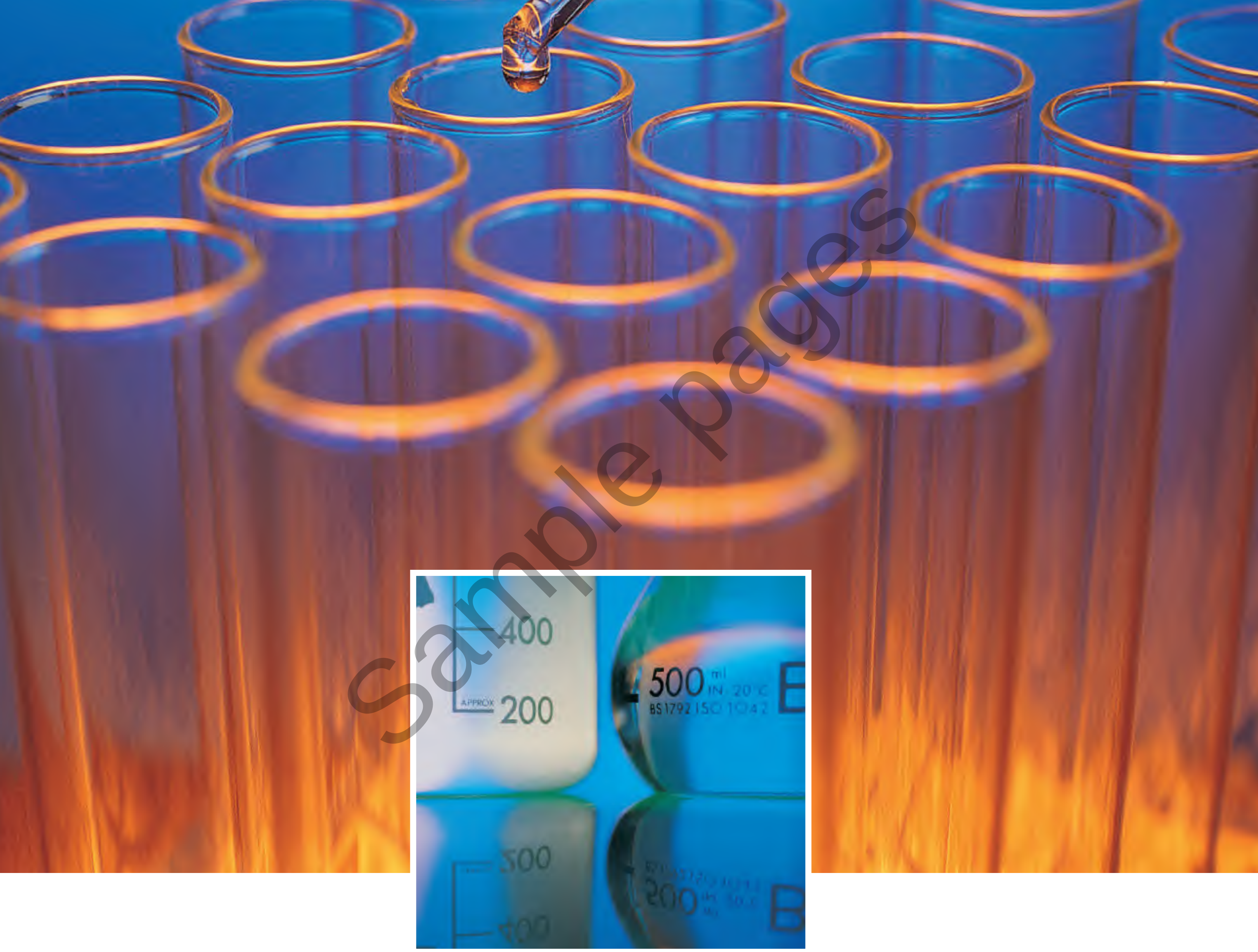
CHEMISTRY
THE CENTRAL SCIENCE



A Custom Edition

University of Canberra

Pearson

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Chapters 1 to 4 come from *Chemistry: The Central Science* 14th Edition, Global Edition in SI Units by Brown, LeMay, Bursten, Murphy, Woodward & Stoltzfus, ISBN 9781292221229.

SECTION TWO

Chapters 5 to 30 come from *Chemistry: The Central Science* 3rd Edition by Brown, LeMay, Bursten, Murphy, Woodward, Langford, Sagatys & George, ISBN 9781442554603.

INTRODUCTION: MATTER, ENERGY, AND MEASUREMENT

WHAT'S AHEAD

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 brilliant fall colors in leaves, the ways our bodies process the food we eat, and the electrical energy that powers our cell phones.

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 like van Gogh and Monet.

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.

◀ **THE MANUFACTURE OF SYNTHETIC PIGMENTS** is one of the oldest examples of industrial chemistry. The impressionist artists made extensive use of the bold colors of the newly available pigments, as exemplified in van Gogh's painting Glass with Roses.

1.1 | The Study of Chemistry

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.

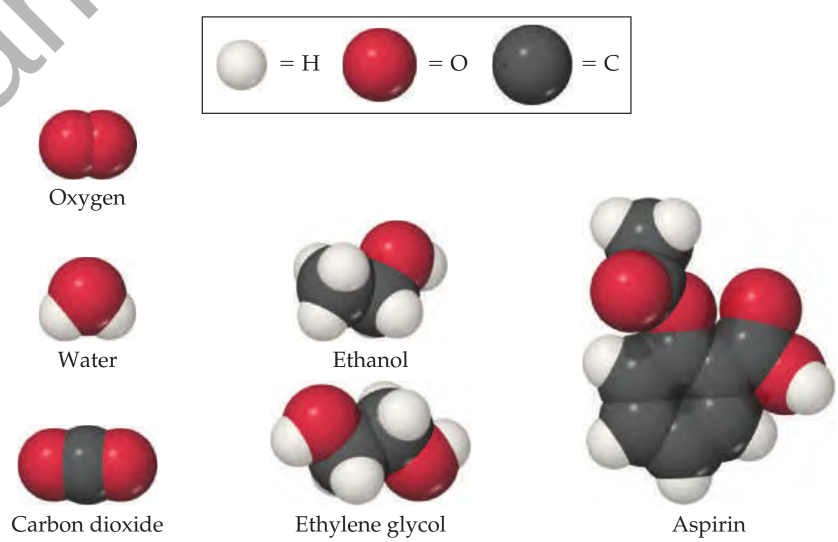
The Atomic and Molecular Perspective of Chemistry

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

Go Figure Which molecule has the most carbon atoms? How many carbon atoms does it contain?



▲ **Figure 1.1 Molecular models.** The white, black, and red spheres represent atoms of hydrogen, carbon, and oxygen, respectively.

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 two realms: the *macroscopic* realm of ordinary-sized objects (*macro* = large) and the *submicroscopic* realm of atoms and molecules. We make our observations in the macroscopic world, but to understand that world, we must visualize how atoms and molecules behave at the submicroscopic 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.

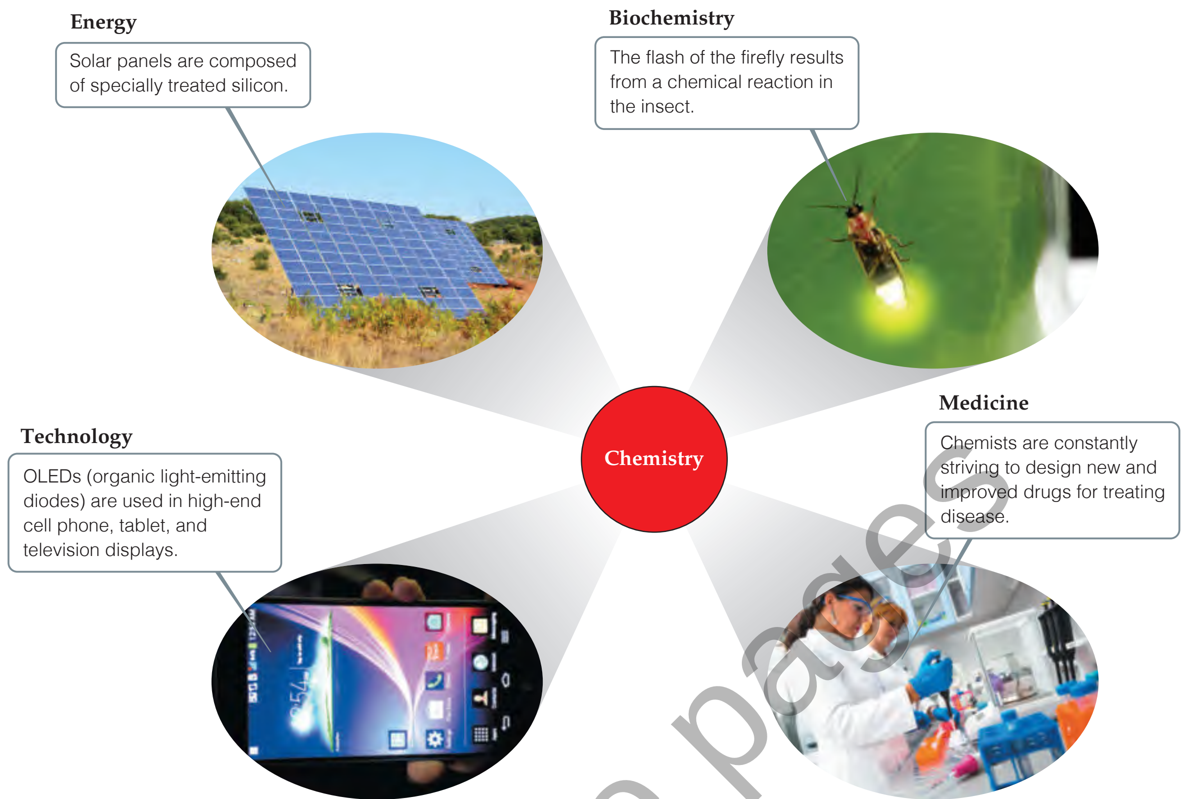
Give It Some Thought

(**a**) Approximately how many elements are there?
(**b**) What submicroscopic particles are the building blocks of matter?

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. We are all familiar with household chemicals, particularly those used for cleaning as shown in Figure 1.3. 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.**

1.2 | Classifications of Matter

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*).

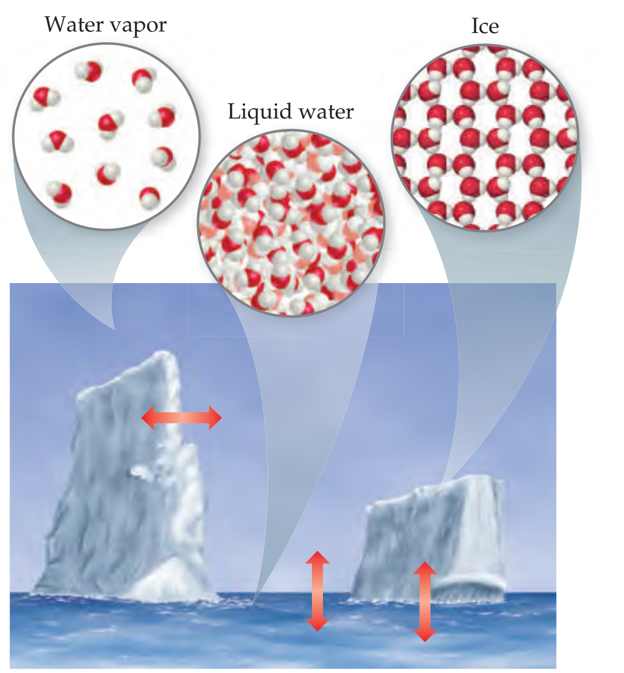
States of Matter

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.

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.

Pure Substances

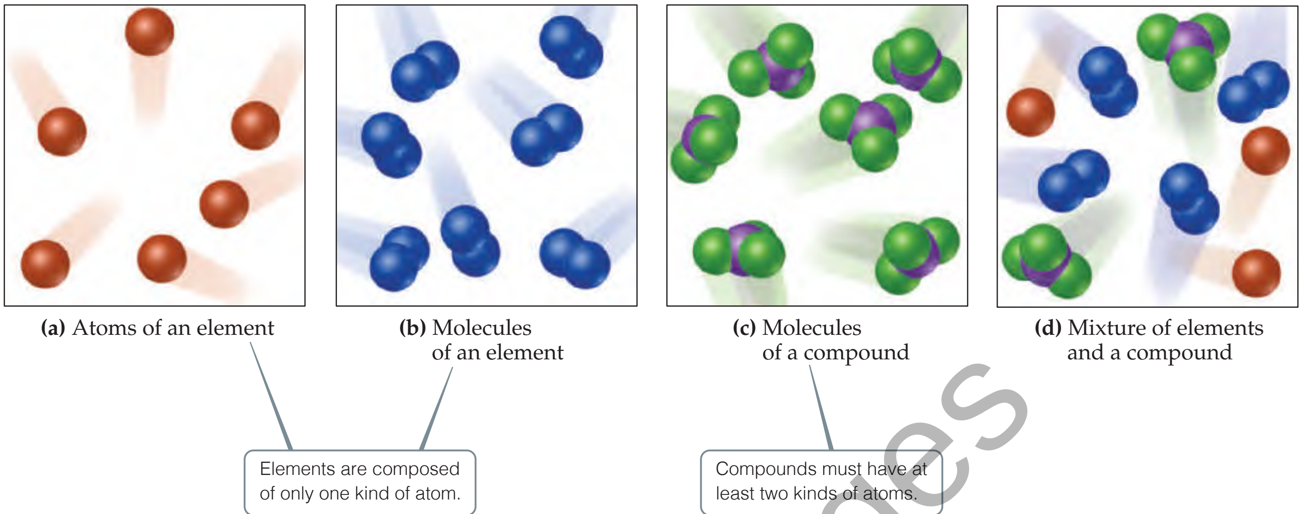
Most forms of matter we encounter—the air we breathe (a gas), the gasoline we burn in our cars (a liquid), and the sidewalk we walk on (a solid)—are not chemically pure. We can, however, separate these forms of matter into pure substances. A **pure substance** (usually referred to simply as a *substance*) 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.

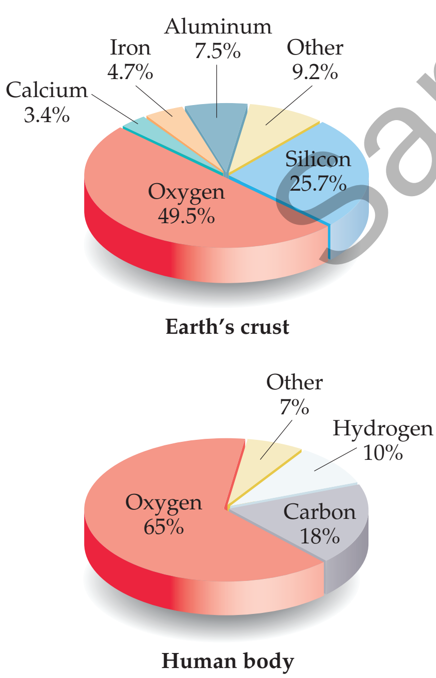
Figure 1.5(**d**) shows a mixture of substances. **Mixtures** are combinations of two or more substances in which each substance retains its chemical identity.

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



▲ **Figure 1.5 Molecular comparison 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

TABLE 1.1 Some Common Elements and Their Symbols

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

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

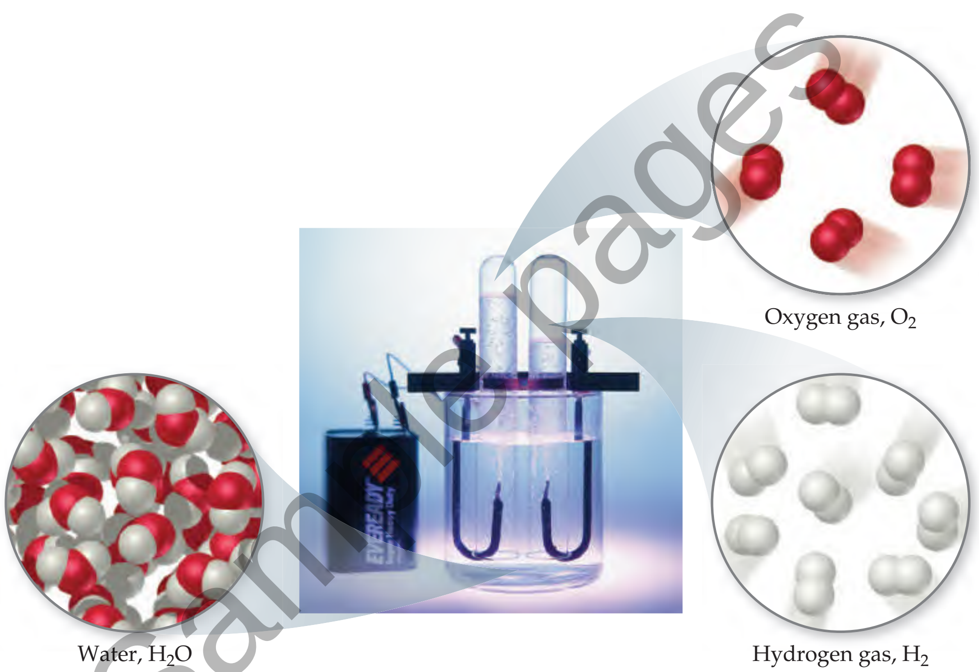
Section 2.5 and consider the periodically repeating properties of the elements in Chapter 7.

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**).

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?



▲ **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.

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. This macroscopic composition corresponds to the molecular composition, which consists of two hydrogen atoms combined with one oxygen atom:

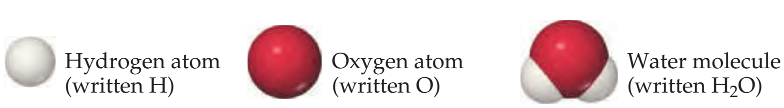


TABLE 1.2 Comparison of Water, Hydrogen, and Oxygen

	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 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.

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.

Give It Some Thought

Hydrogen, oxygen, and water are all composed of molecules. What is it about a molecule of water that makes it a compound, whereas hydrogen and oxygen are elements?

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 homogeneous mixtures [Figure 1.8(**b**)]. Homogeneous mixtures are also called **solutions**. Although the term *solution* conjures an image of a liquid, solutions can be solids, liquids, or gases.

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

(a) (b)

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.

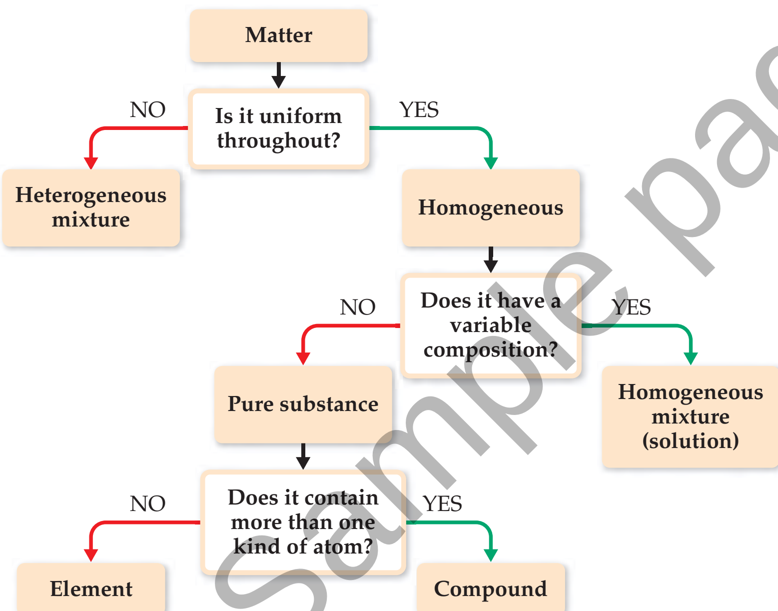


Figure 1.9 Classification of matter. All pure matter is classified ultimately as either an element or a compound.

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 1

Which of the following is the correct description of the inside of a grapefruit?

(**a**) It is a pure compound.
(**b**) It consists of a homogeneous mixture of compounds.
(**c**) It consists of a heterogeneous mixture of compounds.
(**d**) It consists of a heterogeneous mixture of elements and compounds.
(**e**) It consists of a single compound in different states.

Practice Exercise 2

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.

1.3 Properties of Matter

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.

Give It Some Thought

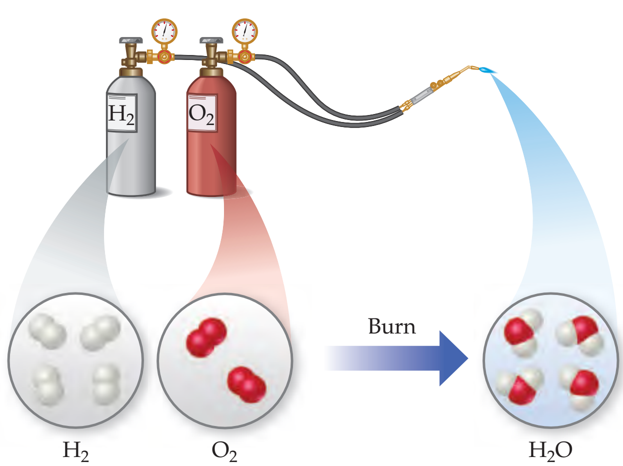
When we say that 10 kg of gold has a larger mass than 1 kg of copper, are we talking about an extensive or intensive property?

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.

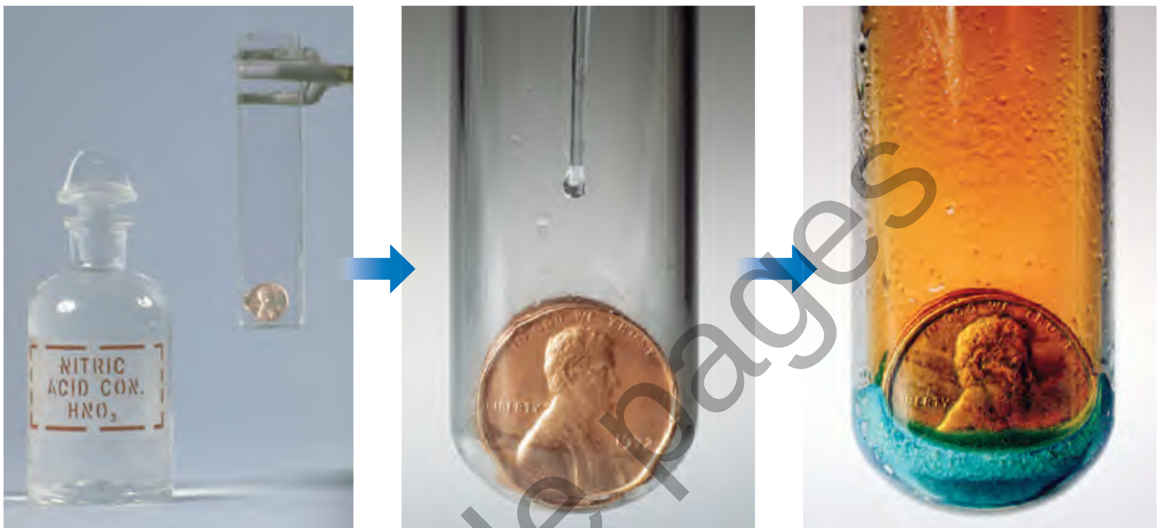
In a **chemical change** (also called a **chemical reaction**), a substance is transformed into a chemically different substance. When hydrogen burns in air, for example, it undergoes a chemical change because it combines with oxygen to form water (**Figure 1.10**).

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.



▲ **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.

Give It Some Thought

Which of these changes are physical and which are chemical? Explain.

(**a**) Iron rusts when it comes into contact with water and oxygen.
(**b**) Dry ice (solid carbon dioxide) goes directly from the solid to gas phase in the open atmosphere.
(**c**) A vitamin C tablet fizzes furiously when dropped into a glass of water.

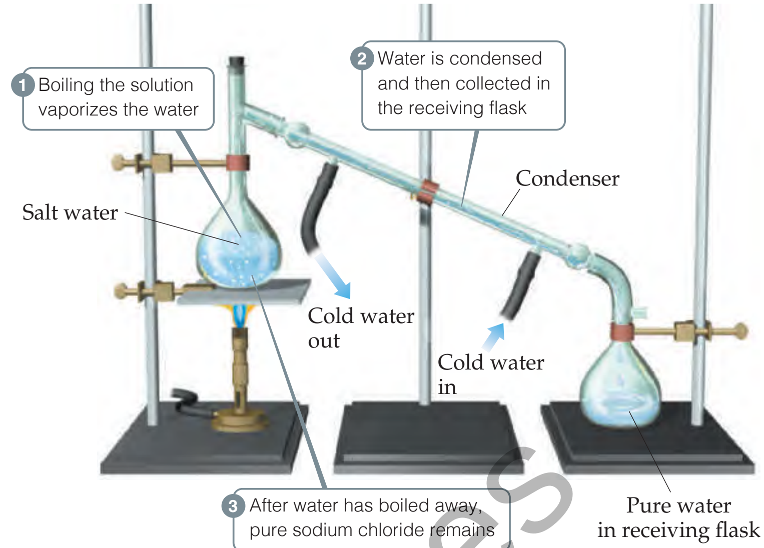
Separation of Mixtures

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 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 *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**.

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

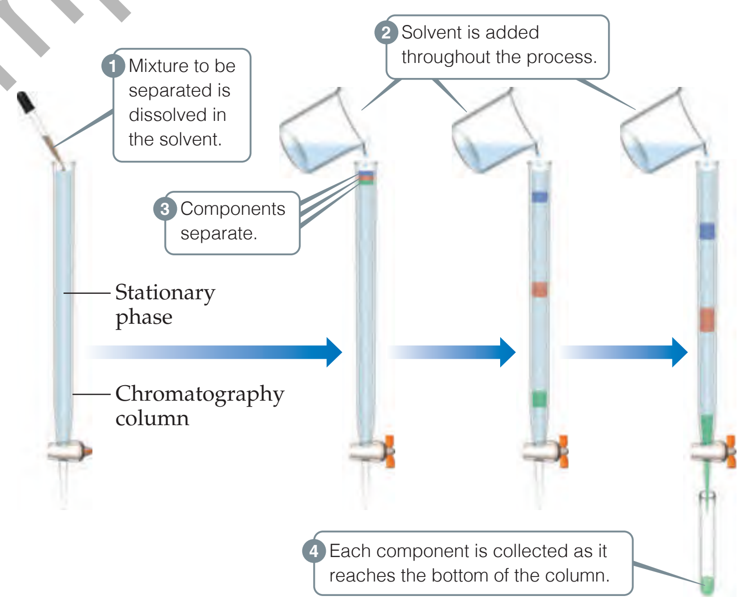
▲ **Figure 1.12 Separation by filtration.** A mixture of a solid and a liquid is poured through filter paper. The liquid passes through the paper while the solid remains on the paper.



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

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**.

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.**

1.4 The Nature of 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:

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

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:

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

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.

Give It Some Thought

Which change will lead to a larger change in the kinetic energy of an object, doubling its mass or doubling its speed?

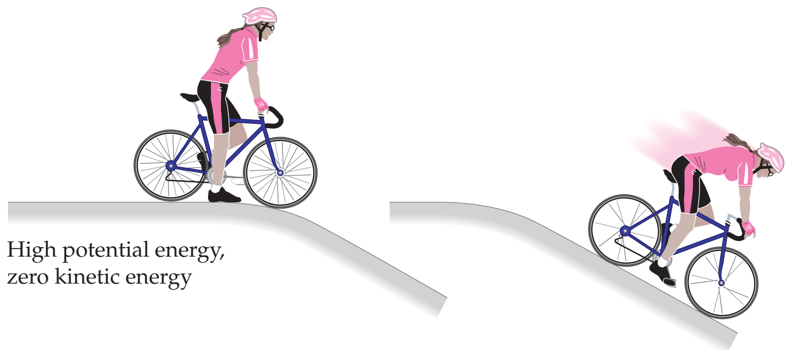
(a)

(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.

**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.



▲ **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.

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 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.

Give It Some Thought

What happens to the chemical energy of a nickel-metal hydride battery as it is recharged after powering an electric razor? Does it increase, decrease, or remain unchanged?

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. 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. The United States has traditionally used the English system, although use of the metric system has become more common (Figure 1.17).

▲ **Figure 1.17 Metric units.** Metric measurements are increasingly common in the United States, as exemplified by the volume printed on this soda can in both English units (fluid ounces, fl oz) and metric units (milliliters, mL).

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 chapter we will consider the base units for length, mass, and temperature.

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

Give It Some Thought

The package of a fluorescent bulb for a table lamp lists the light output in terms of lumens, lm. Which of the seven SI units would you expect to be part of the definition of a lumen?

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.

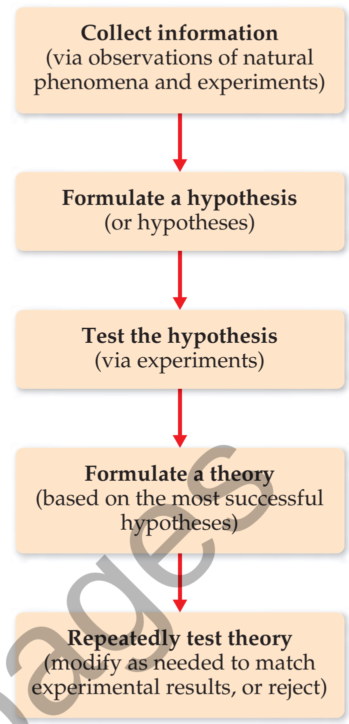
Continued

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 1.18, 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: 1.66, 1.88



▲ Figure 1.18 **The scientific method.**

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 1.4 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	
Peta	P	10^{15}	1 petawatt (PW)	$= 1 \times 10^{15}$ watts[a]
Tera	T	10^{12}	1 terawatt (TW)	$= 1 \times 10^{12}$ watts
Giga	G	10^{9}	1 gigawatt (GW)	$= 1 \times 10^{9}$ watts
Mega	M	10^{6}	1 megawatt (MW)	$= 1 \times 10^{6}$ watts
Kilo	k	10^{3}	1 kilowatt (kW)	$= 1 \times 10^{3}$ watts
Deci	d	10^{-1}	1 deciwatt (dW)	$= 1 \times 10^{-1}$ watt
Centi	c	10^{-2}	1 centiwatt (cW)	$= 1 \times 10^{-2}$ watt
Milli	m	10^{-3}	1 milliwatt (mW)	$= 1 \times 10^{-3}$ watt
Micro	μ[b]	10^{-6}	1 microwatt (μW)	$= 1 \times 10^{-6}$ watt
Nano	n	10^{-9}	1 nanowatt (nW)	$= 1 \times 10^{-9}$ watt
Pico	p	10^{-12}	1 picowatt (pW)	$= 1 \times 10^{-12}$ watt
Femto	f	10^{-15}	1 femtowatt (fW)	$= 1 \times 10^{-15}$ watt
Atto	a	10^{-18}	1 attowatt (aW)	$= 1 \times 10^{-18}$ watt
Zepto	z	10^{-21}	1 zeptowatt (zW)	$= 1 \times 10^{-21}$ watt

[a]The 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); $1\,J = 1\,kg \cdot m^2/s^2$ and $1\,W = 1\,J/s$.
[b]Greek 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.

Give It Some Thought

How many picometers are there in 1 mm?

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), which is equal to about 2.2 pounds (lb). 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 1

Which of the following weights would you expect to be suitable for weighing on an ordinary bathroom scale?

(**a**) 2.0×10^7 mg (**b**) 2500 μg (**c**) 5×10^{-4} kg (**d**) 4×10^6 cg (**e**) 5.5×10^8 dg

Practice Exercise 2

(**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 1.19**).

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

$$K = {}^\circ C + 273.15 \qquad [1.3]$$

The freezing point of water, 0 °C, is 273.15 K (Figure 1.19). Notice that we do not use a degree sign (°) with temperatures on the Kelvin scale.

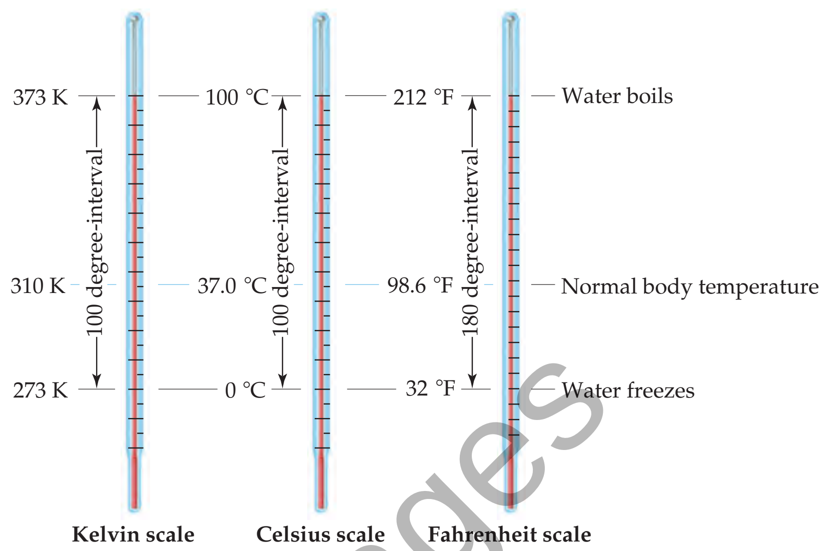
The common temperature scale in the United States is the *Fahrenheit scale*, which is not generally used in science. Water freezes at 32 °F and boils at 212 °F. The Fahrenheit and Celsius scales are related according to

$${}^\circ C = \frac{5}{9}({}^\circ F - 32) \quad \text{or} \quad {}^\circ F = \frac{9}{5}({}^\circ C) + 32 \qquad [1.4]$$

*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

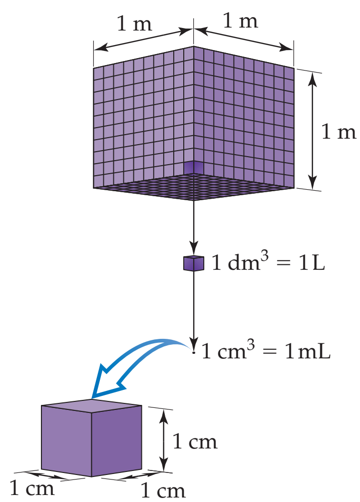
True or false: The "size" of a degree on the Celsius scale is the same as the "size" of a degree on the Kelvin scale.



▲ **Figure 1.19** **Comparison of the Kelvin, Celsius, and Fahrenheit temperature scales.**

Go Figure

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



▲ **Figure 1.20** **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 m^3 = 1000 dm^3. One liter is the same volume as one cubic decimeter, 1 L = 1 dm^3. Each cubic decimeter contains 1000 cubic centimeters, 1 dm^3 = 1000 cm^3. One cubic centimeter equals one milliliter, 1 cm^3 = 1 mL.

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 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 1.20**). 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, and is slightly larger than a quart. (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 mL = 1 cm^3.

Sample Exercise 1.3
Converting Units of Temperature

A weather forecaster predicts the temperature will reach 31 °C. What is this temperature (**a**) in K, (**b**) in °F?

SOLUTION

(**a**) Equation 1.3, we have K = 31 + 273 = 304 K.

(**b**) Using Equation 1.4, we have

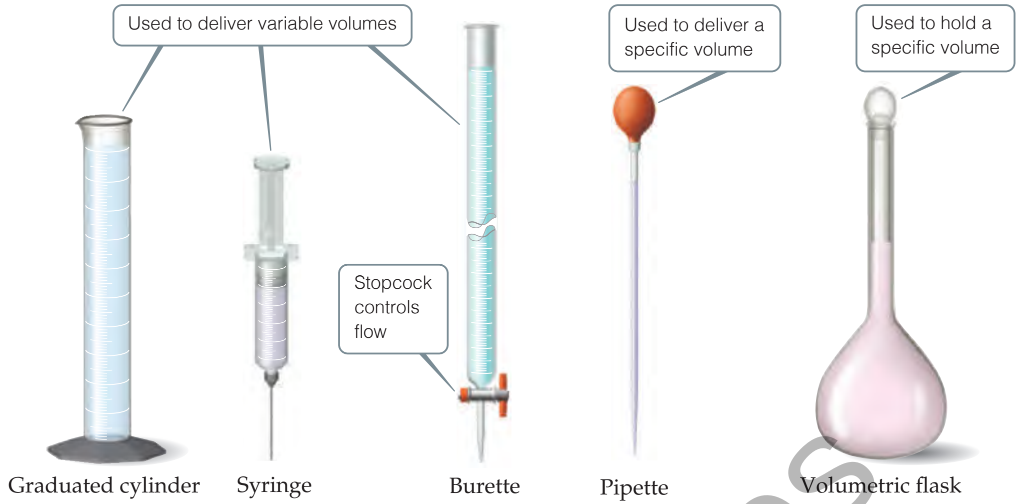
$$°F = \frac{9}{5}(31) + 32 = 56 + 32 = 88\ °F.$$

▶ **Practice Exercise 1**

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.

▶ **Practice Exercise 2**

Ethylene glycol, the major ingredient in antifreeze, freezes at −11.5 °C. What is the freezing point in (**a**) K, (**b**) °F?



▲ **Figure 1.21** **Common volumetric glassware.**

In the lab, you will likely use the devices in **Figure 1.21** 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.

Give It Some Thought

Which of the following quantities represents a volume measurement: 15 m^2; 2.5×10^2 m^3; 5.77 L/s?

Density

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

$$\text{density} = \frac{\text{mass}}{\text{volume}} \quad [1.5]$$

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.

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

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, $1\ J = 1\ kg\text{-}m^2/s^2$. Numerically, a 2-kg mass moving at a velocity of 1 m/s possesses a kinetic energy of 1 J:

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

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 (wood alcohol) if its density is 0.791 g/mL.
(**c**) What is the mass in grams of a cube of gold (density = 19.32 g/cm^3) 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

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

(**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

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

Practice Exercise 1

Platinum, Pt, is one of the rarest of the metals. Worldwide annual production is only about 118 metric tons. Platinum has a density of 21.4 g/cm^3. If thieves were to steal platinum from a bank using a small truck with a maximum payload capacity of 400 kg, how many 1 L bars of the metal could they take?
(**a**) 19 bars (**b**) 2 bars (**c**) 42 bars (**d**) 1 bar (**e**) 47 bars

Practice Exercise 2

(**a**) Calculate the density of a 374.5-g sample of copper if it has a volume of 41.8 cm^3. (**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:

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

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, $g = 9.8$ 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:

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

(**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:

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

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

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 1**

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

▶ **Practice Exercise 2**

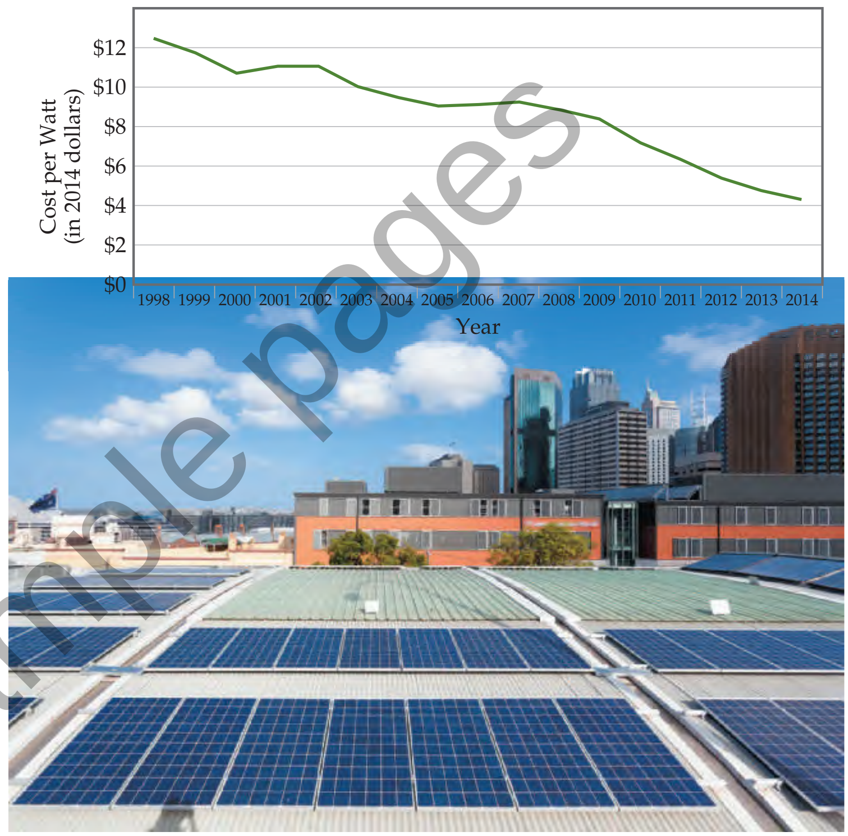
A 355-mL vanilla milkshake at a fast-food restaurant contains 547 Calories. What quantity of energy is this in joules?

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 1.22**).

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.



▲ **Figure 1.22 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.

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.

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 1.23**. 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