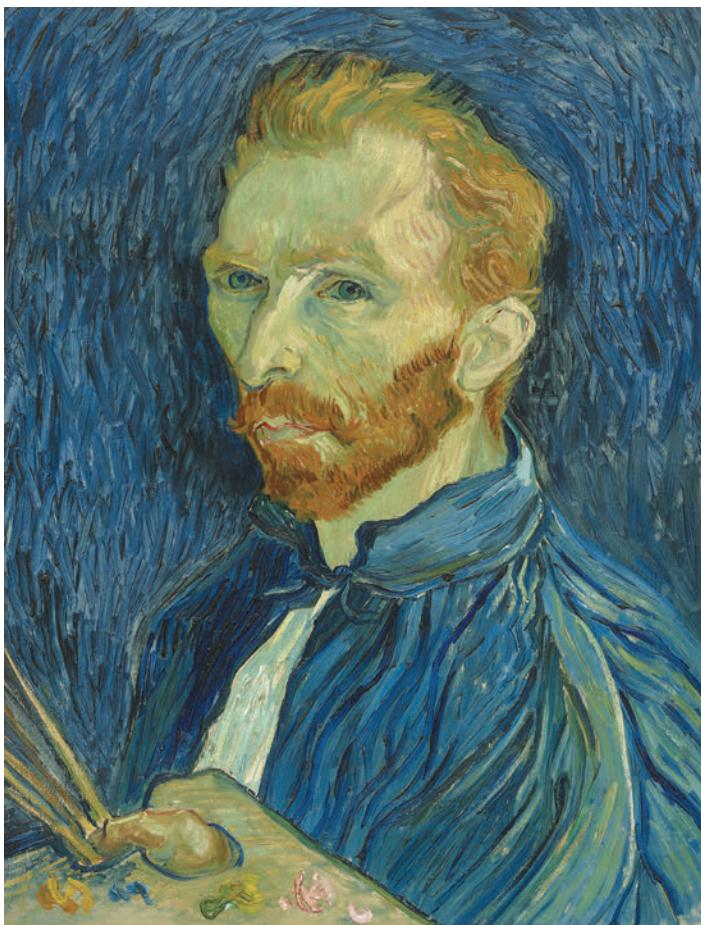
WHAT'S AHEAD

- 1.1 ► The Study of Chemistry
- 1.2 ► Classifications of Matter
- 1.3 ► Properties of Matter
- 1.4 ► The Nature of Energy
- 1.5 ► Units of Measurement
- 1.6 ► Uncertainty in Measurement
- 1.7 ► Dimensional Analysis

1

INTRODUCTION: MATTER, ENERGY, AND MEASUREMENT

1.1 | The Study of Chemistry



The title of this book—*Chemistry: The Central Science*—reflects the fact that much of what goes on in the world around us involves chemistry. Everyday chemical processes include the changes that produce the brilliant colors of flowers, the ways our bodies process the food we eat, and the electrical energy that powers our cell phones.

This first chapter provides an overview of what chemistry is about and what chemists do. The “What’s Ahead” list gives an overview of the chapter organization and of some of the ideas we will consider.

Chemistry is the study of matter, its properties, and the changes that matter undergoes. As you progress in your study, you will come to see how chemical principles operate in all aspects of our lives, from everyday activities like food preparation to more complex processes such as those that operate in the environment. We will also learn how the properties of substances can be tailored for specific applications by controlling their composition and structure. For example, the synthetic pigments chemists developed in the nineteenth century were used extensively by impressionist artists such as van Gogh.

Chemistry is at the heart of many changes we see in the world around us, and it accounts for the myriad different properties we see in matter. To understand how these changes and properties arise, we need to look far beneath the surfaces of our everyday observations.

By the end of this section, you should be able to

- Appreciate the scope of chemistry

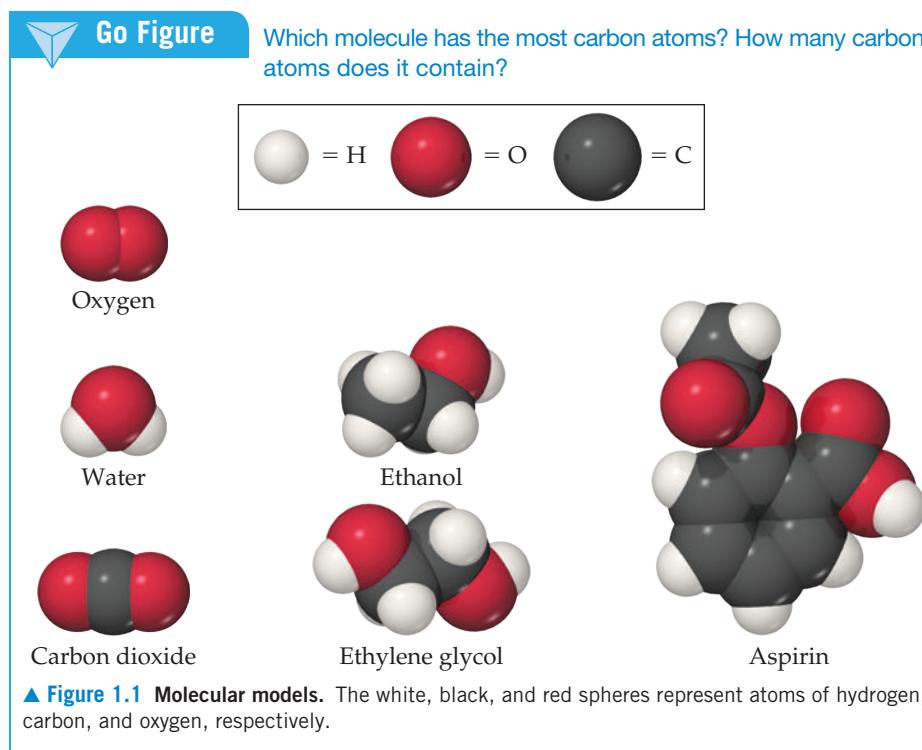
The Atomic and Molecular Perspective of Chemistry

Chemistry is the study of the properties and behavior of matter. **Matter** is the physical material of the universe; it is anything that has mass and occupies space. A **property** is any characteristic that allows us to recognize a particular type of matter and to distinguish it from other types. This book, your body, the air you are breathing, and the clothes you are wearing are all samples of matter. We observe a tremendous variety of matter in our world, but countless experiments have shown that all matter is comprised of combinations of only about 100 substances called **elements**. One of our major goals will be to relate the properties of matter to its composition, that is, to the particular elements it contains.

Chemistry also provides a background for understanding the properties of matter in terms of **atoms**, the almost infinitesimally small building blocks of matter. Each element is composed of a unique kind of atom. We will see that the properties of matter relate to both the kinds of atoms the matter contains (*composition*) and the arrangements of these atoms (*structure*).

In **molecules**, two or more atoms are joined in specific shapes. Throughout this text you will see molecules represented using colored spheres to show how the atoms are connected ([Figure 1.1](#)). The color provides a convenient way to distinguish between atoms of different elements. For example, notice that the molecules of ethanol and ethylene glycol in Figure 1.1 have different compositions and structures. Ethanol contains one oxygen atom, depicted by one red sphere. In contrast, ethylene glycol contains two oxygen atoms.

Even apparently minor differences in the composition or structure of molecules can cause profound differences in properties. For example, let’s compare ethanol and ethylene glycol, which appear in Figure 1.1 to be quite similar. Ethanol is the alcohol in beverages such as beer and wine, whereas ethylene glycol is a viscous liquid used as automobile antifreeze. The properties of these two substances differ in many ways, as do their biological activities. Ethanol is consumed throughout the world, but you should *never* consume ethylene glycol because it is highly toxic. One of the challenges chemists undertake is to alter the composition or structure of molecules in a controlled way, creating new substances with different properties. For example, the common drug aspirin, shown in Figure 1.1, was first synthesized in 1897 in a successful attempt to



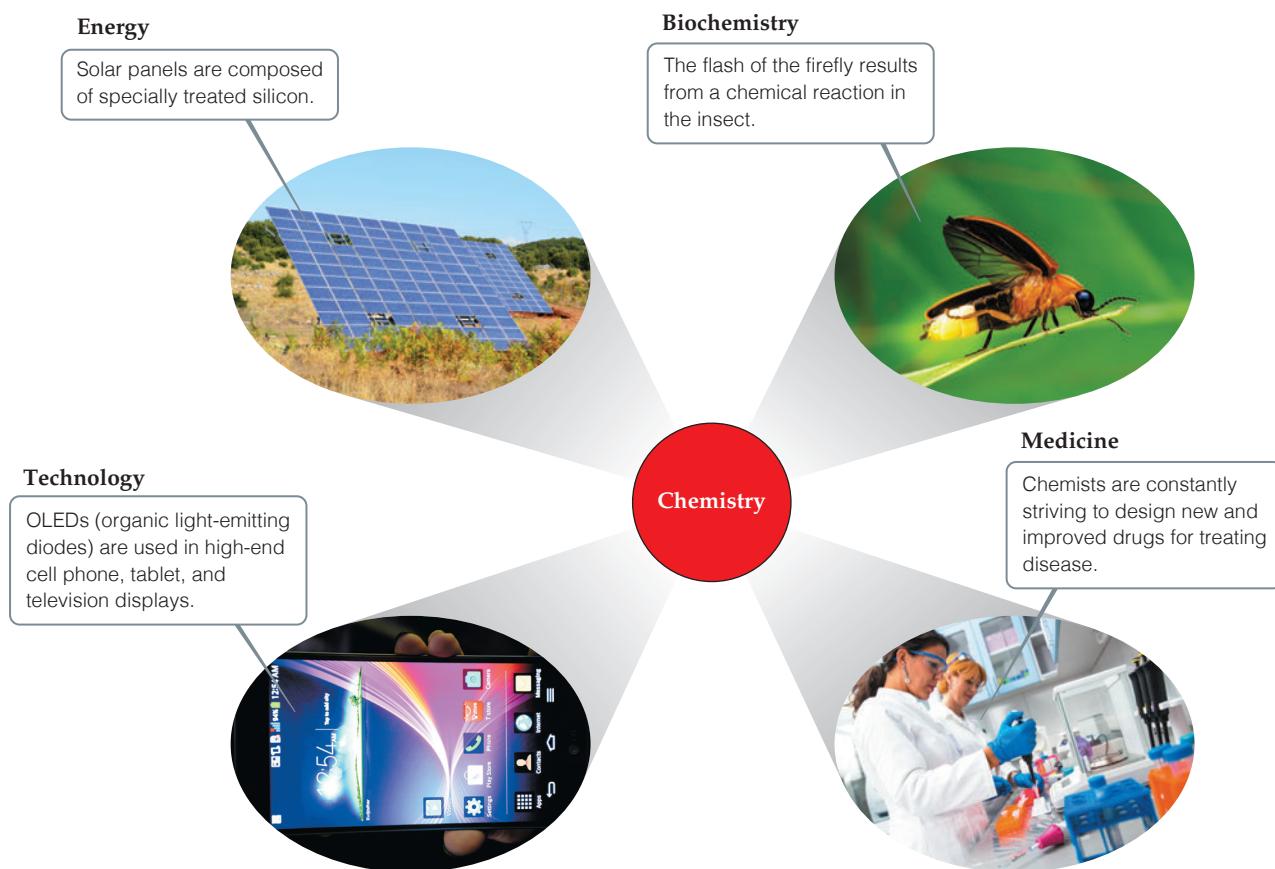
improve on a natural product extracted from willow bark that had long been used to alleviate pain.

Every change in the observable world—from boiling water to the changes that occur as our bodies combat invading viruses—has its basis in the world of atoms and molecules. Thus, as we proceed with our study of chemistry, we will find ourselves thinking in three realms: the *macroscopic* realm of ordinary-sized objects (*macro* = large), the *microscopic* realm of atoms and molecules, and the symbolic realm of how we represent these particles. We make our observations in the macroscopic world, but to understand that world, we must visualize how atoms and molecules behave at the microscopic level. Chemistry is the science that seeks to understand the properties and behavior of matter by studying the properties and behavior of atoms and molecules.

Why Study Chemistry?

Chemistry lies near the heart of many matters of public concern, such as improvement of health care, conservation of natural resources, protection of the environment, and the supply of energy needed to keep society running. Using chemistry, we have discovered and continually improved upon pharmaceuticals, fertilizers and pesticides, plastics, solar panels, light-emitting diodes (LEDs), and building materials. We have also discovered that some chemicals are harmful to our health or the environment. This means that we must be sure that the materials with which we come into contact are safe. As a citizen and consumer, it is in your best interest to understand the effects, both positive and negative, that chemicals can have, in order to arrive at a balanced outlook regarding their uses.

You may be studying chemistry because it is an essential part of your curriculum. Your major might be chemistry, or it could be biology, engineering, pharmacy, agriculture, geology, or some other field. Chemistry is central to a fundamental understanding of governing principles in many science-related fields. For example, our interactions with the material world raise basic questions about the materials around us. **Figure 1.2** illustrates how chemistry is central to several different realms of modern life.



▲ Figure 1.2 Chemistry is central to our understanding of the world around us.

CHEMISTRY PUT TO WORK

Chemistry and the Chemical Industry

Chemistry is all around us. For example, sugar is a chemical extracted from natural sources, refined by industry and sold across the world in a highly purified form. However, few realize the size and importance of the chemical industry. The chemical industry in the United States is estimated to be an \$800 billion enterprise that employs over 800,000 people and accounts for 14% of all U.S. exports.

Who are chemists, and what do they do? People who have degrees in chemistry hold a variety of positions in industry, government, and academia. Those in industry work as laboratory chemists, developing new products (research and development); analyzing materials (quality control); or assisting customers in using products (sales and service). Those with more experience or training may work as managers or company directors. Chemists are important members of the scientific workforce in government (the National Institutes of Health, Department of Energy, and Environmental Protection Agency all employ chemists) and at universities. A chemistry degree is also good preparation for careers in teaching, medicine, biomedical research, information science, environmental work, technical sales, government regulatory agencies, and patent law.

Fundamentally, chemists do three things: They (1) make new types of matter: materials, substances, or combinations of substances with desired properties; (2) measure the properties of matter; and (3) develop models that explain and/or predict the properties of matter. One chemist, for example, may work in the laboratory to discover new drugs. Another may concentrate on the development of new instrumentation to measure properties of matter at the atomic level. Other chemists may use existing

materials and methods to understand how pollutants are transported in the environment or how drugs are processed in the body. Yet another chemist will develop theory, write computer code, and run computer simulations to understand how molecules move and react. The collective chemical enterprise is a rich mix of all of these activities.



▲ Figure 1.3 Common chemicals used for household cleaning.

Self-Assessment Exercise

- 1.1 Which of the following groups of substances involve the use of chemicals? Indicate all that apply.
- (a) Paints, printer toner, food coloring
(b) Computer displays, LED lights, barcode readers

- (c) Antiseptic cream, pain killers, energy drinks
(d) A light-weight bicycle frame, food packaging, a car exhaust catalytic converter
(e) Soap, shampoo, washing powder

1.1 All of them

Answers to Self-Assessment Exercises



1.2 | Classifications of Matter



The study of any science starts with classification. This brings order to what we study much as the sorting of books in a library enables you to find the book you want.

Let's begin our study of chemistry by examining two fundamental ways in which matter is classified. Matter is typically characterized by (1) its physical state (gas, liquid, or solid) and (2) its composition (whether it is an element, a *compound*, or a *mixture*).

By the end of this section, you should be able to

- Understand the basic classification of matter

States of Matter

A sample of matter can be a gas, a liquid, or a solid. These three forms, called the **states of matter**, differ in some of their observable properties.

- A **gas** (also known as vapor) has no fixed volume or shape; rather, it uniformly fills its container. A gas can be compressed to occupy a smaller volume, or it can expand to occupy a larger one.
- A **liquid** has a distinct volume independent of its container, assumes the shape of the portion of the container it occupies, and is not compressible to any appreciable extent.
- A **solid** has both a definite shape and a definite volume and is not compressible to any appreciable extent.

The properties of the states of matter can be understood on the molecular level (Figure 1.4). In a gas the molecules are far apart and moving at high speeds, colliding repeatedly with one another and with the walls of the container. Compressing a gas decreases the amount of space between molecules and increases the frequency of collisions between molecules

but does not alter the size or shape of the molecules. In a liquid, the molecules are packed closely together but still move rapidly. The rapid movement allows the molecules to slide over one another; thus, a liquid pours easily. In a solid the molecules are held tightly together, usually in definite arrangements in which the molecules can wiggle only slightly in their otherwise fixed positions. Thus, the distances between molecules are similar in the liquid and solid states, but while the molecules are for the most part locked in place in a solid, they retain considerable freedom of motion in a liquid. Changes in temperature and/or pressure can lead to conversion from one state of matter to another, illustrated by such familiar processes as ice melting or water vapor condensing.

Pure Substances

Most forms of matter we encounter—the air we breathe (a gas), the fuel we burn in our cars (a liquid), and the road they run on (a solid)—are not chemically pure. We can, however, separate these forms of matter into pure substances. A **pure substance** (usually referred to simply as a *substance*) is matter that has distinct properties and a composition that does not vary from sample to sample. Water and table salt (sodium chloride) are examples of pure substances.

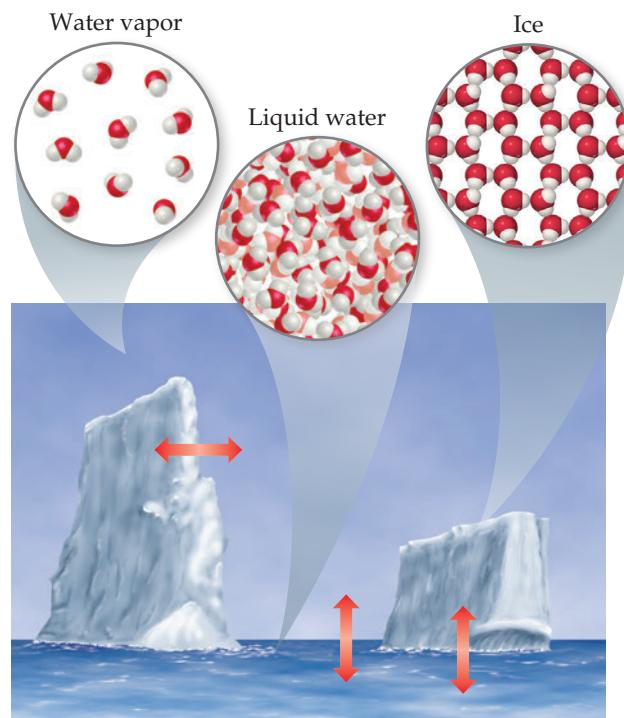
All substances are either elements or compounds.

- **Elements** are substances that cannot be decomposed into simpler substances. On the molecular level, each element is composed of only one kind of atom [Figure 1.5(a and b)].
- **Compounds** are substances composed of two or more elements; they contain two or more kinds of atoms [Figure 1.5(c)]. Water, for example, is a compound composed of two elements: hydrogen and oxygen.

Figure 1.5(d) shows a mixture of substances. **Mixtures** are combinations of two or more substances in which each substance retains its chemical identity.

Go Figure

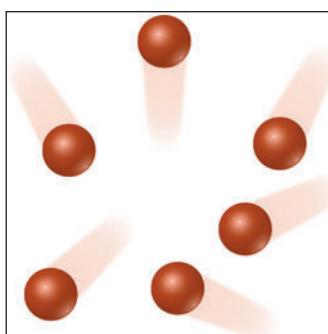
In which form of water are the water molecules farthest apart?



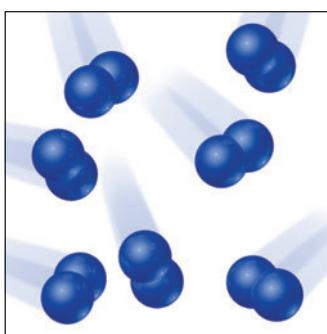
▲ Figure 1.4 The three physical states of water—water vapor, liquid water, and ice. We see the liquid and solid states but cannot see the gas (vapor) state. The red arrows show that the three states of matter interconvert.

Go Figure

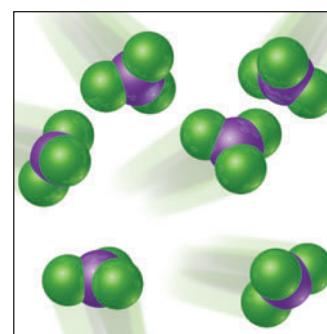
How do the molecules of a compound differ from the molecules of an element?



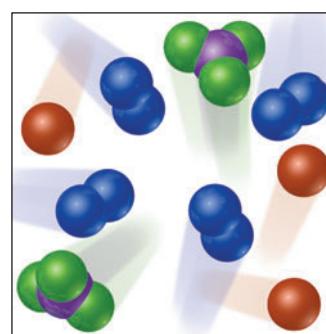
(a) Atoms of an element



(b) Molecules of an element



(c) Molecules of a compound



(d) Mixture of elements and a compound

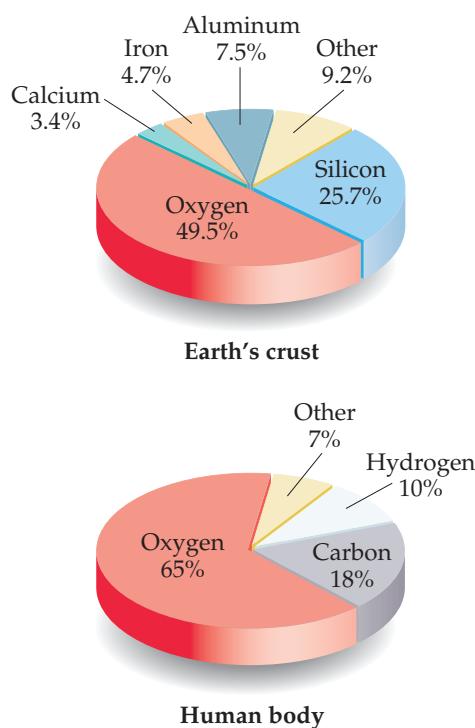
Elements are composed of only one kind of atom.

Compounds must have at least two kinds of atoms.

▲ Figure 1.5 Representation of elements, compounds, and mixtures.

Go Figure

If the lower pie chart was drawn as the percentage in terms of number of atoms rather than the percentage in terms of mass, would the hydrogen slice of the pie get larger or smaller?



▲ Figure 1.6 Relative abundances of elements.* Elements in percent by mass in Earth's crust (including oceans and atmosphere) and the human body.

Elements

Currently, 118 elements are known, though they vary widely in abundance. Hydrogen constitutes about 74% of the mass in the Milky Way galaxy, and helium constitutes 24%. Closer to home, only five elements—oxygen, silicon, aluminum, iron, and calcium—account for over 90% of Earth's crust (including oceans and atmosphere), and only three—oxygen, carbon, and hydrogen—account for over 90% of the mass of the human body (Figure 1.6).

Table 1.1 lists some common elements, along with the chemical *symbols* used to denote them. The symbol for each element consists of one or two letters, with the first letter capitalized. These symbols are derived mostly from the English names of the elements, but sometimes they are derived from a foreign name instead (last column in Table 1.1). You will need to know these symbols and learn others as we encounter them in the text.

All of the known elements and their symbols are listed on the front inside cover of this text in a table known as the *periodic table*. In the periodic table, the elements are arranged in columns so that closely related elements are grouped together. We describe the periodic table in more detail in Section 2.5 and consider the periodically repeating properties of the elements in Chapter 7.

Compounds

Most elements can interact with other elements to form compounds. For example, when hydrogen gas burns in oxygen gas, the elements hydrogen and oxygen combine to form the compound water. Conversely, water can be decomposed into its elements by passing an electrical current through it (Figure 1.7).

Decomposing pure water into its constituent elements shows that it contains 11% hydrogen and 89% oxygen by mass, regardless of its source. This ratio is constant because every water molecule has the same number of hydrogen and oxygen atoms. While the mass percentages make it seem that water is mostly oxygen, there are actually two hydrogen atoms and only one oxygen atom per molecule. The explanation for this apparent discrepancy comes from the fact that hydrogen atoms are much lighter than oxygen atoms. The mass composition, a macroscopic determination, corresponds to

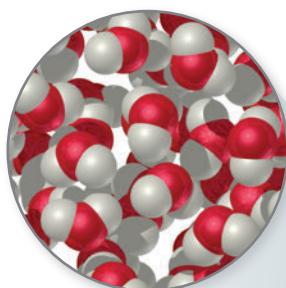
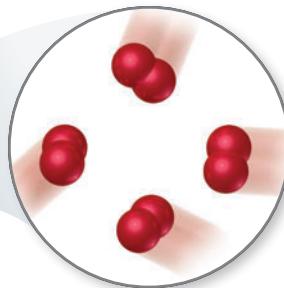
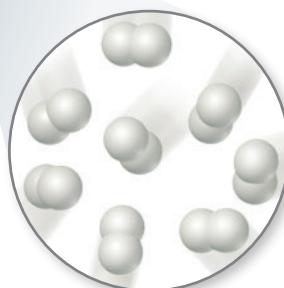
TABLE 1.1 Some Common Elements and Their Symbols

Carbon	C	Aluminum	Al	Copper	Cu (from <i>cuprum</i>)
Fluorine	F	Bromine	Br	Iron	Fe (from <i>ferrum</i>)
Hydrogen	H	Calcium	Ca	Lead	Pb (from <i>plumbum</i>)
Iodine	I	Chlorine	Cl	Mercury	Hg (from <i>hydrargyrum</i>)
Nitrogen	N	Helium	He	Potassium	K (from <i>kalium</i>)
Oxygen	O	Lithium	Li	Silver	Ag (from <i>argentum</i>)
Phosphorus	P	Magnesium	Mg	Sodium	Na (from <i>natrium</i>)
Sulfur	S	Silicon	Si	Tin	Sn (from <i>stannum</i>)

*U.S. Geological Survey Circular 285, U.S Department of the Interior.

**Go Figure**

Is the volume of H_2 produced larger than the volume of O_2 produced because (a) hydrogen atoms are lighter than oxygen atoms, (b) hydrogen atoms are larger than oxygen atoms, or (c) each water molecule contains one oxygen atom and two hydrogen atoms?

Water, H_2O Oxygen gas, O_2 Hydrogen gas, H_2

▲ Figure 1.7 Electrolysis of water. Water decomposes into its component elements, hydrogen and oxygen, when an electrical current is passed through it. The volume of hydrogen, collected in the right test tube, is twice the volume of oxygen.

the molecular composition, which consists of two hydrogen atoms combined with one oxygen atom:

Hydrogen atom
(written H)Oxygen atom
(written O)Water molecule
(written H_2O)

The elements hydrogen and oxygen themselves exist naturally as diatomic (two-atom) molecules:

Oxygen molecule

(written O_2)

Hydrogen molecule

(written H_2)

As seen in **Table 1.2**, the properties of water bear no resemblance to the properties of its component elements. Hydrogen, oxygen, and water are each a unique substance, a consequence of the uniqueness of their respective molecules.

TABLE 1.2 Comparison of Water, Hydrogen, and Oxygen

	Water	Hydrogen	Oxygen
State ^a	Liquid	Gas	Gas
Normal boiling point	100 °C	-253 °C	-183 °C
Density ^a	1000 kg/m ³	0.084 kg/m ³	1.33 kg/m ³
Flammable	No	Yes	No

^a At room temperature and atmospheric pressure.

The observation that the elemental composition of a compound is always the same is known as the **law of constant composition** (or the **law of definite proportions**). French chemist Joseph Louis Proust (1754–1826) first stated the law in about 1800. Although this law has been known for 200 years, the belief persists among some people that a fundamental difference exists between compounds prepared in the laboratory and the corresponding compounds found in nature. This simply is not true. Regardless of its source—nature or a laboratory—a pure compound has the same composition and properties under the same conditions. Both chemists and nature must use the same elements and operate under the same natural laws. When two materials differ in composition or properties, either they are composed of different compounds or they differ in purity.

Mixtures

Most of the matter we encounter consists of mixtures of different substances. Each substance in a mixture retains its chemical identity and properties. In contrast to a pure substance, which by definition has a fixed composition, the composition of a mixture can vary. A cup of sweetened coffee, for example, can contain either a little sugar or a lot. The substances making up a mixture are called *components* of the mixture.

Some mixtures do not have the same composition, properties, and appearance throughout. Rocks and wood, for example, vary in texture and appearance in any typical sample. Such mixtures are *heterogeneous* [Figure 1.8(a)]. Mixtures that are uniform throughout are *homogeneous*. Air is a homogeneous mixture of nitrogen, oxygen, and smaller amounts of other gases. The nitrogen in air has all the properties of pure nitrogen because both the pure substance and the mixture contain the same nitrogen molecules. Salt, sugar, and many other substances dissolve in water to form **solutions**, which are homogeneous mixtures [Figure 1.8(b)]. Although the term *solution* conjures an image of a liquid, the term can also apply to solids, liquids, or gases.

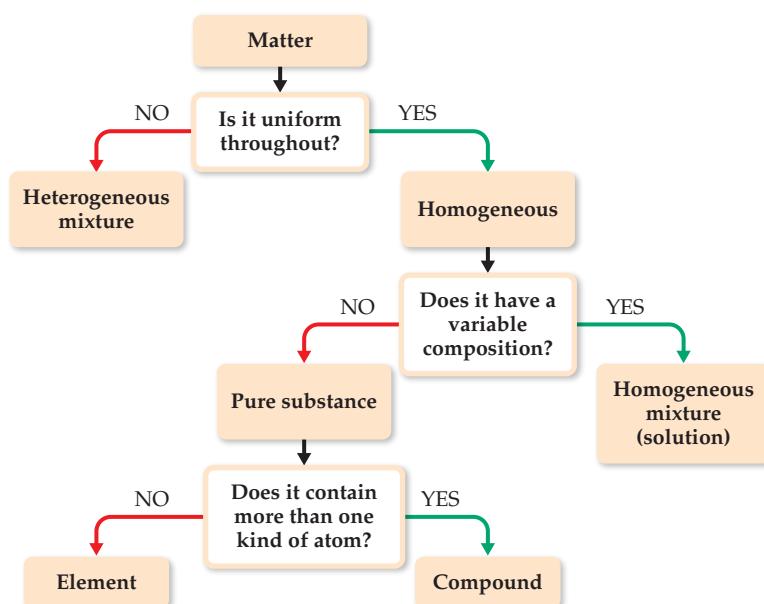
Figure 1.9 summarizes the classification of matter into elements, compounds, and mixtures.

► **Figure 1.8 Mixtures.** (a) Many common materials, including rocks, are heterogeneous mixtures. This photograph of granite shows a heterogeneous mixture of silicon dioxide and other metal oxides. (b) Homogeneous mixtures are called solutions. Many substances, including the blue solid shown here [copper(II) sulfate pentahydrate], dissolve in water to form solutions.



(a)

(b)



◀ **Figure 1.9 Classification of matter.** All pure matter is classified ultimately as either an element or a compound.

Sample Exercise 1.1

Distinguishing among Elements, Compounds, and Mixtures

"White gold" contains gold and a "white" metal, such as palladium. Two samples of white gold differ in the relative amounts of gold and palladium they contain. Both samples are uniform in composition throughout. Use Figure 1.9 to classify white gold.

SOLUTION

Because the material is uniform throughout, it is homogeneous. Because its composition differs for the two samples, it cannot be a compound. Instead, it must be a homogeneous mixture.

► Practice Exercise

Aspirin is composed of 60.0% carbon, 4.5% hydrogen, and 35.5% oxygen by mass, regardless of its source. Use Figure 1.9 to classify aspirin.

Self-Assessment Exercise

- 1.2** Carbonated water, 'Soda Water', is water in which carbon dioxide has been dissolved under pressure. Classify soda water.

- (a) A homogeneous solution of elements
- (b) A homogeneous solution of compounds
- (c) A heterogeneous mixture of elements
- (d) A heterogeneous mixture of compounds

Exercises

- 1.3** Classify each of the following as a pure substance or a mixture. If a mixture, indicate whether it is homogeneous or heterogeneous: (a) air, (b) chocolate with almond, (c) aluminum, (d) iodine tincture.
- 1.4** Give the chemical symbol or name for the following elements, as appropriate: (a) helium, (b) platinum, (c) cobalt, (d) tin, (e) silver, (f) Sb, (g) Pb, (h) Br, (i) V, (j) Hg.

- 1.5** A solid white substance A is heated strongly in the absence of air. It decomposes to form a new white substance B and a gas C. The gas has exactly the same properties as the product obtained when carbon is burned in an excess of oxygen. Based on these observations, can we determine whether solids A and B and gas C are elements or compounds?

1.2 (b)



1.3 | Properties of Matter



We are familiar with many of the properties of materials around us—the hardness of steel, the flexibility of rubber, or the glow of a wood fire. If we delve a little deeper, we can classify these properties as one of two types. By the end of this section, you should be able to

- Distinguish between physical and chemical changes

Every substance has unique properties. For example, the properties listed in Table 1.2 allow us to distinguish hydrogen, oxygen, and water from one another. The properties of matter can be categorized as physical or chemical. **Physical properties** can be observed without changing the identity and composition of the substance. These properties include color, odor, density, melting point, boiling point, and hardness. **Chemical properties** describe the way a substance may change, or *react*, to form other substances. A common chemical property is flammability, the ability of a substance to burn in the presence of oxygen.

Some properties, such as temperature and melting point, are *intensive properties*. **Intensive properties** do not depend on the amount of sample being examined and are particularly useful in chemistry because many intensive properties can be used to *identify* substances. **Extensive properties** depend on the amount of sample, with two examples being mass and volume. Extensive properties relate to the *amount* of substance present.

Physical and Chemical Changes

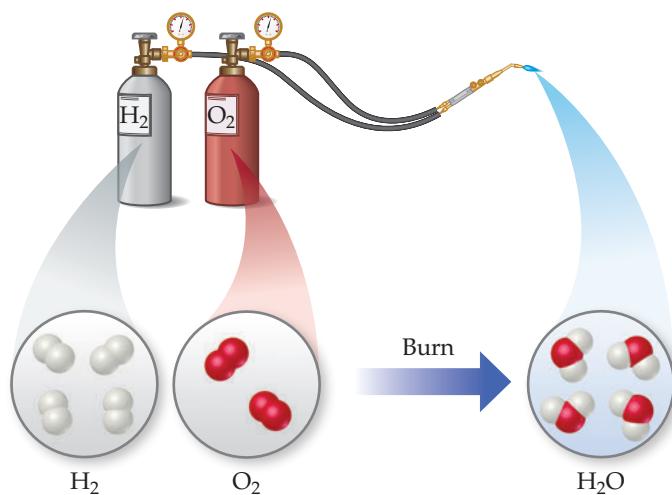
The changes substances undergo are either physical or chemical. During a **physical change**, a substance changes its physical appearance but not its composition. (That is, it is the same substance before and after the change.) The evaporation of water is a physical change. When water evaporates, it changes from the liquid state to the gas state, but it is still composed of water molecules, as depicted in Figure 1.4. All **changes of state** (for example, from liquid to gas or from liquid to solid) are physical changes.

In a **chemical change** (also called a **chemical reaction**), a substance is transformed into a chemically different substance. When hydrogen burns in air, for example, it undergoes a chemical change because it combines with oxygen to form water (Figure 1.10).

Chemical changes can be dramatic. In the account given in Figure 1.11, Ira Remsen, author of a popular chemistry text published in 1901, describes his first experiences with chemical reactions.

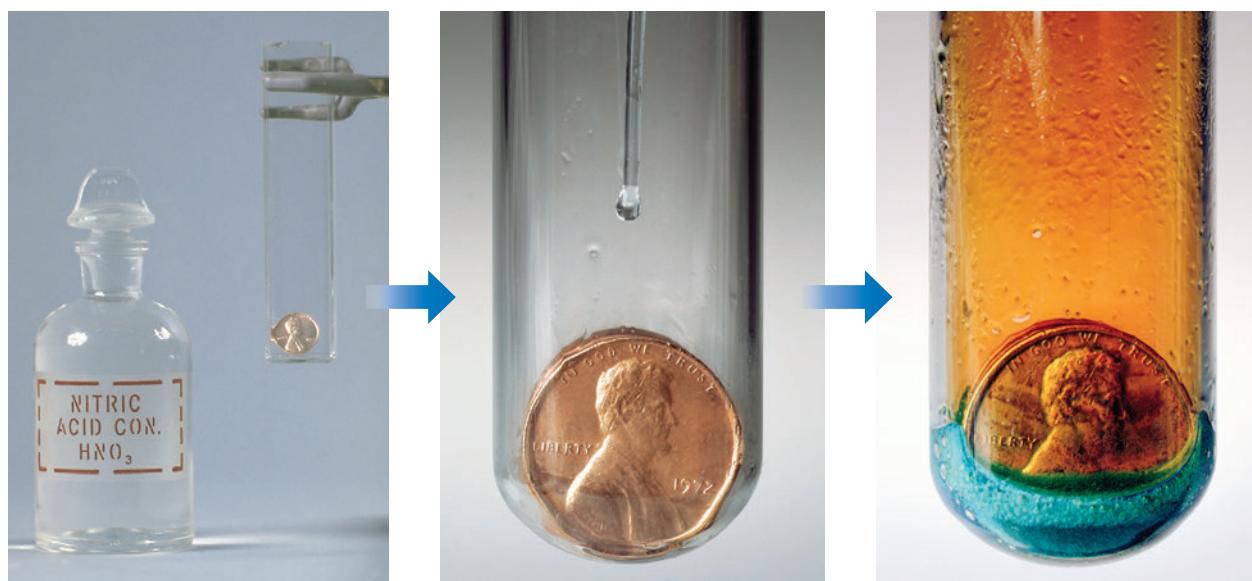
Separation of Mixtures

We can separate a mixture into its components by taking advantage of differences in their properties. For example, a heterogeneous mixture of iron filings and gold filings could be sorted by color into iron and gold. A less tedious approach would be to use a magnet to



▲ Figure 1.10 A chemical reaction.

While reading a textbook of chemistry, I came upon the statement “nitric acid acts upon copper,” and I determined to see what this meant. Having located some nitric acid, I had only to learn what the words “act upon” meant. In the interest of knowledge I was even willing to sacrifice one of the few copper cents then in my possession. I put one of them on the table, opened a bottle labeled “nitric acid,” poured some of the liquid on the copper, and prepared to make an observation. But what was this wonderful thing which I beheld? The cent was already changed, and it was no small change either. A greenish-blue liquid foamed and fumed over the cent and over the table. The air became colored dark red. How could I stop this? I tried by picking the cent up and throwing it out the window. I learned another fact: nitric acid acts upon fingers. The pain led to another unpremeditated experiment. I drew my fingers across my trousers and discovered nitric acid acts upon trousers. That was the most impressive experiment I have ever performed. I tell of it even now with interest. It was a revelation to me. Plainly the only way to learn about such remarkable kinds of action is to see the results, to experiment, to work in the laboratory.*



▲ Figure 1.11 The chemical reaction between a copper penny and nitric acid. The dissolved copper produces the blue-green solution; the reddish brown gas produced is nitrogen dioxide.

attract the iron filings, leaving the gold ones behind. We can also take advantage of an important chemical difference between these two metals: Many acids dissolve iron but not gold. Thus, if we put our mixture into an appropriate acid, the acid would dissolve the iron and the solid gold would be left behind. The two could then be separated by

*Remsen, *The Principles of Theoretical Chemistry*, 1887.

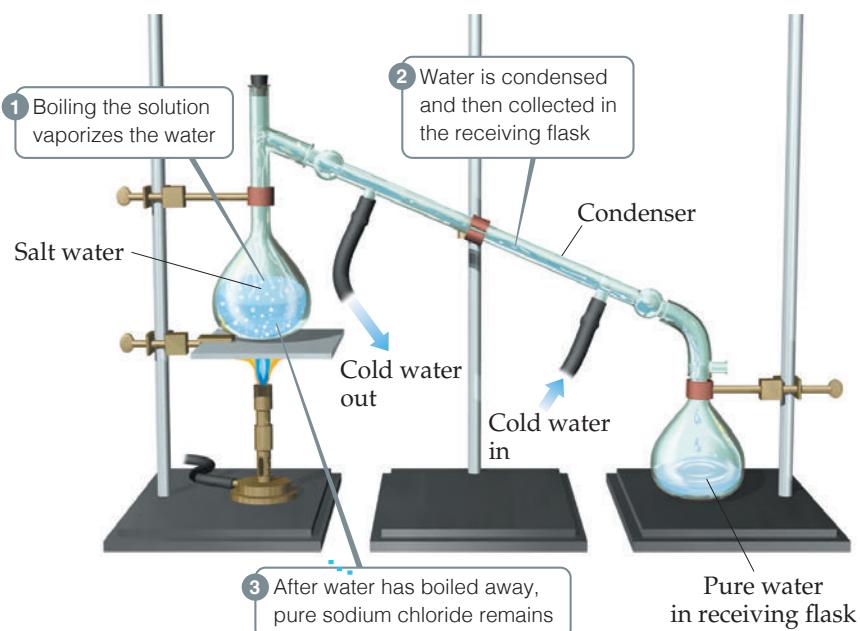
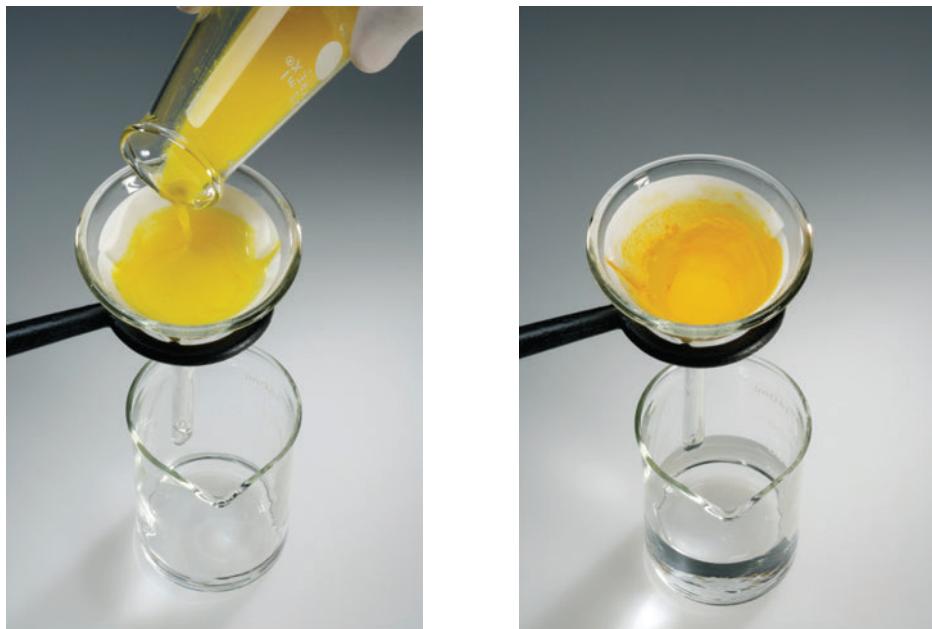
filtration (Figure 1.12). We would have to use other chemical reactions, which we will learn about later, to transform the dissolved iron back into metal.

An important method of separating the components of a homogeneous mixture is **distillation**, a process that depends on the different abilities of substances to form gases. For example, if we boil a solution of salt and water, the water evaporates, forming a gas, and the salt is left behind. The gaseous water can be converted back to a liquid on the walls of a condenser, as shown in Figure 1.13.

The differing abilities of substances to adhere to the surfaces of solids can also be used to separate mixtures. This ability is the basis of *chromatography*, a technique shown in Figure 1.14.

► **Figure 1.12 Separation by filtration.**

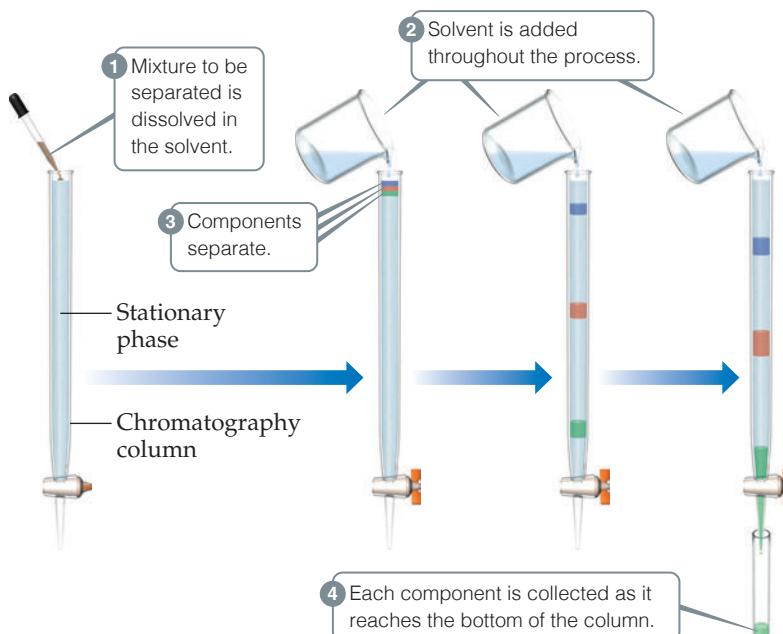
A mixture of a solid and a liquid is poured through filter paper. The liquid passes through the paper while the solid remains on the paper.



▲ **Figure 1.13 Distillation.** Apparatus for separating a sodium chloride solution (salt water) into its components.

Go Figure

Is the separation of the mixture shown here a physical or chemical process?



▲ Figure 1.14 Separation of three substances using column chromatography.

Self-Assessment Exercise

- 1.6** When water boils in a pot, bubbles of water vapor can be seen rising through the liquid. What type of change does this represent?

- (a) a physical change
(b) a chemical change

Exercises

- 1.7** Zirconia, an oxide of zirconium, is often used as an affordable diamond substitute. Just like diamond, it is a colorless crystal which sparkles under sunlight. Which of the following physical properties do you think would help in differentiating between diamond and Zirconia—melting point, density, or physical state?
1.8 In the process of attempting to characterize a substance, a chemist makes the following observations: The substance is a silvery white, lustrous metal. It melts at 649°C and boils at 1105°C . Its density at 20°C is 1.738 g/cm^3 . The substance burns in air, producing an intense white light. It reacts with

chlorine to give a brittle white solid. The substance can be pounded into thin sheets or drawn into wires. It is a good conductor of electricity. Which of these characteristics are physical properties, and which are chemical properties?

- 1.9** Label each of the following as either a physical process or a chemical process: (a) crushing a metal can, (b) production of urine in the kidneys, (c) melting a piece of chocolate, (d) burning fossil fuel, (e) discharging a battery.
1.10 Which separation method is better suited for obtaining sugar from cane juice—filtration or evaporation?

1.6 (a)

Answers to Self-Assessment Exercises

1.4 | The Nature of Energy



Since the earliest time, humans have recognized fire as a form of energy. In chemistry, it is convenient to distinguish between different types of energy, and, by the end of this section, you should be able to

- Recognize the difference between kinetic and potential energy

All objects in the universe are made of matter, but matter alone is not enough to describe the behavior of the world around us. The water in an alpine lake and a pot of boiling water are both made from the same substance, but your body will experience a very different sensation if you put your hand in each. The difference between the two is their energy content; boiling water has more energy than chilled water. To understand chemistry, we must also understand energy and the changes in energy that accompany chemical processes.

Unlike matter, energy does not have mass and cannot be held in our hands, but its effects can be observed and measured. **Energy** is defined as the capacity to do work or transfer heat. **Work** is the energy transferred when a force exerted on an object causes a displacement of that object, and **heat** is the energy used to cause the temperature of an object to increase (Figure 1.15). Although the temperature of an object is intuitive to most people, the definition of work is less apparent. We define work, w , as the product of the force exerted on the object, F , and the distance, d , that it moves:

$$w = F \times d \quad [1.1]$$

where **force** is defined as any push or pull exerted on the object.* Familiar examples include gravity and the attraction between opposite poles of a bar magnet. It takes work to lift an object off of the floor, or to pull apart two magnets that have come together at the opposite poles.

Kinetic Energy and Potential Energy

To understand energy, we need to grasp its two fundamental forms, kinetic energy and potential energy. Objects, whether they are automobiles, soccer balls, or molecules, can possess **kinetic energy**, the energy of motion. The magnitude of kinetic energy, E_k , of an object depends on its mass, m , and velocity, v :

$$E_k = \frac{1}{2} mv^2 \quad [1.2]$$

(a)



Heat added by burner to water makes water temperature rise

(b)

▲ **Figure 1.15** Work and heat, two forms of energy. (a) *Work* is energy used to cause an object to move against an opposing force. (b) *Heat* is energy used to increase the temperature of an object.

*In using this equation, only the component of the force that is acting parallel to the distance traveled is used. That will generally be the case for problems we will encounter in this chapter.

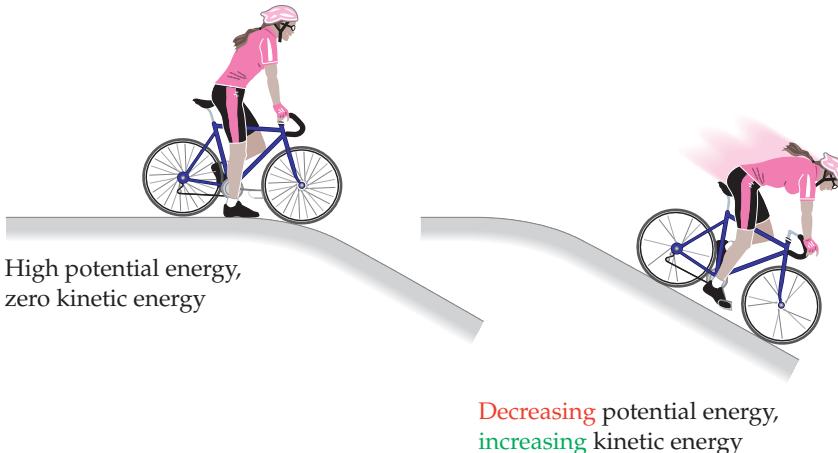
Thus, the kinetic energy of an object increases as its velocity or speed** increases. For example, a car has greater kinetic energy moving at 65 kilometers per hour (km/h) than it does at 25 km/h. For a given velocity, the kinetic energy increases with increasing mass. Thus, a large truck traveling at 65 km/h has greater kinetic energy than a motorcycle traveling at the same velocity because the truck has the greater mass.

In chemistry, we are interested in the kinetic energy of atoms and molecules. Although these particles are too small to be seen, they have mass and are in motion, and therefore, possess kinetic energy. When a substance is heated, be it a pot of water on the stove or an aluminum can sitting in the sun, the atoms and molecules in that substance gain kinetic energy and their average speed increases. Hence, we see that the transfer of heat is simply the transfer of kinetic energy at the molecular level.

All other forms of energy—the energy stored in a stretched spring, in a weight held above your head, or in a chemical bond—are classified as potential energy. An object has **potential energy** by virtue of its position relative to other objects. Potential energy is, in essence, the “stored” energy that arises from the attractions and repulsions an object experiences in relation to other objects.

We are familiar with many instances in which potential energy is converted into kinetic energy. For example, think of a cyclist poised at the top of a hill (**Figure 1.16**). Because of the attractive force of gravity, the potential energy of the bicycle is greater at the top of the hill than at the bottom. As a result, the bicycle easily rolls down the hill with increasing speed. As it does so, potential energy is converted into kinetic energy. The potential energy decreases as the bicycle rolls down the hill, while at the same time its kinetic energy increases as it picks up speed (Equation 1.2). This example illustrates that kinetic and potential energy are interconvertible.

Gravitational forces play a negligible role in the ways that atoms and molecules interact with one another. Forces that arise from electrical charges are more important when dealing with atoms and molecules. One of the most important forms of potential energy in chemistry is *electrostatic potential energy*, which arises from the interactions between charged particles. You are probably familiar with the fact that opposite charges attract each other and like charges repel. The strength of this interaction increases as the magnitude of the charges increase, and decreases as the



▲ Figure 1.16 Potential energy and kinetic energy. The potential energy initially stored in the motionless bicycle and rider at the top of the hill is converted to kinetic energy as the bicycle moves down the hill and loses potential energy.

**Strictly speaking, velocity is a vector quantity that has a direction; that is, it tells you how fast an object is moving and in what direction. Speed is a scalar quantity that tells you how fast an object is moving but not the direction of the motion. Unless otherwise stated, we will not be concerned with the direction of motion, and thus velocity and speed are interchangeable quantities for the treatment in this book.

distance between charges increases. We will return to electrostatic energy several times throughout the book.

One of our goals in chemistry is to relate the energy changes seen in the macroscopic world to the kinetic or potential energy of substances at the molecular level. Many substances, fuels for example, release energy when they react. The *chemical energy* of a fuel is due to the potential energy stored in the arrangements of its atoms. As we will learn in later chapters, *chemical energy is released when bonds between atoms are formed, and consumed when bonds between atoms are broken*. When a fuel burns, some bonds are broken and others are formed, but the net effect is to convert chemical potential energy to thermal energy, the energy associated with temperature. The increase in thermal energy arises from increased molecular motion and hence increased kinetic energy at the molecular level.

Self-Assessment Exercise

- 1.11** What happens to the chemical energy of a nickel-metal hydride battery as it is recharged after powering an electric razor?

- (a) It increases
(b) It decreases
(c) It remains the same

Exercises

- 1.12** Two positively charged particles are first brought close together and then released. Once released, the repulsion between particles causes them to move away from each other. (a) This is an example of potential energy being converted into what form of energy? (b) Does the potential energy of the two particles prior to release increase or decrease as the distance between them is increased.

- 1.13** For each of the following processes, does the potential energy of the object(s) increase or decrease? (a) The charge of two oppositely charged particles is increased. (b) H_2O molecule is split into two oppositely charged ions, H^+ and OH^- . (c) A person skydives from a height of 600 meters.

1.11 (a)

Answers to Self-Assessment Exercises

1.5 | Units of Measurement



Many properties of matter are *quantitative*, that is, associated with numbers. When a number represents a measured quantity, the units of that quantity must be specified. The volume of a soft drink expressed in different units demonstrates how important this is. To say that the length of a pencil is 17.5 is meaningless. Expressing the number with its units, 17.5 centimeters (cm), properly specifies the length. The units used for scientific measurements are those of the **metric system**.

The metric system, developed in France during the late eighteenth century, is used as the system of measurement in most countries. By the end of this section, you should be able to

- Convert between different multiples of SI units

SI Units

In 1960, an international agreement was reached specifying a particular choice of metric units for use in scientific measurements. These preferred units are called **SI units**, after the French *Système International d'Unités*. This system has seven *base units* from which all other units are derived (Table 1.3). In this section, we will consider the base units for length, mass, and temperature.

With SI units, prefixes are used to indicate decimal fractions or multiples of various units. For example, the prefix *milli-* represents a 10^{-3} fraction, one-thousandth, of a unit: A milligram (mg) is 10^{-3} gram (g), a millimeter (mm) is 10^{-3} meter (m), and so forth.

TABLE 1.3 SI Base Units

Physical Quantity	Name of Unit	Abbreviation
Length	Meter	m
Mass	Kilogram	kg
Temperature	Kelvin	K
Time	Second	s or sec
Amount of substance	Mole	mol
Electric current	Ampere	A or amp
Luminous intensity	Candela	cd

A CLOSER LOOK The Scientific Method

Where does scientific knowledge come from? How is it acquired? How do we know it is reliable? How do scientists add to it, or modify it?

There is nothing mysterious about how scientists work. The first idea to keep in mind is that scientific knowledge is gained through observations of the natural world. A principal aim of the scientist is to organize these observations by identifying patterns and regularity, making measurements, and associating one set of observations with another. The next step is to ask *why* nature behaves in the manner we observe. To answer this question, the scientist constructs a model, known as a **hypothesis**, to explain the observations. Initially, the hypothesis is likely to be pretty tentative. There could be more than one reasonable hypothesis. If a hypothesis is correct, then certain results and observations should follow from it. In this way, hypotheses can stimulate the design of experiments to learn more about the system being studied. Scientific creativity comes into play in thinking of hypotheses that are fruitful in suggesting good experiments to do, ones that will shed new light on the nature of the system.

As more information is gathered, the initial hypotheses get winnowed down. Eventually, just one may stand out as most consistent with a body of accumulated evidence. We then begin to call this hypothesis a **theory**, a model that has predictive powers and that accounts for all the available observations. A theory also generally is consistent with other, perhaps larger and more general theories. For example, a theory of what goes on inside a volcano has to be consistent with more general theories regarding heat transfer, chemistry at high temperature, and so forth.

We will be encountering many theories as we proceed through this book. Some of them have been found over and over again to be consistent with observations. However, no theory can be proven to be absolutely true. We can treat it as though it is, but there always remains a possibility that there is some respect in which a theory is wrong. A famous example is Isaac Newton's theory of mechanics, which yielded such precise results for the mechanical behavior of matter that no exceptions to it were found before the twentieth century. But Albert Einstein showed that Newton's theory of the nature of space and time is incorrect.

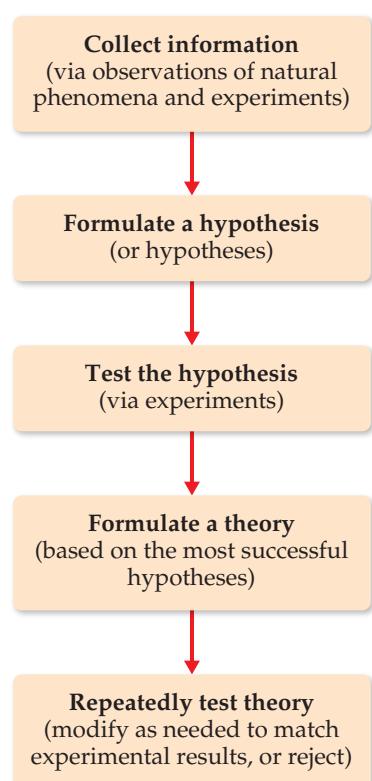
Continued

Einstein's theory of relativity represented a fundamental shift in how we think of space and time. He predicted where the exceptions to predictions based on Newton's theory might be found. Although only small departures from Newton's theory were predicted, they *were* observed. Einstein's theory of relativity became accepted as the correct model. However, for most uses, Newton's laws of motion are quite accurate enough.

The overall process we have just considered, illustrated in **Figure 1.17**, is often referred to as *the scientific method*. But there is no single scientific method. Many factors play a role in advancing scientific knowledge. The one unvarying requirement is that our explanations be consistent with observations and that they depend solely on natural phenomena.

When nature behaves in a certain way over and over again, under all sorts of different conditions, we can summarize that behavior in a **scientific law**. For example, it has been repeatedly observed that in a chemical reaction there is no change in the total mass of the materials reacting as compared with the materials that are formed; we call this observation the *law of conservation of mass*. It is important to make a distinction between a theory and a scientific law. On the one hand, a scientific law is a statement of what always happens, to the best of our knowledge. A theory, on the other hand, is an *explanation* for what happens. If we discover some law fails to hold true, then we must assume the theory underlying that law is wrong in some way.

Related Exercises: 1.74, 1.96



▲ **Figure 1.17** The scientific method.

Table 1.4 presents the prefixes commonly encountered in chemistry. In using SI units and in working problems throughout this text, you must be comfortable using exponential notation. If you are unfamiliar with exponential notation or want to review it, refer to Appendix A.1.

TABLE 1.4 Prefixes Used in the Metric System and with SI Units

Prefix	Abbreviation	Meaning	Example	= $1 \times 10^{\text{exponent}}$ watt
Peta	P	10^{15}	1 petawatt (PW)	= 1×10^{15} watts ^a
Tera	T	10^{12}	1 terawatt (TW)	= 1×10^{12} watts
Giga	G	10^9	1 gigawatt (GW)	= 1×10^9 watts
Mega	M	10^6	1 megawatt (MW)	= 1×10^6 watts
Kilo	k	10^3	1 kilowatt (kW)	= 1×10^3 watts
Deci	d	10^{-1}	1 deciwatt (dW)	= 1×10^{-1} watt
Centi	c	10^{-2}	1 centiwatt (cW)	= 1×10^{-2} watt
Milli	m	10^{-3}	1 milliwatt (mW)	= 1×10^{-3} watt
Micro	μ ^b	10^{-6}	1 microwatt (μ W)	= 1×10^{-6} watt
Nano	n	10^{-9}	1 nanowatt (nW)	= 1×10^{-9} watt
Pico	p	10^{-12}	1 picowatt (pW)	= 1×10^{-12} watt
Femto	f	10^{-15}	1 femtowatt (fW)	= 1×10^{-15} watt
Atto	a	10^{-18}	1 attowatt (aW)	= 1×10^{-18} watt
Zepto	z	10^{-21}	1 zeptowatt (zW)	= 1×10^{-21} watt

^aThe watt (W) is the SI unit of power, which is the rate at which energy is either generated or consumed.

The SI unit of energy is the joule (J); $1 \text{ J} = 1 \text{ kg} \cdot \text{m}^2/\text{s}^2$ and $1 \text{ W} = 1 \text{ J/s}$.

^bGreek letter mu, pronounced "mew."

Although non-SI units are being phased out, some are still commonly used by scientists. Whenever we first encounter a non-SI unit in the text, the SI unit will also be given. The relations between the non-SI and SI units we will use most frequently in this text appear on the back inside cover. We will discuss how to convert from one to the other in Section 1.7.

Length and Mass

The SI base unit of *length* is the meter, a distance slightly longer than a yard. **Mass*** is a measure of the amount of material in an object. The SI base unit of mass is the kilogram (kg). This base unit is unusual because it uses a prefix, *kilo-*, instead of the word *gram* alone. We obtain other units for mass by adding prefixes to the word *gram*.



Sample Exercise 1.2

Using SI Prefixes



What is the name of the unit that equals (a) 10^{-9} gram, (b) 10^{-6} second, (c) 10^{-3} meter?

SOLUTION

We can find the prefix related to each power of ten in Table 1.4:
(a) nanogram, ng; **(b)** microsecond, μs ; **(c)** millimeter, mm.

► Practice Exercise

- (a)** How many picometers are there in 1 m? **(b)** Express 6.0×10^3 m using a prefix to replace the power of ten.
- (c)** Use exponential notation to express 4.22 mg in grams.
- (d)** Use decimal notation to express 4.22 mg in grams.

Temperature

Temperature, a measure of the hotness or coldness of an object, is a physical property that determines the direction of heat flow. Heat always flows spontaneously from a substance at higher temperature to one at lower temperature. Thus, the influx of heat we feel when we touch a hot object tells us that the object is at a higher temperature than our hand.

The temperature scales commonly employed in science are the Celsius and Kelvin scales. The **Celsius scale** was originally based on the assignment of 0 °C to the freezing point of water and 100 °C to its boiling point at sea level (*Figure 1.18*).

The **Kelvin scale** is the SI temperature scale, and the SI unit of temperature is the *kelvin* (K). Zero on the Kelvin scale is the temperature at which all thermal motion ceases, a temperature referred to as **absolute zero**. On the Celsius scale, absolute zero has the value –273.15 °C. The Celsius and Kelvin scales have equal-sized units—that is, a kelvin is the same size as a degree Celsius. Thus, the Kelvin and Celsius scales are related according to

$$K = ^\circ C + 273.15 \quad [1.3]$$

The freezing point of water, 0 °C, is 273.15 K (*Figure 1.18*). Notice that we do not use a degree sign (°) with temperatures on the Kelvin scale.

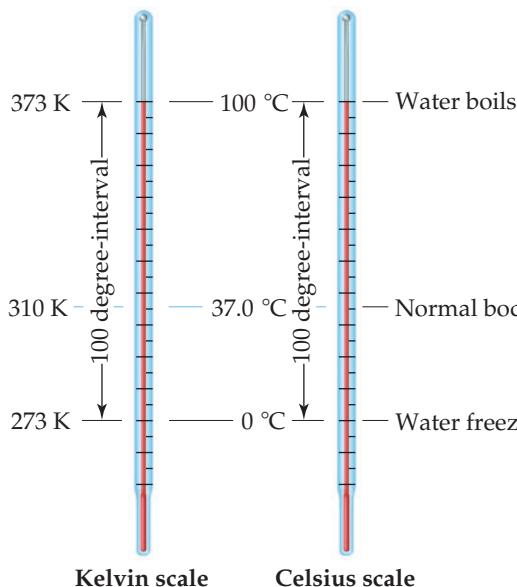
Derived SI Units

The SI base units are used to formulate *derived units*. A **derived unit** is obtained by multiplication or division of one or more of the base units. We begin with the defining

*Mass and weight are not the same. Mass is a measure of the amount of matter; weight is the force exerted on this mass by gravity. For example, an astronaut weighs less on the Moon than on Earth because the Moon's gravitational force is less than Earth's. The astronaut's mass on the Moon, however, is the same as it is on Earth.

**Go Figure**

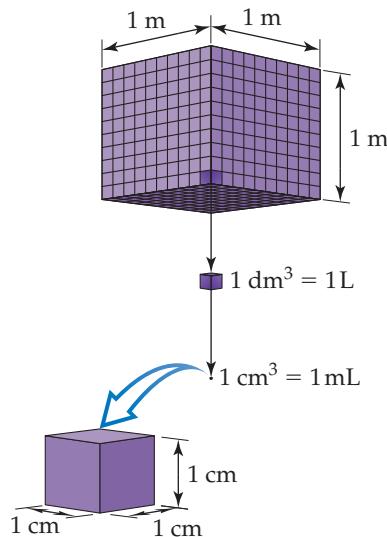
True or false: The “size” of a degree on the Celsius scale is the same as the “size” of a degree on the Kelvin scale.



▲ Figure 1.18 Comparison of the Kelvin and Celsius temperature scales.



How many 1-L bottles are required to contain 1 m^3 of liquid?



▲ Figure 1.19 Volume relationships. The volume occupied by a cube 1 m on each edge is one cubic meter, 1 m^3 . Each cubic meter contains 1000 dm^3 , $1\text{ m}^3 = 1000\text{ dm}^3$. One liter is the same volume as one cubic decimeter, $1\text{ L} = 1\text{ dm}^3$. Each cubic decimeter contains 1000 cubic centimeters, $1\text{ dm}^3 = 1000\text{ cm}^3$. One cubic centimeter equals one milliliter, $1\text{ cm}^3 = 1\text{ mL}$.

equation for a quantity and, then substitute the appropriate base units. For example, *speed* is defined as the ratio of distance traveled to elapsed time. Thus, the derived SI unit for speed is the SI unit for distance (length), m, divided by the SI unit for time, s, which gives m/s , read “meters per second.” Two common derived units in chemistry are those for volume and density.

Volume

The *volume* of a cube is its length cubed, length^3 . Thus, the derived SI unit of volume is the SI unit of length, m, raised to the third power. The cubic meter, m^3 , is the volume of a cube that is 1 m on each edge (Figure 1.19). Smaller units, such as cubic centimeters, cm^3 (sometimes written cc), are frequently used in chemistry. Another volume unit used in chemistry is the *liter* (L), which equals a cubic decimeter, dm^3 . (The liter is the first metric unit we have encountered that is *not* an SI unit.) There are 1000 milliliters (mL) in a liter, and 1 mL is the same volume as 1 cm^3 : $1\text{ mL} = 1\text{ cm}^3$.



Sample Exercise 1.3

Converting Units of Temperature

A weather forecaster predicts the temperature will reach $31\text{ }^\circ\text{C}$. What is this temperature in K?

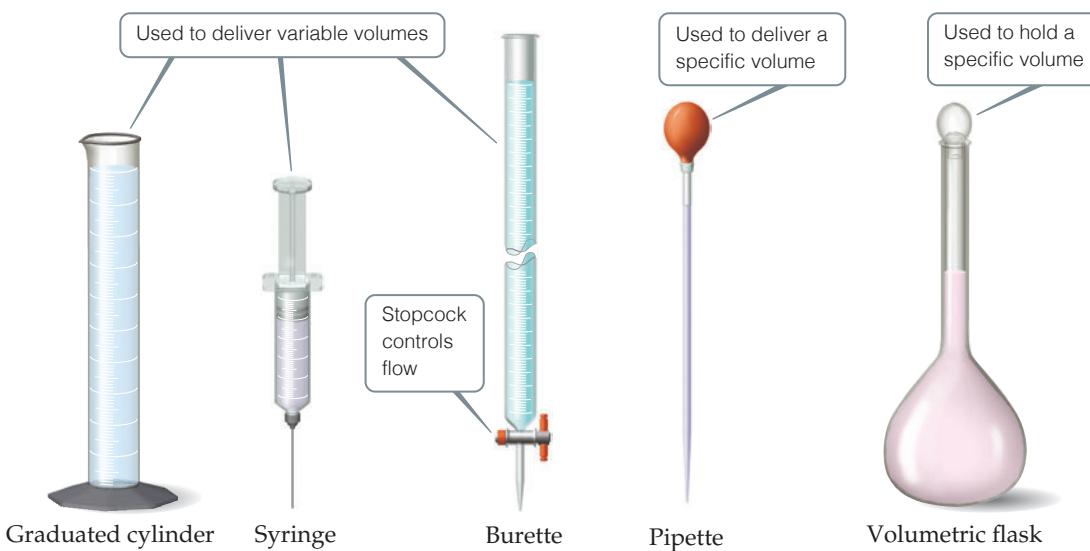
SOLUTION

Equation 1.3, we have $K = 31 + 273 = 304\text{ K}$.

► Practice Exercise

Using Wolfram Alpha (<http://www.wolframalpha.com/>) or some other reference, determine which of these elements

would be liquid at 525 K (assume samples are protected from air): (a) bismuth, Bi; (b) platinum, Pt; (c) selenium, Se; (d) calcium, Ca; (e) copper, Cu.



▲ Figure 1.20 Common volumetric glassware.

In the lab, you will likely use the devices in **Figure 1.20** to measure and deliver volumes of liquids. Syringes, burettes, and pipettes deliver amounts of liquids with more precision than graduated cylinders. Volumetric flasks are used to contain specific volumes of liquid.

Density

Density is defined as the amount of mass in a unit volume of a substance:

$$\text{density} = \frac{\text{mass}}{\text{volume}} \quad [1.4]$$

The densities of solids and liquids are commonly expressed in either grams per cubic centimeter (g/cm^3) or grams per milliliter (g/mL). The densities of some common substances are listed in **Table 1.5**. It is no coincidence that the density of water is 1.00 g/mL ; the gram was originally defined as the mass of 1 mL of water at a specific temperature. Because most substances change volume when they are heated or cooled, densities are temperature dependent, and so temperature should be specified when reporting densities. If no temperature is reported, we assume 25°C , close to normal room temperature.

The terms *density* and *weight* are sometimes confused. A person who says that iron weighs more than air generally means that iron has a higher density than air— 1 kg of air has the same mass as 1 kg of iron, but the iron occupies a smaller volume, thereby giving it a higher density. If we combine two liquids that do not mix, the less dense liquid will float on the denser liquid.

Units of Energy

The SI unit for energy is the **joule** (pronounced “jool”), J, in honor of James Joule (1818–1889), a British scientist who investigated work and heat. If we return to Equation 1.2 where kinetic energy was defined, we immediately see that joules are a derived unit, $1 \text{ J} = 1 \text{ kg}\cdot\text{m}^2/\text{s}^2$. Numerically, a 2-kg mass moving at a velocity of 1 m/s possesses a kinetic energy of 1 J :

$$E_k = \frac{1}{2}mv^2 = \frac{1}{2}(2 \text{ kg})(1 \text{ m/s})^2 = 1 \text{ kg}\cdot\text{m}^2/\text{s}^2 = 1 \text{ J}$$

TABLE 1.5 Densities of Selected Substances at 25°C

Substance	Density (g/cm^3)
Air	0.001
Balsa wood	0.16
Ethanol	0.79
Water	1.00
Ethylene glycol	1.09
Table sugar	1.59
Table salt	2.16
Iron	7.9
Gold	19.32



Sample Exercise 1.4

Determining Density and Using Density to Determine Volume or Mass

- (a) Calculate the density of mercury if 1.00×10^2 g occupies a volume of 7.36 cm^3 .
 (b) Calculate the volume of 65.0 g of liquid methanol if its density is 0.791 g/mL.
 (c) What is the mass in grams of a cube of gold (density = 19.32 g/cm³) if the length of the cube is 2.00 cm?

SOLUTION

- (a) We are given mass and volume, so Equation 1.3 yields

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{1.00 \times 10^2 \text{ g}}{7.36 \text{ cm}^3} = 13.6 \text{ g/cm}^3$$

- (b) Solving Equation 1.3 for volume and then using the given mass and density gives

$$\text{Volume} = \frac{\text{mass}}{\text{density}} = \frac{65.0 \text{ g}}{0.791 \text{ g/mL}} = 82.2 \text{ mL}$$

- (c) We can calculate the mass from the volume of the cube and its density. The volume of a cube is given by its length cubed:

$$\text{Volume} = (2.00 \text{ cm})^3 = (2.00)^3 \text{ cm}^3 = 8.00 \text{ cm}^3$$

Solving Equation 1.3 for mass and substituting the volume and density of the cube, we have

$$\text{Mass} = \text{volume} \times \text{density} = (8.00 \text{ cm}^3)(19.32 \text{ g/cm}^3) = 155 \text{ g}$$

► Practice Exercise

- (a) Calculate the density of a 374.5-g sample of copper if it has a volume of 41.8 cm³. (b) A student needs 15.0 g of ethanol for an experiment. If the density of ethanol is 0.789 g/mL, how many milliliters of ethanol are needed? (c) What is the mass, in grams, of 25.0 mL of mercury (density = 13.6 g/mL)?

Because a joule is not a very large amount of energy, we often use *kilojoules* (kJ) in discussing the energies associated with chemical reactions. For example, the amount of heat released when hydrogen and oxygen react to form 1 g of water is 16 kJ.

It is still quite common in chemistry, biology, and biochemistry to find energy changes associated with chemical reactions expressed in the non-SI unit of calories. A **calorie** (cal) was originally defined as the amount of energy required to raise the temperature of 1 g of water from 14.5 to 15.5 °C. It has since been defined in terms of a joule:

$$1 \text{ cal} = 4.184 \text{ J} \text{ (exactly)}$$

A related energy unit that is familiar to anyone who has read a food label is the nutritional *Calorie* (note the capital C), which is 1000 times larger than calorie with a lowercase c: 1 Cal = 1000 cal = 1 kcal.



Sample Exercise 1.5

Identifying and Calculating Energy Changes

- A standard propane (C_3H_8) tank used in an outdoor grill holds approximately 9.0 kg of propane. When the grill is operating, propane reacts with oxygen to form carbon dioxide and water. For every gram of propane that reacts with oxygen in this way, 46 kJ of energy is released as heat. (a) How much energy is released if the entire contents of the propane tank react with oxygen? (b) As the propane reacts, does the potential energy stored in chemical bonds increase or decrease? (c) If you were to store an equivalent amount of potential energy by pumping water to an elevation of 75 m above the ground, what mass of water would be needed? (Note: The force due to gravity acting on the water, which is the water's weight, is $F = m \times g$, where m is the mass of the object and g is the gravitational constant, $g = 9.8 \text{ m/s}^2$.)

SOLUTION

- (a) We can calculate the amount of energy released from the propane as heat by converting the mass of propane from kg to g and then using the fact that 46 kJ of heat are released per gram:

$$E = 9.0 \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} \times \frac{46 \text{ kJ}}{1 \text{ g}} = 4.1 \times 10^5 \text{ kJ} = 4.1 \times 10^8 \text{ J}$$

- (b) When propane reacts with oxygen, the potential energy stored in the chemical bonds is converted to an alternate form of energy, heat. Therefore, the potential energy stored as chemical energy must decrease.
 (c) The amount of work done to pump the water to a height of 75 m can be calculated using Equation 1.1:

$$w = F \times d = (m \times g) \times d$$

rearranging to solve for the mass of water:

$$m = \frac{w}{g \times d} = \frac{4.1 \times 10^8 \text{ J}}{(9.8 \text{ m/s}^2)(75 \text{ m})} = \frac{4.1 \times 10^8 \text{ kg} \cdot \text{m}^2/\text{s}^2}{(9.8 \text{ m/s}^2)(75 \text{ m})}$$

$$= 5.6 \times 10^5 \text{ kg}$$

At 25 °C, this mass of water would have a volume of 560,000 L. Thus, we see that large amounts of potential energy can be stored as chemical energy.

► Practice Exercise

Which of the following objects has the greatest kinetic energy?

- (a) a 500-kg motorcycle moving at 100 km/h
- (b) a 1,000-kg car moving at 50 km/h
- (c) a 1500-kg car moving at 30 km/h
- (d) a 5000 kg truck moving at 10 km/h
- (e) a 10,000-kg truck moving at 5 km/h

CHEMISTRY PUT TO WORK

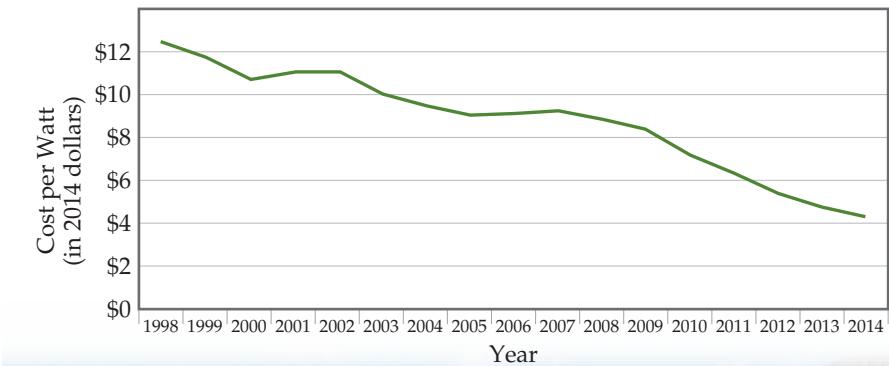
Chemistry in the News

Solar Energy Steps Up. Given the challenges society faces in trying to mitigate the effects of climate change, the need for affordable clean energy has never been greater. Given the massive amount of energy our planet receives from the Sun, solar energy has long been touted as a technology of the future. The problem for decades has been the relatively high cost of solar power compared to energy generated from burning fossil fuels. However, in recent years the cost of solar energy has decreased more rapidly than most people thought possible, decreasing more than 50% in just the past five years (**Figure 1.21**).

Not surprisingly, over the last six years the number of solar panel installations worldwide has increased sixfold. Over the same period, the number of coal plants operating in the United States has decreased by 38%, from 523 to 323. China is on track to install plants capable of generating more than 18 gigawatts of solar energy in 2015, nearly equal to the entire solar energy capacity of the United States. Furthermore, with the price of solar energy dropping rapidly, there is reason to believe that developing nations may forgo building fossil fuel power plants and skip straight to green energy technologies like solar and wind.

Recently, chemists have discovered a new class of materials called halide perovskites that have the potential to bring the cost of solar energy down even further. Solar cells made with halide perovskites have been shown to be nearly as efficient as single crystal silicon solar cells, but can be prepared from inexpensive solution methods that differ from the more costly and energy-intensive methods used to produce silicon solar cells. There are still many challenges to be overcome before halide perovskite solar cells are produced commercially, but their future looks very promising.

Slowing a Progressive Disease. Proteins are very large molecules that play an essential role in biology and the functioning of living organisms. Out of the myriad different types of proteins, scientists have identified one class of proteins called *prions* that play an important role in certain neurological diseases. We all have prion proteins in our brains, and for nearly all of us, they cause no harmful effects. For a small fraction of people, though, something causes the prions to change form and to adopt an incorrectly folded molecular shape. This process, once begun, is cumulative, propagating throughout the brain; the misfolded proteins somehow trigger the same misfolding in other prion proteins. Eventually, the misfolded proteins aggregate into clusters that can destroy neurons, producing symptoms such as those seen in Alzheimer's and Parkinson's diseases.

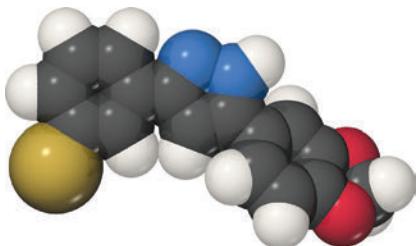
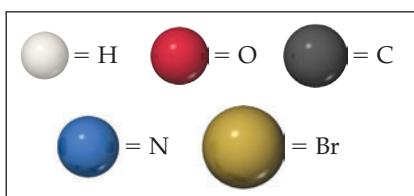


▲ **Figure 1.21** The cost of energy generated by photovoltaic modules. The median price per watt of electricity generated has dropped by two-thirds since 1998. The cost figures here include the cost for fully installed residential solar panels.

Currently, no therapies have been developed to stop the progression of prion diseases. However, there are encouraging signs that certain small molecules could disrupt propagation of the disease. They might be able to do this by blocking the interaction between one prion protein and another that causes the misfolding to propagate. So far, the experimental work has involved studies of mice infected with prions. One such molecule, called Anle-138b, is shown in **Figure 1.22**. Administration of this compound more than doubled the life spans of treated mice.

Compounds such as this one are not suitable for use in treating prion disease in humans, but studies conducted thus far point the way toward discovery of molecules that might in the future provide

Continued



▲ Figure 1.22 Anle-138b, a promising molecule for the treatment of degenerative brain diseases.

effective therapies. This work illustrates the complex and sometimes tortuous path that takes the scientist from promising beginnings to a much desired goal.

Chemical Approaches to Art Restoration. Works of art are made from materials that over time can undergo chemical reactions that alter their appearance or undermine their mechanical stability. Layers of dirt and pollutants can build up on the surface, discoloring statues, murals, and paintings. Plasters that serve as the base for frescoes react with gases in the atmosphere. One approach to art restoration is to use selected chemical reactions to undo or reverse the effects of detrimental reactions that have occurred over time.

One such example involves murals at the ancient Mayan ruins of Mayapan in Mexico's Yucatán peninsula. In the 1960s, polymer coatings were applied to these murals in an attempt to preserve them, but within a decade it became clear that the polymer coatings were doing more harm than good. Unfortunately, years of oxidation and cross-linking reactions had rendered the polymer coatings insoluble to nearly every type of organic solvent. In 2008, a team of Italian chemists found a way to remove the polymer coatings without damaging the underlying fresco. They treated the murals with a microemulsion, a special type of mixture where molecules called surfactants encapsulate and carry nanometer-sized droplets of organic solvents through an aqueous solution to the surface of the fresco. Once the droplets reach the surface, they can dissolve and carry away the unwanted polymer coating.

Self-Assessment Exercise

- 1.14** A standard Olympic-size swimming pool contains 2.5 ML of water. What volume is this in m^3 ?

- (a) 2.5 m^3
(b) 250 m^3

- (c) $2\ 500 \text{ m}^3$
(d) $2\ 500\ 000 \text{ m}^3$

Exercises

- 1.15** Convert the following expressions into exponential notation:
(a) 3 terameters (Tm) (b) 2.5 femtoseconds (fs) (c) 57 micrometers (μm) (d) 8.3 megagrams (Mg).

- 1.16** Make the following conversions: (a) 25°C to K (b) 110°C to K
(c) 300 K to $^\circ\text{C}$ (d) 1100 K to $^\circ\text{C}$

- 1.17** (a) A sample of tetrachloroethene, a liquid used in dry cleaning that is being phased out because of its potential to cause cancer, has a mass of 40.55 g and a volume of 25.0 mL at 25°C . What is its density at this temperature? Will tetrachloroethene float on water? (Materials that are less dense than water will float.)
(b) Carbon dioxide (CO_2) is a gas at room temperature and pressure. However, carbon dioxide can be put under pressure to become a "supercritical fluid" that is a much safer dry-cleaning agent than tetrachloroethylene. At a certain pressure, the density of supercritical CO_2 is 0.469 g/cm^3 . What is the mass of a 25.0-mL sample of supercritical CO_2 at this pressure?

- 1.18** (a) To identify a liquid substance, a student determined its density. Using a graduated cylinder, she measured out

a 45-mL sample of the substance. She then measured the mass of the sample, finding that it weighed 38.5 g. She knew that the substance had to be either isopropyl alcohol (density 0.785 g/mL) or toluene (density 0.866 g/mL). What are the calculated density and the probable identity of the substance? (b) An experiment requires 45.0 g of ethylene glycol, a liquid whose density is 1.114 g/mL . Rather than weigh the sample on a balance, a chemist chooses to dispense the liquid using a graduated cylinder. What volume of the liquid should he use? (c) Is a graduated cylinder such as that shown in Figure 1.20 likely to afford the accuracy of measurement needed? (d) A cubic piece of metal measures 5.00 cm on each edge. If the metal is nickel, whose density is 8.90 g/cm^3 , what is the mass of the cube?

- 1.19** If on a certain year, an estimated amount of 4 million metric tons (1 metric ton = 1000 kg) of nitrous oxide (N_2O) was emitted worldwide due to agricultural activities, express this mass of N_2O in grams without exponential notation, using an appropriate metric prefix.

1.14 (c)



1.6 | Uncertainty in Measurement



Weighing a single postage stamp on a pair of old-fashioned scales is unlikely to give you a meaningful result. As scientists, when we make measurements, we must always consider the appropriateness of the numbers we report. There are specific ways in which to do this that can give us confidence in interpreting data collected by ourselves and others. By the end of this section, you should be able to

- Determine the correct number of significant figures to report at the conclusion of an experiment

Two kinds of numbers are encountered in scientific work: *exact numbers* (those whose values are known exactly) and *inexact numbers* (those whose values have some uncertainty). Most of the exact numbers we will encounter in this book have defined values. For example, there are exactly 12 eggs in a dozen and exactly 1000 g in a kilogram. The number 1 in any conversion factor, such as 1 m = 100 cm or 1 kg = 1000 g, is an exact number. Exact numbers can also result from counting objects. For example, we can count the exact number of marbles in a jar or the exact number of people in a classroom.

Numbers obtained by measurement are always *inexact*. The equipment used to measure quantities always has inherent limitations (equipment errors), and there are differences in how different people make the same measurement (human errors). Suppose ten students with ten balances are to determine the mass of the same coin. The ten measurements will probably vary slightly for various reasons. The balances might be calibrated slightly differently, and there might be differences in how each student reads the mass from the balance. Remember: *Uncertainties always exist in measured quantities*.

Precision and Accuracy

The terms *precision* and *accuracy* are often used in discussing the uncertainties of measured values. **Precision** is a measure of how closely individual measurements agree with one another. **Accuracy** refers to how closely individual measurements agree with the correct, or “true,” value. The dart analogy in [Figure 1.23](#) illustrates the difference between these two concepts.

In the laboratory, we often perform several “trials” of an experiment and average the results. The precision of the measurements is often expressed in terms of the *standard deviation* (Appendix A.5), which reflects how much the individual measurements differ



High precision can be achieved on a scale like this one, which has 0.1 mg accuracy.


Go Figure

How would the darts be positioned on the target for the case of “good accuracy, poor precision”?



Good accuracy
Good precision



Poor accuracy
Good precision



Poor accuracy
Poor precision

▲ Figure 1.23 Precision and accuracy.

from the average. We gain confidence in our measurements if we obtain nearly the same value each time—that is, when the standard deviation is small. Figure 1.23 reminds us, however, that precise measurements can be inaccurate. For example, if a very sensitive balance is poorly calibrated, the masses we measure will be consistently either high or low. They will be inaccurate even if they are precise.

Significant Figures

Suppose you determine the mass of a coin on a balance capable of measuring to the nearest 0.0001 g. You could report the mass as 2.2405 ± 0.0001 g. The \pm notation (read “plus or minus”) expresses the magnitude of the uncertainty of your measurement. In much scientific work, we drop the \pm notation with the understanding that *there is always some uncertainty in the last digit reported for any measured quantity*.

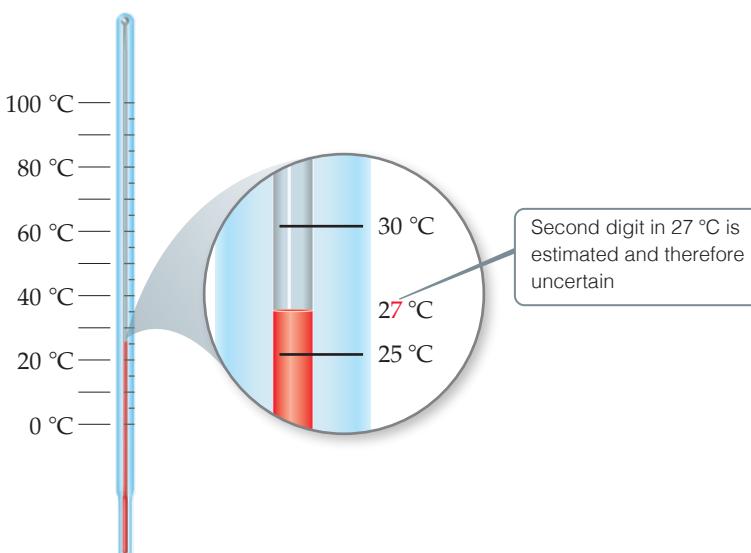
Figure 1.24 shows a thermometer with its liquid column between two scale marks. We can read the certain digits from the scale and estimate the uncertain one. Seeing that the liquid is between the 25°C and 30°C marks, we estimate the temperature to be 27°C , being uncertain of the second digit of our measurement. By *uncertain* we mean that the temperature is reliably 27°C and not 28°C or 26°C , but we can’t say that it is *exactly* 27°C .

All digits of a measured quantity, including the uncertain one, are called **significant figures**. A measured mass reported as 2.2 g has two significant figures, whereas one reported as 2.2405 g has five significant figures. The greater the number of significant figures, the greater the precision implied for the measurement.

To determine the number of significant figures in a reported measurement, read the number from left to right, counting the digits starting with the first digit that is not zero. *In any measurement that is properly reported, all nonzero digits are significant.* Because zeros can be used either as part of the measured value or merely to locate the decimal point, they may or may not be significant:

- Zeros *between* nonzero digits are always significant— 1005 kg (four significant figures); 7.03 cm (three significant figures).
- Zeros *at the beginning* of a number are never significant; they merely indicate the position of the decimal point— 0.02 g (one significant figure); 0.0026 cm (two significant figures).
- Zeros *at the end* of a number are significant if the number contains a decimal point— 0.0200 g (three significant figures); 3.0 cm (two significant figures).

A problem arises when a number ends with zeros but contains no decimal point. In such cases, it is normally assumed that the zeros are not significant. Exponential notation (Appendix A.1) can be used to indicate whether end zeros are significant.



▲ Figure 1.24 Uncertainty and significant figures in a measurement.

For example, a mass of 10 300 g can be written to show three, four, or five significant figures depending on how the measurement is obtained:

1.03×10^4 g	(three significant figures)
1.030×10^4 g	(four significant figures)
1.0300×10^4 g	(five significant figures)

In these numbers, all the zeros to the right of the decimal point are significant (rules 1 and 3). (The exponential term 10^4 does not add to the number of significant figures.)

Sample Exercise 1.6

Assigning Appropriate Significant Figures

The US state of Colorado is listed in a road atlas as having a population of 5,546,574 and an area of 269,837 square kilometers. Do the numbers of significant figures in these two quantities seem reasonable? If not, what seems to be wrong with them?

SOLUTION

The population of Colorado must vary from day to day as people move in or out, are born, or die. Thus, the reported number suggests a much higher degree of *accuracy* than is possible. Moreover, it would not be feasible to actually count every individual resident in the state at any given time. Thus, the reported number suggests far greater *precision* than is possible. A reported number of 5,500,000 would better reflect the actual state of knowledge.

The area of Colorado does not normally vary from time to time, so the question here is whether the accuracy of the measurements is

good to six significant figures. It would be possible to achieve such accuracy using satellite technology, provided the legal boundaries are known with sufficient accuracy.

► Practice Exercise

The back inside cover of the book tells us that there are 1000 cm^3 in 1 liter. Does this make the liter an exact volume?

Sample Exercise 1.7

Determining the Number of Significant Figures in a Measurement

How many significant figures are in each of the following numbers (assume that each number is a measured quantity):

- (a) 4.003, (b) 6.023×10^{23} , (c) 5000?

SOLUTION

(a) Four; the zeros are significant figures. (b) Four; the exponential term does not add to the number of significant figures. (c) One; we assume that the zeros are not significant when there is no decimal point shown. If the number has more significant figures, a decimal point should be employed or the number written in exponential notation. Thus, 5000. has four significant figures, whereas 5.00×10^3 has three.

► Practice Exercise

An object is determined to have a mass of 0.01080 g. How many significant figures are there in this measurement?

- (a) 2 (b) 3 (c) 4 (d) 5 (e) 6

Significant Figures in Calculations

Apply the following rule when carrying measured quantities through calculations.

The least certain measurement limits the certainty of the calculated quantity and thereby determines the number of significant figures in the final answer.

The final answer should be reported with only one uncertain digit. To keep track of significant figures in calculations, we will make frequent use of two rules: one for addition and subtraction, and another for multiplication and division.

1. For addition and subtraction, the result has the same number of decimal places as the measurement with the fewest decimal places. When the result contains more than the correct number of significant figures, it must be rounded off. Consider the following

example where the first three numbers are added to give the fourth and in which the uncertain digits appear in color:

This number limits the number of the significant figures in the result	20.4 ² 1.32 ² 83. ¹ 104. ⁸⁴²	← two decimal places ← three decimal places ← one decimal places ← round off to one decimal place (104.8)
--	---	--

We report the result as 104.8 because 83.1 has only one decimal place.

2. For multiplication and division, the result contains the same number of significant figures as the measurement with the *fewest significant figures*. When the result contains more than the correct number of significant figures, it must be rounded off. For example, the area of a rectangle whose measured edge lengths are 6.221 and 5.2 cm should be reported with two significant figures, 32 cm^2 , even though a calculator shows the product to have more digits:

$$\text{Area} = (6.221 \text{ cm})(5.2 \text{ cm}) = 32.3492 \text{ cm}^2 \Rightarrow \text{round off to } 32 \text{ cm}^2$$

because 5.2 has two significant figures.

In determining the final answer for a calculated quantity, *exact numbers* are assumed to have an infinite number of significant figures. Thus, when we say, “There are 1000 meters in 1 kilometer,” the number 1000 is exact, and we need not worry about the number of significant figures in it.

When *rounding off numbers* look at the leftmost digit to be removed:

- If the leftmost digit removed is less than 5, the preceding number is left unchanged. Thus, rounding off 7.248 to two significant figures gives 7.2.
- If the leftmost digit removed is 5 or greater, the preceding number is increased by 1. Rounding off 4.735 to three significant figures gives 4.74, and rounding 2.376 to two significant figures gives 2.4.*

Sample Exercise 1.8

Determining the Number of Significant Figures in a Calculated Quantity



The width, length, and height of a small box are 15.5, 27.3, and 5.4 cm, respectively. Calculate the volume of the box, using the correct number of significant figures in your answer.

SOLUTION

In reporting the volume, we can show only as many significant figures as given in the dimension with the fewest significant figures, which is that for the height (two significant figures):

$$\begin{aligned}\text{Volume} &= \text{width} \times \text{length} \times \text{height} \\ &= (15.5 \text{ cm})(27.3 \text{ cm})(5.4 \text{ cm}) \\ &= 2285.01 \text{ cm}^3 \Rightarrow 2.3 \times 10^3 \text{ cm}^3\end{aligned}$$

A calculator used for this calculation shows 2285.01, which we must round off to two significant figures. Because the resulting

number is 2300, it is best reported in exponential notation, 2.3×10^3 , to clearly indicate two significant figures.

► Practice Exercise

It takes 10.5 s for a sprinter to run 100.00 m. Calculate her average speed in meters per second and express the result to the correct number of significant figures.

*Your instructor may want you to use a slight variation on the rule when the leftmost digit to be removed is exactly 5, with no following digits or only zeros following. One common practice is to round up to the next higher number if that number will be even and otherwise leave the number unchanged. Thus, 4.7350 would be rounded to 4.74, and 4.7450 would also be rounded to 4.74.



Sample Exercise 1.9

Determining the Number of Significant Figures in a Calculated Quantity

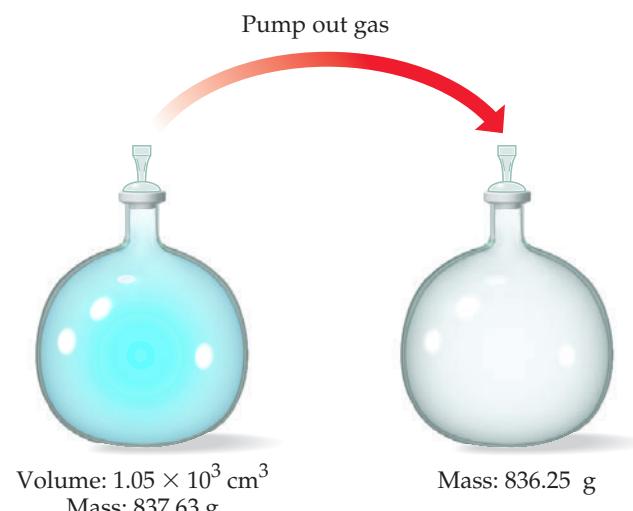
A vessel containing a gas at 25 °C is weighed, emptied, and then reweighed as depicted in **Figure 1.25**. From the data provided, calculate the density of the gas at 25 °C.

SOLUTION

To calculate the density, we must know both the mass and the volume of the gas. The mass of the gas is just the difference in the masses of the full and empty container:

$$\begin{array}{r} 837.63 \text{ g} \\ -836.25 \text{ g} \\ \hline 1.38 \text{ g} \end{array}$$

In subtracting numbers, we determine the number of significant figures in our result by counting decimal places in each quantity.



▲ Figure 1.25 Uncertainty and significant figures in a measurement.

In this case each quantity has two decimal places. Thus, the mass of the gas, 1.38 g, has two decimal places.

Using the volume given in the question, $1.05 \times 10^3 \text{ cm}^3$, and the definition of density, we have

$$\begin{aligned} \text{Density} &= \frac{\text{mass}}{\text{volume}} = \frac{1.38 \text{ g}}{1.05 \times 10^3 \text{ cm}^3} \\ &= 1.31 \times 10^{-3} \text{ g/cm}^3 = 0.00131 \text{ g/cm}^3 \end{aligned}$$

In dividing numbers, we determine the number of significant figures our result should contain by counting the number of significant figures in each quantity. There are three significant figures in our answer, corresponding to the number of significant figures in the two numbers that form the ratio. Notice that in this example, following the rules for determining significant figures gives an answer containing only three significant figures, even though the measured masses contain five significant figures.

► Practice Exercise

You are asked to determine the mass of a piece of copper using its reported density, 8.96 g/mL, and a 150-mL graduated cylinder. First, you add 105 mL of water to the graduated cylinder; then you place the piece of copper in the cylinder and record a volume of 137 mL. What is the mass of the copper reported with the correct number of significant figures? (a) 287 g (b) 3.5×10^{-3} g/mL (c) 286.72 g/mL (d) 3.48×10^{-3} g/mL (e) 2.9×10^2 g/mL

When a calculation involves two or more steps and you write answers for intermediate steps, retain at least one nonsignificant digit for the intermediate answers. This procedure ensures that small errors from rounding at each step do not combine to affect the final result. When using a calculator, you may enter the numbers one after another, rounding only the final answer. Accumulated rounding-off errors may account for small differences among results you obtain and answers given in the text for numerical problems.

Self-Assessment Exercise

- 1.20** Do the following numbers, all representing the thickness of a typical piece of paper, have the same number of significant figures: 0.10 mm, 1.0×10^{-4} m, 100 μm ?
- (a) Yes
(b) No

Exercises

- 1.21** Indicate which of the following are exact numbers: (a) the mass of a 7.5 by 12.5 cm index card, (b) the number of grams in a kilogram, (c) the volume of a cup of Seattle's Best coffee, (d) the number of centimeters in a kilometer, (e) the number of microseconds in a week, (f) the number of pages in this book.
- 1.22** What is the number of significant figures in each of the following measured quantities? (a) 902.5 kg, (b) 3×10^{-6} m, (c) 0.0096 L, (d) $2.94 \times 10^3 \text{ m}^2$, (e) 92.03 km (f) 782.234 g.
- 1.23** Round each of the following numbers to three significant figures and express the result in standard exponential

notation: (a) 2048732.23 (b) 0.000292945 (c) -82454.09
 (d) 942.057024 (e) -0.00000324683.

- 1.24** Carry out the following operations and express the answers with the appropriate number of significant numbers.

(a) $43.029 + 0.02348$ (b) $952.72 - 73.4201$
 (c) $(2.93 \times 10^3) (0.732)$ (d) $0.06324/0.624$

- 1.25** You weigh an object on a balance and read the mass in grams according to the picture. How many significant figures are in this measurement?



1.20 (b)

Answers to Self-Assessment Exercises



1.7 | Dimensional Analysis



Anyone who has travelled to another country and looked at the price of fuel there will experience the need to convert from one set of units to another. Most often, this conversion is between currencies but, in some cases, the standard volume may have different units—liters or gallons for example.

The SI system of units provides a common ‘language’ in which to communicate measurement. However, there are still some industries and some countries in which historic units are employed. By the end of this section, you should be able to

- Convert between different units for a particular measurement

Because measured quantities have units associated with them, it is important to keep track of units as well as numerical values when using the quantities in calculations. Throughout the text we use dimensional analysis in solving problems. In **dimensional analysis**, units are multiplied together or divided into each other along with the numerical values. Equivalent units cancel each other. Using dimensional analysis helps ensure that solutions to problems yield the proper units. Moreover, it provides a systematic way of solving many numerical problems and of checking solutions for possible errors.

Conversion Factors

The key to using dimensional analysis is the correct use of *conversion factors* to change one unit into another. A **conversion factor** is a fraction whose numerator and denominator are the same quantity expressed in different units. For example, 2.54 cm and 1 inch are the same length: 2.54 cm = 1 in. This relationship allows us to write two conversion factors:

$$\frac{2.54 \text{ cm}}{1 \text{ in.}} \quad \text{and} \quad \frac{1 \text{ in.}}{2.54 \text{ cm}}$$

We use the first factor to convert inches to centimeters. For example, the length in centimeters of an object that is 8.50 in. long is

$$\text{Number of centimeters} = (8.50 \text{ in.}) \frac{2.54 \text{ cm}}{1 \text{ in.}} = 21.6 \text{ cm}$$

Desired unit
Given unit

The unit inches in the denominator of the conversion factor cancels the unit inches in the given data (8.50 *inches*), so that the centimeters unit in the numerator of the conversion factor becomes the unit of the final answer. Because the numerator and denominator of a conversion factor are equal, multiplying any quantity by a conversion factor is equivalent to multiplying by the number 1 and so does not change the intrinsic value of the quantity. The length 8.50 in. is the same as the length 21.6 cm.

In general, we begin any conversion by examining the units of the given data and the units we desire. We then ask ourselves what conversion factors we have available to take us from the units of the given quantity to those of the desired one. When we multiply a quantity by a conversion factor, the units multiply and divide as follows:

$$\text{Given unit} \times \frac{\text{desired unit}}{\text{given unit}} = \text{desired unit}$$

If the desired units are not obtained in a calculation, an error must have been made somewhere. Careful inspection of units often reveals the source of the error.

Sample Exercise 1.10

Converting Units

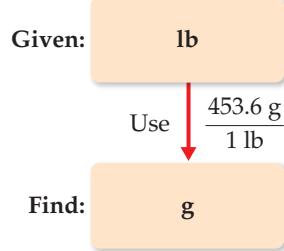
If a woman has a mass of 115 lb, what is her mass in grams? (The measurement of a ‘pound’ has the units ‘lb’. 1 lb = 453.6 g)

SOLUTION

Because we want to change from pounds to grams, we look for a relationship between these units of mass. The conversion factor table found on the back inside cover tells us that 1 lb = 453.6 g. To cancel pounds and leave grams, we write the conversion factor with grams in the numerator and pounds in the denominator:

$$\text{Mass in grams} = (115 \text{ lb}) \left(\frac{453.6 \text{ g}}{1 \text{ lb}} \right) = 5.22 \times 10^4 \text{ g}$$

The answer can be given to only three significant figures, the number of significant figures in 115 lb. The process we have used is diagrammed on the top right column.



► Practice Exercise

At a particular instant in time, the Earth is judged to be 92,955,000 miles from the Sun. What is the distance in kilometers to four significant figures? (1 mile = 1.6093 km.)

- (a) $5763 \times 10^4 \text{ km}$ (b) $1.496 \times 10^8 \text{ km}$ (c) $1.49596 \times 10^8 \text{ km}$
 (d) $1.483 \times 10^4 \text{ km}$ (e) 57,759,000 km

STRATEGIES FOR SUCCESS Estimating Answers

Calculators are wonderful devices; they enable you to get to the wrong answer very quickly. Of course, that's not the destination you want. You can take certain steps to avoid putting that wrong answer into your homework set or on an exam. One is to keep track of the units in a calculation and use the correct conversion factors. Second, you can do a quick mental check to be sure that your answer is reasonable: you can try to make a "ballpark" estimate.

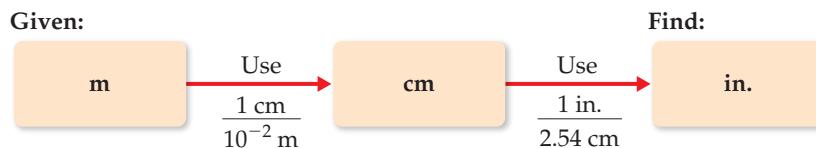
A ballpark estimate involves making a rough calculation using numbers that are rounded off in such a way that the arithmetic can be done without a calculator. Even though this approach does not

give an exact answer, it gives one that is roughly the correct size. By using dimensional analysis and by estimating answers, you can readily check the reasonableness of your calculations.

You can get better at making estimates by practicing in everyday life. How far is it from your house to the chemistry lecture hall? How many bikes are there on campus? If you respond "I have no idea" to these questions, you're giving up too easily. Try estimating familiar quantities and you'll get better at making estimates in science and in other aspects of your life where a misjudgment can be costly.

Using Two or More Conversion Factors

It is often necessary to use several conversion factors in solving a problem. As an example, let's convert the length of an 8.00 m rod to inches. We may be given the relationship between centimeters and inches (1 in. = 2.54 cm). We combine this with our knowledge of SI prefixes, ie $1 \text{ cm} = 10^{-2} \text{ m}$. Thus, we can convert step by step, first from meters to centimeters and then from centimeters to inches:



Combining the given quantity (8.00 m) and the two conversion factors, we have

$$\text{Number of inches} = (8.00 \text{ m}) \left(\frac{1 \text{ cm}}{10^{-2} \text{ m}} \right) \left(\frac{1 \text{ in.}}{2.54 \text{ cm}} \right) = 315 \text{ in.}$$

The first conversion factor is used to cancel meters and convert the length to centimeters. Thus, meters are written in the denominator and centimeters in the numerator. The second conversion factor is used to cancel centimeters and convert the length to inches, so it has centimeters in the denominator and inches, the desired unit, in the numerator.

Note that you could have used $100 \text{ cm} = 1 \text{ m}$ as a conversion factor as well in the second parentheses. As long as you keep track of your given units and cancel them properly to obtain the desired units, you are likely to be successful in your calculations.



Sample Exercise 1.11

Converting Units Using Two or More Conversion Factors



The average speed of a nitrogen molecule in air at 25 °C is 515 m/s. Convert this speed to miles per hour.

SOLUTION

To go from the given units, m/s, to the desired units, mi/hr, we must convert meters to miles and seconds to hours. From our knowledge of SI prefixes we know that $1 \text{ km} = 10^3 \text{ m}$ and we are given $1 \text{ mi} = 1.6093 \text{ km}$. Thus, we can convert m to km and then convert km to mi. From our knowledge of time we know that $60 \text{ s} = 1 \text{ min}$ and $60 \text{ min} = 1 \text{ hr}$. Thus, we can convert s to min and then convert min to hr. The overall process is

Applying first the conversions for distance and then those for time, we can set up one long equation in which unwanted units are canceled:

$$\begin{aligned} \text{Speed in mi/hr} &= \left(515 \frac{\text{m}}{\text{s}} \right) \left(\frac{1 \text{ km}}{10^3 \text{ m}} \right) \left(\frac{1 \text{ mi}}{1.6093 \text{ km}} \right) \left(\frac{60 \text{ s}}{1 \text{ min}} \right) \left(\frac{60 \text{ min}}{1 \text{ hr}} \right) \\ &= 1.15 \times 10^3 \text{ mi/hr} \end{aligned}$$

Given:



Our answer has the desired units. We can check our calculation, using the estimating procedure described in the “Strategies in Chemistry” box. The given speed is about 500 m/s. Dividing by 1000 converts m to km, giving 0.5 km/s. Because 1 mi is about 1.6 km, this speed corresponds to $0.5 / 1.6 = 0.3$ mi/s. Multiplying by 60 gives about $0.3 \times 60 = 20$ mi/min. Multiplying again by 60 gives $20 \times 60 = 1200$ mi/hr. The approximate solution (about 1200 mi/hr) and the detailed solution (1150 mi/hr) are reasonably close.

The answer to the detailed solution has three significant figures, corresponding to the number of significant figures in the given speed in m/s.

► Practice Exercise

A car travels 28 mi per gallon of gasoline. What is the mileage in kilometers per liter?

Conversions Involving Volume

The conversion factors previously noted convert from one unit of a given measure to another unit of the same measure, such as from length to length. We also have conversion factors that convert from one measure to a different one. The density of a substance, for example, can be treated as a conversion factor between mass and volume. Suppose we want to know the mass in grams of 2 cubic inches (2.00 in.^3) of gold, which has a density of 19.3 g/cm^3 . The density gives us the conversion factors:

$$\frac{19.3 \text{ g}}{1 \text{ cm}^3} \quad \text{and} \quad \frac{1 \text{ cm}^3}{19.3 \text{ g}}$$

Because we want a mass in grams, we use the first factor, which has mass in grams in the numerator. To use this factor, however, we must first convert cubic inches to cubic centimeters. The relationship between in.^3 and cm^3 may not be given directly, but the relationship between inches and centimeters is given: 1 in. = 2.54 cm (exactly). Cubing both sides of this equation gives $(1 \text{ in.})^3 = (2.54 \text{ cm})^3$, from which we write the desired conversion factor:

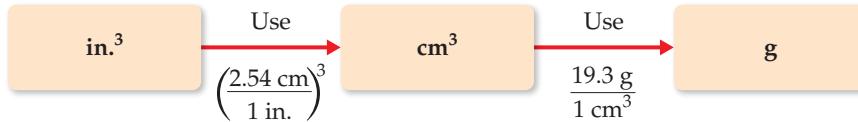
$$\frac{(2.54 \text{ cm})^3}{(1 \text{ in.})^3} = \frac{(2.54)^3 \text{ cm}^3}{(1)^3 \text{ in.}^3} = \frac{16.39 \text{ cm}^3}{1 \text{ in.}^3}$$

Notice that both the numbers and the units are cubed. Also, because 2.54 is an exact number, we can retain as many digits of $(2.54)^3$ as we need. We have used four, one more than the number of digits in the density (19.3 g/cm^3). Applying our conversion factors, we can now solve the problem:

$$\text{Mass in grams} = (2.00 \text{ in.}^3) \left(\frac{16.39 \text{ cm}^3}{1 \text{ in.}^3} \right) \left(\frac{19.3 \text{ g}}{1 \text{ cm}^3} \right) = 633 \text{ g}$$

The procedure is diagrammed here. The final answer is reported to three significant figures, the same number of significant figures as in 2.00 in.³ and 19.3 g.

Given:





Sample Exercise 1.12

Converting Volume Units

Earth's oceans contain approximately $1.36 \times 10^9 \text{ km}^3$ of water. Calculate the volume in liters.

SOLUTION

We know $1 \text{ L} = 10^{-3} \text{ m}^3$ (Figure 1.19), but there is no relationship listed involving km^3 . From our knowledge of SI prefixes, however, we know $1 \text{ km} = 10^3 \text{ m}$ and we can use this relationship between lengths to write the desired conversion factor between volumes:

$$\left(\frac{10^3 \text{ m}}{1 \text{ km}}\right)^3 = \frac{10^9 \text{ m}^3}{1 \text{ km}^3}$$

Thus, converting from km^3 to m^3 to L , we have

$$\begin{aligned}\text{Volume in liters} &= (1.36 \times 10^9 \text{ km}^3) \left(\frac{10^9 \text{ m}^3}{1 \text{ km}^3}\right) \left(\frac{1 \text{ L}}{10^{-3} \text{ m}^3}\right) \\ &= 1.36 \times 10^{21} \text{ L}\end{aligned}$$



How many liters of water do Earth's oceans contain?

► Practice Exercise

A barrel of oil as measured on the oil market is equal to 1.333 U.S. barrels. A U.S. barrel is equal to 31.5 gal. If oil is on the market at \$94.0 per barrel, what is the price in dollars per gallon?

- (a) \$2.24/gal (b) \$3.98/gal (c) \$2.98/gal (d) \$1.05/gal
(e) \$8.42/gal

STRATEGIES FOR SUCCESS The Importance of Practice

If you have ever played a musical instrument or participated in athletics, you know that the keys to success are practice and discipline. You cannot learn to play a piano merely by listening to music, and you cannot learn how to play basketball merely by watching games on television. Likewise, you cannot learn chemistry by merely watching your instructor give lectures. Simply reading this book, listening to lectures, or reviewing notes will not usually be sufficient when exam time comes around. Your task is to master chemical concepts to a degree that you can put them to use in solving problems and answering questions. Solving problems correctly takes practice—actually, a fair amount of it. You will do well in your chemistry course if you embrace the idea that you need to master the materials presented and then learn how to apply them in solving problems. Even if you're a brilliant student, this will take time; it's what being a student is all about. Almost no one fully absorbs new material on a first reading, especially when new concepts are being presented. You are sure to master the content of the chapters more fully by reading

them through at least twice, even more for passages that present you with difficulties in understanding.

Throughout the book, we have provided sample exercises in which the solutions are shown in detail. For practice exercises, we supply only the answer, at the back of the book. It is important that you use these exercises to test yourself.

The practice exercises in this text and the homework assignments given by your instructor provide the minimal practice that you will need to succeed in your chemistry course. Only by working all the assigned problems will you face the full range of difficulty and coverage that your instructor expects you to master for exams. There is no substitute for a determined and perhaps lengthy effort to work problems on your own. If you are stuck on a problem, however, ask for help from your instructor, a teaching assistant, a tutor, or a fellow student. Spending an inordinate amount of time on a single exercise is rarely effective unless you know that it is particularly challenging and is expected to require extensive thought and effort.

STRATEGIES FOR SUCCESS The Features of This Book

If, like most students, you haven't yet read the part of the Preface to this text entitled TO THE STUDENT, *you should do it now*. In less than two pages of reading you will encounter valuable advice on how to navigate your way through this book and through the course. We're serious! This is advice you can use.

The TO THE STUDENT section describes how text features such as "What's Ahead," Key Terms, Learning Outcomes, and Key Equations will help you remember what you have learned. If you have registered for MasteringChemistry®, you will have access to many helpful animations, tutorials, and additional problems correlated to specific topics and sections of each chapter. An eBook is also available online. In addition, the Pearson eText brings to life the content of each chapter with animations and videos. Interactive end-of-section Self-Assessment Exercises are featured with specific wrong-answer feedback.

As previously mentioned, working exercises is very important—in fact, essential. You will find a large variety of exercises at the end of each section and chapter that are designed to test your problem-solving skills in chemistry. Your instructor will very likely assign some of these exercises as homework.

- The first few exercises called "Visualizing Concepts" are meant to test how well you understand a concept without plugging a lot of numbers into a formula.
- Additional Exercises appear after the regular exercises; the chapter sections that they cover are not identified.
- Integrative Exercises, which start appearing in Chapter 3, are problems that require skills learned in previous chapters.

- Also first appearing in Chapter 3 are Design an Experiment exercises consisting of problem scenarios that challenge you to design experiments to test hypotheses.

Many chemical databases are available, usually on the Web.

- The *CRC Handbook of Chemistry and Physics* is the standard reference for many types of data and is available in libraries.

- The *Merck Index* is a standard reference for the properties of many organic compounds, especially those of biological interest.
- WebElements (<http://www.webelements.com/>) is a good website for looking up the properties of the elements.
- Wolfram Alpha (<http://www.wolframalpha.com/>) can also be a source of useful information on substances, numerical values, and other data.

Self-Assessment Exercise

- 1.26** A soft drink indicates that it contains 21 kJ of energy. Given $1\text{ J} = 0.2390$ calories, what is the energy content in calories (hint: mind the significant figures)?

- (a) 5.0 kcal
(b) 88 cal

- (c) 5.0 kcal
(d) 5019 cal

Exercises

- 1.27** Using your knowledge of metric units and the information given in the chapter, write down the conversion factors needed to convert (a) in. to cm (b) lb to g (c) μg to g (d) ft^2 to cm^2 .

- 1.28** (a) A bumblebee flies with a ground speed of 15.2 m/s. Calculate its speed in km/hr. (b) The lung capacity of the blue whale is 5.0×10^3 L. Convert this volume into gallons. (c) The Statue of Liberty is 151 ft tall. Calculate its height in meters. (d) Bamboo can grow up to 60.0 cm/day. Convert this growth rate into inches per hour.

- 1.29** Perform the following conversions: (a) 5.00 days to s, (b) 0.0550 miles to m, (c) \$1.89/gal to dollars per liter, (d) 0.510 in/ms to km/h, (e) 22.50 gal/min to L/s, (f) 0.02500 ft^3 to cm^3 .

- 1.30** (a) How many liters of wine can be held in a wine barrel whose capacity is 31 gal? (b) The recommended adult dose of Elixophyllin®, a drug used to treat asthma, is 6 mg/kg of

body mass. Calculate the dose in milligrams for a 185 lb person. (c) If an automobile is able to travel 400 km on 47.3 L of fuel, what is the gas mileage in miles per gallon? (d) When the coffee is brewed according to directions, a pound of coffee beans yields 50 cups of coffee. How many kg of coffee are required to produce 200 cups of coffee?

- 1.31** The density of air at ordinary atmospheric pressure and 25 °C is 1.19 g/L. What is the mass, in kilograms, of the air in a room that measures 4.5 m × 5.0 m × 2.5 m?

- 1.32** Gold can be hammered into extremely thin sheets called gold leaf. An architect wants to cover a 30 m × 25 m ceiling with gold leaf that is twelve-millionths of a centimeter thick. The density of gold is 19.32 g/cm³, and gold costs \$1654 per troy ounce (1 troy ounce = 31.1034768 g). How much will it cost the architect to buy the necessary gold?

1.26 (c)

Answers to Self-Assessment Exercises

Chapter Summary and Key Terms

THE STUDY OF CHEMISTRY (SECTION 1.1) Chemistry is the study of the composition, structure, properties, and changes of matter. The composition of matter relates to the kinds of elements it contains. The structure of matter relates to the ways the atoms of these elements are arranged. A property is any characteristic that gives a sample of matter its unique identity. A molecule is an entity composed of two or more atoms, with the atoms attached to one another in a specific way.

The scientific method is a dynamic process used to answer questions about the physical world. Observations and experiments lead to tentative explanations or hypotheses. As a hypothesis is tested and refined, a theory may be developed that can predict the results of future observations and experiments. When observations repeatedly lead to the same consistent results, we speak of a scientific law, a general rule that summarizes how nature behaves.

CLASSIFICATIONS OF MATTER (SECTION 1.2) Matter exists in three physical states, gas, liquid, and solid, which are known as the states of matter. There are two kinds of pure substances: elements and compounds. Each element has a single kind of atom and is represented by a chemical symbol consisting of one or two letters, with the first letter capitalized. Compounds are composed of two or more elements joined

chemically. The law of constant composition, also called the law of definite proportions, states that the elemental composition of a pure compound is always the same. Most matter consists of a mixture of substances. Mixtures have variable compositions and can be either homogeneous or heterogeneous; homogeneous mixtures are called solutions.

PROPERTIES OF MATTER (SECTION 1.3) Each substance has a unique set of physical properties and chemical properties that can be used to identify it. During a physical change, matter does not change its composition. Changes of state are physical changes. In a chemical change (chemical reaction), a substance is transformed into a chemically different substance. Intensive properties are independent of the amount of matter examined and are used to identify substances. Extensive properties relate to the amount of substance present. Differences in physical and chemical properties are used to separate substances.

THE NATURE OF ENERGY (SECTION 1.4) Energy is defined as the capacity to do work or transfer heat. Work is the energy transferred when a force exerted on an object causes a displacement of that object, and heat is the energy used to cause the temperature of an object to increase. An object can possess energy in two forms: kinetic energy,

which is the energy associated with the motion of an object, and **potential energy**, which is the energy that an object possesses by virtue of its position relative to other objects. Important forms of potential energy include gravitational energy and electrostatic energy.

UNITS OF MEASUREMENT (SECTION 1.5) Measurements in chemistry are made using the **metric system**. Special emphasis is placed on **SI units**, which are based on the meter, the kilogram, and the second as the basic units of length, **mass**, and time, respectively. SI units use prefixes to indicate fractions or multiples of base units. The **SI temperature scale** is the **Kelvin scale**, although the **Celsius scale** is frequently used as well. **Absolute zero** is the lowest temperature attainable. It has the value 0 K. A **derived unit** is obtained by multiplication or division of SI base units. Derived units are needed for defined quantities such as speed or volume. **Density** is an important derived unit that equals mass divided by volume.

UNCERTAINTY IN MEASUREMENT (SECTION 1.6) All measured quantities are inexact to some extent. The **precision** of a measurement

indicates how closely different measurements of a quantity agree with one another. The **accuracy** of a measurement indicates how well a measurement agrees with the accepted or “true” value. The **significant figures** in a measured quantity include one estimated digit, the last digit of the measurement. The significant figures indicate the extent of the uncertainty of the measurement. Certain rules must be followed so that a calculation involving measured quantities is reported with the appropriate number of significant figures.

DIMENSIONAL ANALYSIS (SECTION 1.7) In the **dimensional analysis** approach to problem solving, we keep track of units as we carry measurements through calculations. The units are multiplied together, divided into each other, or canceled like algebraic quantities. Obtaining the proper units for the final result is an important means of checking the method of calculation. When converting units and when carrying out several other types of problems, **conversion factors** can be used. These factors are ratios constructed from valid relations between equivalent quantities.

Learning Outcomes

After studying this chapter, you should be able to:

- Distinguish among elements, compounds, and mixtures. (Section 1.2) *Related Exercises: 1.3, 1.5, 1.45*
- Identify symbols of common elements. (Section 1.2) *Related Exercises: 1.4, 1.46*
- Distinguish between chemical and physical changes. (Section 1.3) *Related Exercises: 1.8, 1.9, 1.47, 1.48*
- Distinguish between kinetic and potential energy. (Section 1.4) *Related Exercises: 1.12, 1.13*

- Calculate the kinetic energy of an object. (Section 1.4). *Related Exercises: 1.50, 1.51, 1.52, 1.53*
- Identify common metric prefixes. (Section 1.5) *Related Exercises: 1.15, 1.54*
- Demonstrate the use of significant figures, scientific notation, and SI units in calculations. (Section 1.6) *Related Exercises: 1.22, 1.24, 1.63, 1.65*
- Use appropriate SI units for defined quantities, and employ dimensional analysis in calculations. (Sections 1.5 and 1.7) *Related Exercises: 1.28, 1.30, 1.68, 1.70*

Key Equations

- $w = F \times d$ [1.1]
- $E_k = \frac{1}{2}mv^2$ [1.2]
- $K = {}^\circ C + 273.15$ [1.3]
- Density = $\frac{\text{mass}}{\text{volume}}$ [1.4]

Work done by a force in the direction of displacement

Kinetic energy

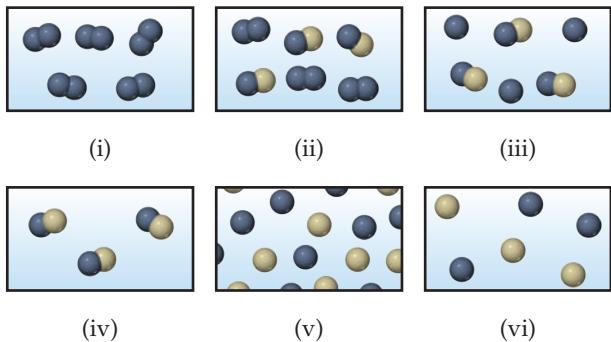
Converting between Celsius (${}^\circ C$) and Kelvin (K) temperature scales

Definition of density

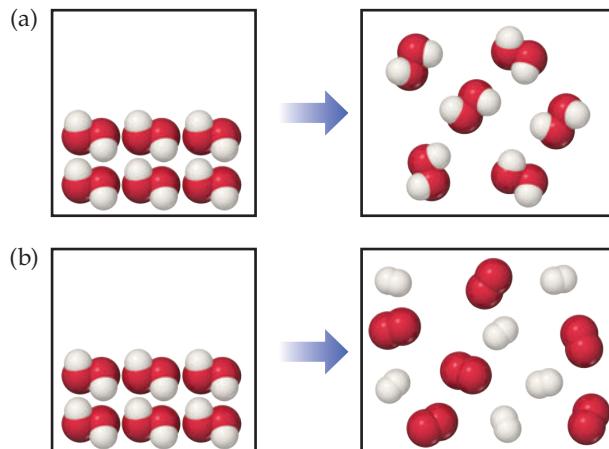
Exercises

Visualizing Concepts

- 1.33** Which of the following figures represents (a) a pure element, (b) a mixture of two elements, (c) a pure compound, (d) a mixture of an element and a compound? (More than one picture might fit each description.) [Section 1.2]



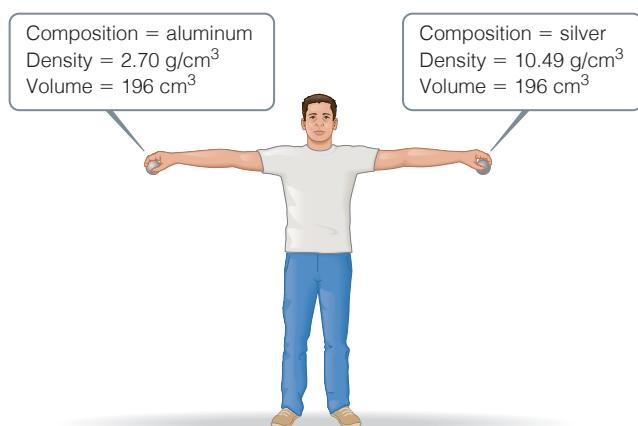
- 1.34** Which of the following diagrams represents a chemical change? [Section 1.3]



- 1.35** Musical instruments like trumpets and trombones are made from an alloy called brass. Brass is composed of copper and zinc atoms and appears homogeneous under an optical microscope. The approximate composition of most brass objects is a 2:1 ratio of copper to zinc atoms, but the exact ratio varies somewhat from one piece of brass to another. **(a)** Would you classify brass as an element, a compound, a homogeneous mixture, or a heterogeneous mixture? **(b)** Would it be correct to say that brass is a solution? [Section 1.2]



- 1.36** Consider the two spheres shown here, one made of silver and the other of aluminum. **(a)** What is the mass of each sphere in kg? **(b)** The force of gravity acting on an object is $F = mg$, where m is the mass of an object and g is the acceleration of gravity (9.8 m/s^2). How much work do you do on each sphere if you raise it from the floor to a height of 2.2 m? **(c)** Does the act of lifting the sphere off the ground increase the potential energy of the aluminum sphere by a larger, smaller, or same amount as the silver sphere? **(d)** If you release the spheres simultaneously, they will have the same velocity when they hit the ground. Will they have the same kinetic energy? If not, which sphere will have more kinetic energy? [Section 1.4]



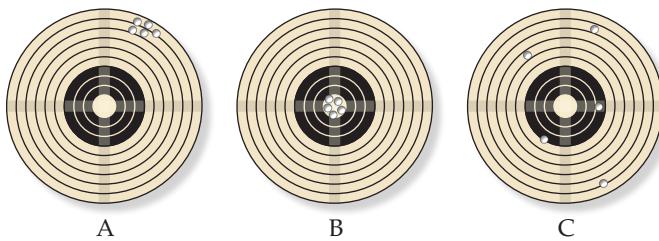
- 1.37** Is the separation method used in brewing a cup of coffee best described as distillation, filtration, or chromatography? [Section 1.3]



- 1.38** Identify each of the following as measurements of length, area, volume, mass, density, time, or temperature: **(a)** 25 ps, **(b)** 374.2 mg, **(c)** 77 K, **(d)** $100,000 \text{ km}^2$, **(e)** $1.06 \mu\text{m}$, **(f)** 16 nm^2 , **(g)** -78°C , **(h)** 2.56 g/cm^3 , **(i)** 28 cm^3 . [Section 1.5]

- 1.39** **(a)** Three spheres of equal size are composed of aluminum (density = 2.70 g/cm^3), silver (density = 10.49 g/cm^3), and nickel (density = 8.90 g/cm^3). List the spheres from lightest to heaviest. **(b)** Three cubes of equal mass are composed of gold (density = 19.32 g/cm^3), platinum (density = 21.45 g/cm^3), and lead (density = 11.35 g/cm^3). List the cubes from smallest to largest. [Section 1.5]

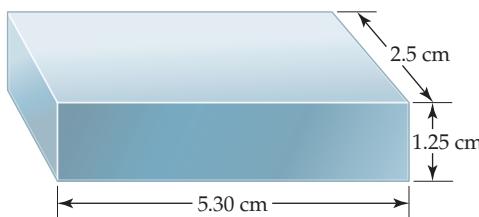
- 1.40** The three targets from a rifle range shown here were produced by: **(A)** the instructor firing a newly acquired target rifle; **(B)** the instructor firing his personal target rifle; and **(C)** a student who has fired his target rifle only a few times. **(a)** Comment on the accuracy and precision for each of these three sets of results. **(b)** For the A and C results in the future to look like those in B, what needs to happen? [Section 1.6]



- 1.41** **(a)** What is the length of the pencil in the following figure if the ruler reads in centimeters? How many significant figures are there in this measurement? **(b)** An automobile speedometer with circular scales reading both miles per hour and kilometers per hour is shown. What speed is indicated, in both units? How many significant figures are in the measurements? [Section 1.6]



- 1.42** **(a)** How many significant figures should be reported for the volume of the metal bar shown here? **(b)** If the mass of the bar is 104.72 g, how many significant figures should be reported when its density is determined using the calculated volume? [Section 1.6]



- 1.43** Consider the jar of jelly beans in the photo. To get an estimate of the number of beans in the jar you weigh six beans and obtain masses of 3.15, 3.12, 2.98, 3.14, 3.02, and 3.09 g. Then you weigh the jar with all the beans in it, and obtain a mass of 2082 g. The empty jar has a mass of 653 g. Based on these data, estimate the number of beans in the jar. Justify the number of significant figures you use in your estimate. [Section 1.6]



- 1.44** This photo shows a picture of an agate stone. Jack, who picked up the stone on the Lake Superior shoreline and polished it, insists that agate is a chemical compound. Ellen argues that it cannot be a compound. Discuss the relative merits of their positions. [Section 1.2]



Classification and Properties of Matter (Sections 1.2 and 1.3)

- 1.45** Classify each of the following as a pure substance or a mixture. If a mixture, indicate whether it is homogeneous or heterogeneous: **(a)** milk, **(b)** beer, **(c)** diamond, **(d)** mayonnaise.
- 1.46** Give the chemical symbol or name for each of the following elements, as appropriate: **(a)** rhenium, **(b)** tungsten, **(c)** caesium, **(d)** hydrogen, **(e)** indium, **(f)** As, **(g)** Xe, **(h)** Kr, **(i)** Te, **(j)** Ge.
- 1.47** **(a)** Read the following description of the element zinc and indicate which are physical properties and which are chemical properties.



Zinc melts at 420 °C. When zinc granules are added to dilute sulfuric acid, hydrogen is given off and the metal dissolves. Zinc has a hardness on the Mohs scale of 2.5 and a density of 7.13 g/cm³ at 25 °C. It reacts slowly with oxygen gas at elevated temperatures to form zinc oxide, ZnO.

- (a)** Which properties of zinc can you describe from the photo? Are these physical or chemical properties?

- 1.48** A match is lit to light a candle. The following observations are made: **(a)** The candle burns. **(b)** Some wax melts. **(c)** Melted wax solidifies on the candleholder. **(d)** Soot (carbon) is produced by the burning of the match and the candle. Which of these occurrences are due to physical changes, and which are due to chemical changes?

- 1.49** A silvery metal is put inside a beaker of water. Bubbles form on the surface of the metal and it dissolves gradually. **(a)** Is this an example of a chemical or a physical change? **(b)** Do you expect the remaining solution to be a pure substance or a mixture?

The Nature of Energy (Section 1.4)

- 1.50** **(a)** Calculate the kinetic energy, in joules, of a 15-g bullet moving at 120 m/s. **(b)** When the bullet is stopped by a bulletproof vest, which form of energy does the kinetic energy of the bullet convert to?

- 1.51** **(a)** A baseball weighs 145.4 g. What is the kinetic energy, in joules, of this baseball when it is thrown by a major league pitcher at 150 km/h? **(b)** By what factor will the kinetic energy change if the speed of the baseball is decreased to 90 km/h? **(c)** What happens to the kinetic energy when the baseball is caught by the catcher? Is it converted mostly to heat or to some form of potential energy?

- 1.52** What is the kinetic energy and velocity of the aluminum sphere in Problem 1.36 at the moment it hits the ground? (Assume that energy is conserved during the fall and that 100% of the sphere's initial potential energy is converted to kinetic energy by the time impact occurs.)

- 1.53** What is the kinetic energy and velocity of the silver sphere in Problem 1.36 at the moment it hits the ground? (Assume that energy is conserved during the fall and that 100% of the sphere's initial potential energy is converted to kinetic energy by the time impact occurs.)

Units of Measurement (Section 1.5)

- 1.54** Use appropriate metric prefixes to write the following measurements without use of exponents: **(a)** 7.29 × 10⁶ g **(b)** 6.1 × 10⁻¹⁰ m **(c)** 1.828 × 10⁻³ s **(d)** 3.523 × 10⁹ m³ **(e)** 9.62 × 10² m/s **(f)** 8.923 × 10⁻¹² kg **(g)** 3.552 × 10¹² L.

- 1.55** Make the following conversions: **(a)** 83 °F to °C **(b)** 29 °C to °F **(c)** 294 °C to K **(d)** 832 K to °C **(e)** 721 K to °F **(f)** 35 °F to K.

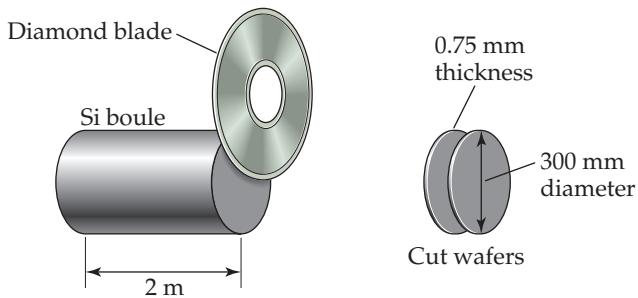
- 1.56** **(a)** A child has a fever of 101 °F. What is the temperature in °C? **(b)** In a desert, the temperature can be as high as 45 °C, what is the temperature in °F? **(c)** During winter, the temperature of the Arctic region can drop below –50 °C, what is the temperature in degree Fahrenheit and in Kelvin? **(d)** The sublimation temperature of dry ice is –78.5 °C. Convert this temperature to degree Fahrenheit and Kelvin. **(e)** Ethanol boils at 351 K. Convert this temperature to degree Fahrenheit and degree Celsius.

- 1.57** **(a)** What is the mass of a silver cube whose edges measure 2.00 cm each at 25 °C? The density of silver is 10.49 g/cm³ at 25 °C. **(b)** The density of aluminum is 2.70 g/cm³ at 25 °C. What is the weight of the aluminum foil with an area of 0.5 m² and a thickness of 0.5 mm? **(c)** The density of hexane is 0.655 g/mL at 25 °C. Calculate the mass of 1.5 L of hexane at this temperature.

- 1.58** **(a)** After the label fell off a bottle containing a clear liquid believed to be benzene, a chemist measured the density of the liquid to verify its identity. A 25.0-mL portion of the liquid

had a mass of 21.95 g. A chemistry handbook lists the density of benzene at 15 °C as 0.8787 g/mL. Is the calculated density in agreement with the tabulated value? (b) An experiment requires 15.0 g of cyclohexane, whose density at 25 °C is 0.7781 g/mL. What volume of cyclohexane should be used? (c) A spherical ball of lead has a diameter of 5.0 cm. What is the mass of the sphere if lead has a density of 11.34 g/cm³? (The volume of a sphere is $(4/3)\pi r^3$, where r is the radius.)

- 1.59** Silicon for computer chips is grown in large cylinders called “boules” that are 300 mm in diameter and 2 m in length, as shown. The density of silicon is 2.33 g/cm³. Silicon wafers for making integrated circuits are sliced from a 2.0 m boule and are typically 0.75 mm thick and 300 mm in diameter. (a) How many wafers can be cut from a single boule? (b) What is the mass of a silicon wafer? (The volume of a cylinder is given by $\pi r^2 h$, where r is the radius and h is its height.)



- 1.60** Use of the British thermal unit (Btu) is common in some types of engineering work. A Btu is the amount of heat required to raise the temperature of 1 lb of water by 1°F. Calculate the number of joules in a Btu.
- 1.61** A watt is a measure of power (the rate of energy change) equal to 1 J/s. (a) Calculate the number of joules in a kilowatt-hour. (b) An adult person radiates heat to the surroundings at about the same rate as a 100-watt electric incandescent light bulb. What is the total amount of energy in kcal radiated to the surroundings by an adult over a 24 h period?

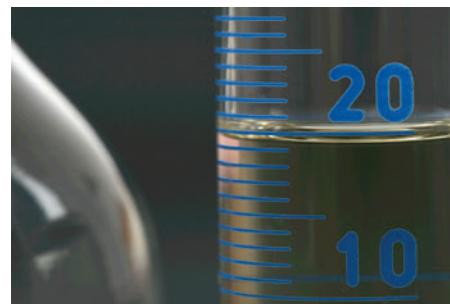
Uncertainty in Measurement (Section 1.6)

- 1.62** Indicate which of the following are exact numbers: (a) the mass of a 945-mL can of coffee, (b) the number of students in your chemistry class, (c) the temperature of the surface of the Sun, (d) the mass of a postage stamp, (e) the number of milliliters in a cubic meter of water, (f) the average height of NBA basketball players.
- 1.63** Indicate the number of significant figures in each of the following measured quantities: (a) 62.65 km/hr, (b) 78.00 K, (c) 36.9 mL, (d) 250 mm, (e) 89.2 metric tons, (f) 6.4224×10^2 m³.
- 1.64** (a) The diameter of Earth at the equator is 12756.27 km. Round this number to three significant figures and express it in standard exponential notation. (b) The circumference of Earth through the poles is 40,008 km. Round this number to four significant figures and express it in standard exponential notation.
- 1.65** Carry out the following operations and express the answers with the appropriate number of significant numbers.
- (a) $(6.234 + 8.72) \times 0.6746$
- (b) $732.1 - (892.5/8.2)$

(c) $[(3.696 \times 10^5) - (6.234 \times 10^3)] \times 0.0742$

(d) $0.006438 \times 108 - (8.639 + 8.52)$

- 1.66** You have a graduated cylinder that contains a liquid (see photograph). Write the volume of the liquid, in milliliters, using the proper number of significant figures.



Dimensional Analysis (Section 1.7)

Note: 12 inches = 1 foot

1 m = 39.37 inches

1 mile = 1.609 km

1 pound = 454 g

1 gallon = 3.7854 L

1 m³ = 264 gallon

- 1.67** Using your knowledge of metric units and the given information, write down the conversion factors needed to convert (a) km/h to m/s (b) mL to μ L (c) ps to s (d) m³ to gal.

- 1.68** (a) The speed of light in a vacuum is 2.998×10^8 m/s. Calculate its speed in miles per hour. (b) The Sears Tower in Chicago is 1454 ft tall. Calculate its height in meters. (c) The Vehicle Assembly Building at the Kennedy Space Center in Florida has a volume of 3,666,500 m³. Convert this volume to liters and express the result in standard exponential notation. (d) An individual suffering from a high cholesterol level in her blood has 242 mg of cholesterol per 100 mL of blood. If the total blood volume of the individual is 5.2 L, how many grams of total blood cholesterol does the individual's body contain?

- 1.69** Carry out the following conversions: (a) 0.105 in. to mm, (b) 8.75 mm/s to km/h, (c) \$3.99/lb to dollars per kg, (d) 8.75 lb/ft³ to g/mL.

- 1.70** (a) In March 1989, the *Exxon Valdez* ran aground and spilled 240,000 barrels of crude petroleum off the coast of Alaska. One barrel of petroleum is equal to 42 gal. How many liters of petroleum were spilled?

- 1.71** The indoor concentration of ozone above 300 μ g/m³ is considered to be unhealthy. What mass of ozone in grams is present in a room measuring 3.2 m \times 2.8 m \times 4.1 m?

- 1.72** A copper refinery produces a copper ingot weighing 70 kg. If the copper is drawn into wire whose diameter is 7.50 mm, how many meters of copper can be obtained from the ingot? The density of copper is 8.94 g/cm³. (Assume that the wire is a cylinder whose volume $V = \pi r^2 h$, where r is its radius and h is its height or length.)

Additional Exercises

- 1.73** Classify each of the following as a pure substance, a solution, or a heterogeneous mixture: (a) a leaf, (b) a 999 gold bar, (c) stainless steel.
- 1.74** (a) Which is more likely to eventually be shown to be incorrect: an hypothesis or a theory? (b) A(n) _____ reliably predicts the behavior of matter, while a(n) _____ provides an explanation for that behavior.
- 1.75** A sample of ascorbic acid (vitamin C) is synthesized in the laboratory. It contains 1.50 g of carbon and 2.00 g of oxygen. Another sample of ascorbic acid isolated from citrus fruits contains 6.35 g of carbon. According to the law of constant composition, how many grams of oxygen does it contain?
- 1.76** Ethyl chloride is sold as a liquid (see photo) under pressure for use as a local skin anesthetic. Ethyl chloride boils at 12 °C at atmospheric pressure. When the liquid is sprayed onto the skin, it boils off, cooling and numbing the skin as it vaporizes. (a) What changes of state are involved in this use of ethyl chloride? (b) What is the boiling point of ethyl chloride in degrees Fahrenheit? (c) The bottle shown contains 103.5 mL of ethyl chloride. The density of ethyl chloride at 25 °C is 0.765 g/cm³. What is the mass of ethyl chloride in the bottle?



- 1.77** Two students determine the percentage of lead in a sample as a laboratory exercise. The true percentage is 22.52%. The students' results for three determinations are as follows:
- 22.52, 22.48, 22.54
 - 22.64, 22.58, 22.62
- (a) Calculate the average percentage for each set of data and state which set is the more accurate based on the average. (b) Precision can be judged by examining the average of the deviations from the average value for that data set. (Calculate the average value for each data set; then calculate the average value of the absolute deviations of each measurement from the average.) Which set is more precise?
- 1.78** Is the use of significant figures in each of the following statements appropriate? (a) The 2005 circulation of *National Geographic* was 7,812,564. (b) On July 1, 2005, the population of Cook County, Illinois, was 5,303,683. (c) In the United

States, 0.621% of the population has the surname Brown. (d) You calculate your grade point average to be 3.87562.

- 1.79** What type of quantity (for example, length, volume, density) do the following units indicate? (a) m³, (b) ns, (c) mm, (d) g/dm³, (e) °C, (f) ms⁻¹, (g) Pa.
- 1.80** Give the derived SI units for each of the following quantities in base SI units:
- acceleration = distance/time²
 - force = mass × acceleration
 - work = force × distance
 - pressure = force/area
 - power = work/time
 - velocity = distance/time
 - energy = mass × (velocity)²
- 1.81** The distance from Earth to the Moon is approximately 240,000 mi. (a) What is this distance in meters? (b) The peregrine falcon has been measured as traveling up to 350 km/hr in a dive. If this falcon could fly to the Moon at this speed, how many seconds would it take? (c) The speed of light is 3.00×10^8 m/s. How long does it take for light to travel from Earth to the Moon and back again? (d) Earth travels around the Sun at an average speed of 29.783 km/s. Convert this speed to miles per hour.
- 1.82** Which of the following would you characterize as pure or nearly pure substance? (a) stomach acid; (b) dry ice; (c) ice-cream; (d) stainless steel; (e) petroleum; (f) distilled water; (g) carbon monoxide gas; (h) compressed air in balloon.
- 1.83** The U.S. quarter has a mass of 5.67 g and is approximately 1.55 mm thick. (a) How many quarters would have to be stacked to reach 575 ft, the height of the Washington Monument? (b) How much would this stack weigh? (c) How much money would this stack contain? (d) The U.S. National Debt Clock showed the outstanding public debt to be \$16,213,166,914,811 on October 28, 2012. How many stacks like the one described would be necessary to pay off this debt?
- 1.84** In the United States, water used for irrigation is measured in acre-feet. An acre-foot of water covers an acre to a depth of exactly 1 ft. An acre is 4840 yd². An acre-foot is enough water to supply two typical households for 1.00 yr. (a) If desalinated water costs \$1950 per acre-foot, how much does desalinated water cost per liter? (b) How much would it cost one household per day if it were the only source of water?
- 1.85** By using estimation technique, determine which of the following is the heaviest and which is the lightest: a 10-lb bag of fertilizer, a 10-kg bag of rice, or 2 gal of olive oil (density = 0.918 g/cm³).
- 1.86** Suppose you decide to define your own temperature scale with units of O, using the freezing point (13 °C) and boiling point (360 °C) of oleic acid, the main component of olive oil. If you set the freezing point of oleic acid as 0 °O and the boiling point as 100 °O, what is the freezing point of water on this new scale?
- 1.87** Hexane (density = 0.659 g/mL) and acetic acid (density = 1.0446 g/mL) do not form a solution when mixed but are separate in distinct layers. A piece of oak wood (density = 900 kg/m³) is placed inside a test tube containing hexane and acetic acid solution; sketch how the three substances would position themselves.

- 1.88** Two spheres of equal volume are placed on the scales as shown. Which one is more dense?



- 1.89** Water has a density of 0.997 g/cm^3 at 25°C ; ice has a density of 0.917 g/cm^3 at -10°C . **(a)** If a soft-drink bottle whose volume is 1.50 L is completely filled with water and then frozen to -10°C , what volume does the ice occupy? **(b)** Can the ice be contained within the bottle?

- 1.90** A 32.65-g sample of a solid is placed in a flask. Toluene, in which the solid is insoluble, is added to the flask so that the total volume of solid and liquid together is 50.00 mL . The solid and toluene together weigh 58.58 g . The density of toluene at the temperature of the experiment is 0.864 g/mL . What is the density of the solid?

- 1.91** A thief plans to steal a cylindrical platinum medal with a radius of 2.3 cm and a thickness of 0.8 cm from a jewellery store. If the platinum has a density of 21.45 g/cm^3 , what is the mass of the medal in kg? [The volume of a cylinder is $V = \pi r^2 h$.]

- 1.92** Saline solution used in hospital contains 0.9% sodium chloride by mass. Calculate the number of grams of sodium chloride in 0.5 gal of saline solution if the solution has a density of 1.01 g/mL .

- 1.93** A 40-lb container of peat moss measures $14 \times 20 \times 30 \text{ in.}$ A 40-lb container of topsoil has a volume of 1.9 gal . **(a)** Calculate the average densities of peat moss and topsoil in units of g/cm^3 . Would it be correct to say that peat moss is “lighter” than topsoil? **(b)** How many bags of peat moss are needed to cover an area measuring $15.0 \text{ ft} \times 20.0 \text{ ft}$ to a depth of 3.0 in. ?

- 1.94** A 10.0 g block of gold is hammered into a thin gold sheet which has an area of 150 cm^2 . Given the density of gold is 19.3 g/cm^3 , what is the approximate thickness of the gold sheet in millimeters?

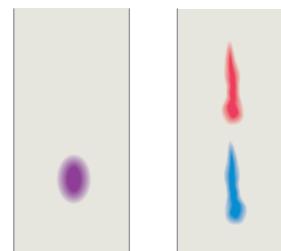
- 1.95** The total rate at which power is used by humans worldwide is approximately 15 TW (terawatts). The solar flux averaged over the sunlit half of Earth is 680 W/m^2 (assuming no clouds). The area of Earth’s disc as seen from the Sun is $1.28 \times 10^{14} \text{ m}^2$. The surface area of Earth is approximately $197,000,000$ square miles. How much of Earth’s surface would we need to cover with solar energy collectors to power the planet for use by all humans? Assume that the solar energy collectors can convert only 10% of the available sunlight into useful power.

- 1.96** In 2005, J. Robin Warren and Barry J. Marshall shared the Nobel Prize in Medicine for discovering the bacterium *Helicobacter pylori* and for establishing experimental proof that it plays a major role in gastritis and peptic ulcer disease. The story began when Warren, a pathologist, noticed that bacilli were associated with the tissues taken from patients suffering from ulcers. Look up the history of this case and describe Warren’s first hypothesis. What sorts of evidence did it take to create a credible theory based on it?

- 1.97** A 30.0-cm -long cylindrical plastic tube, sealed at one end, is filled with acetic acid. The mass of acetic acid needed to fill the tube is found to be 89.24 g . The density of acetic acid is 1.05 g/mL . Calculate the inner diameter of the tube in centimeters.

- 1.98** Gold is alloyed (mixed) with other metals to increase its hardness in making jewelry. **(a)** Consider a piece of gold jewelry that weighs 9.85 g and has a volume of 0.675 cm^3 . The jewelry contains only gold and silver, which have densities of 19.3 and 10.5 g/cm^3 , respectively. If the total volume of the jewelry is the sum of the volumes of the gold and silver that it contains, calculate the percentage of gold (by mass) in the jewelry. **(b)** The relative amount of gold in an alloy is commonly expressed in units of carats. Pure gold is 24 carat, and the percentage of gold in an alloy is given as a percentage of this value. For example, an alloy that is 50% gold is 12 carat. State the purity of the gold jewelry in carats.

- 1.99** Paper chromatography is a simple but reliable method for separating a mixture into its constituent substances. You have a mixture of two vegetable dyes, one red and one blue, that you are trying to separate. You try two different chromatography procedures and achieve the separations shown in the figure. Which procedure worked better? Can you suggest a method to quantify how good or poor the separation was?



- 1.100** Judge the following statements as true or false. If you believe a statement to be false, provide a corrected version.

- (a)** Air and water are both elements.
- (b)** All mixtures contain at least one element and one compound.
- (c)** Compounds can be decomposed into two or more other substances; elements cannot.
- (d)** Elements can exist in any of the three states of matter.
- (e)** When yellow stains in a kitchen sink are treated with bleach water, the disappearance of the stains is due to a physical change.
- (f)** A hypothesis is more weakly supported by experimental evidence than a theory.
- (g)** The number 0.0033 has more significant figures than 0.033 .
- (h)** Conversion factors used in converting units always have a numerical value of one.
- (i)** Compounds always contain at least two different elements.

- 1.101** You are assigned the task of separating a desired granular material with a density of 3.62 g/cm^3 from an undesired granular material that has a density of 2.04 g/cm^3 . You want to do this by shaking the mixture in a liquid in which the heavier material will fall to the bottom and the lighter material will float. A solid will float on any liquid that is more dense. Using an Internet-based source or a handbook of chemistry, find the densities of the following substances: carbon tetrachloride, hexane, and diiodomethane. Which of these

liquids will serve your purpose, assuming no chemical interaction takes place between the liquid and the solids?

- 1.102** In 2009, a team from Northwestern University and Western Washington University reported the preparation of a new “spongy” material composed of nickel, molybdenum, and sulfur that excels at removing mercury from water. The density of this new material is 0.20 g/cm^3 , and its surface area is 1242 m^2 per gram of material. (a) Calculate the volume

of a 10.0-mg sample of this material. (b) Calculate the surface area for a 10.0-mg sample of this material. (c) A 10.0-mL sample of contaminated water had 7.748 mg of mercury in it. After treatment with 10.0 mg of the new spongy material, 0.001 mg of mercury remained in the contaminated water. What percentage of the mercury was removed from the water? (d) What is the final mass of the spongy material after the exposure to mercury?