CHAPTER 1

Matter, Energy, and Measurement

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A woman climbing a frozen waterfall in British Columbia.

1.1 Why Do We Call Chemistry the Study of Matter?

The world around us is made of chemicals. Our food, our clothing, the buildings in which we live are all made of chemicals. Our bodies are made of chemicals, too. To understand the human body, its diseases, and its cures, we must know all we can about those chemicals. There was a time—only a few hundred years ago—when physicians were powerless to treat many diseases. Cancer, tuberculosis, smallpox, typhus, plague, and many other sicknesses struck people seemingly at random. Doctors, who had no idea what caused any of these diseases, could do little or nothing about them. Doctors treated them with magic as well as by such measures as bleeding, laxatives, hot plasters, and pills made from powdered stag horn, saffron, or gold. None of these treatments was effective, and the doctors, because they came into direct contact with highly contagious diseases, died at a much higher rate than the general public.

A woman being bled by a leech on her left forearm; a bottle of leeches is on the table. From a 1639 woodcut.

Medicine has made great strides since those times. We live much longer, and many once-feared diseases have been essentially eliminated or are curable. Smallpox has been eradicated, and polio, typhus, bubonic plague, diphtheria, and other diseases that once killed millions no longer pose a serious problem, at least not in the developed countries.

How has this medical progress come about? The answer is that diseases could not be cured until they were understood, and this understanding has emerged through greater knowledge of how the body functions. It is progress in our understanding of the principles of biology, chemistry, and physics that has led to these advances in medicine. Because so much of modern medicine depends on chemistry, it is essential that students who intend to enter the health professions have some understanding of basic chemistry. This book was written to help you achieve that goal. Even if you choose a different profession, you will find that the chemistry you learn in this course will greatly enrich your life.

The universe consists of matter, energy, and empty space. **Matter** is anything that has mass and takes up space. **Chemistry** is the science that deals with matter: the structure and properties of matter and the transformations from one form of matter to another. We will discuss energy in Section 1.8.

It has long been known that matter can change, or be made to change, from one form to another. In a **chemical change,** more commonly called a **chemical reaction,** substances are used up (disappear) and others are formed to take their places. An example is the burning of the mixture of hydrocarbons usually called "bottled gas." In this mixture of hydrocarbons, the main component is propane. When this chemical change takes place, propane and oxygen from the air are converted to carbon dioxide and water. Figure 1.1 shows another chemical change.

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Matter also undergoes other kinds of changes, called **physical changes.** These changes differ from chemical reactions in that the identities of the substances do not change. Most physical changes involve changes of state—for example, the melting of solids and the boiling of liquids. Water remains water whether it is in the liquid state or in the form of ice or steam. The conversion from one state to another is a physical—not a chemical—change. Another important type of physical change involves making or separating mixtures. Dissolving sugar in water is a physical change.

When we talk about the **chemical properties** of a substance, we mean the chemical reactions that it undergoes. **Physical properties** are all properties that do not involve chemical reactions. For example, density, color, melting point, and physical state (liquid, solid, gas) are all physical properties.

(*a*) (*b*) (*c*)

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Active Figure 1.1 A chemical reaction. (*a*) Bromine, an orange-brown liquid, and aluminum metal. (*b*) These two substances react so vigorously that the aluminum becomes molten and glows white hot at the bottom of the beaker. The yellow vapor consists of vaporized bromine and some of the product of the reaction, white aluminum bromide. (*c*) Once the reaction is complete, the beaker is coated with aluminum bromide and the products of its reaction with atmospheric moisture. (*Note:* This reaction is dangerous! Under no circumstances should it be done except under properly supervised conditions.) **See a simulation based on this figure, and take a short quiz on the concepts at http://now.brookscole.com/gob8 or on the CD.**

1.2 What Is the Scientific Method?

Scientists learn by using a tool called the **scientific method.** The heart of the scientific method is the testing of theories. It was not always so, however. Before about 1600, philosophers often believed statements just because they sounded right. For example, the great philosopher Aristotle (384–322 BCE) believed that if you took the gold out of a mine it would grow back. He believed this idea because it fitted in with a more general picture that he had about the workings of nature. In ancient times, most thinkers behaved in this way. If a statement sounded right, they believed it without testing it.

About 1600 CE, the scientific method came into use. Let us look at an example to see how the scientific method operates. The Greek physician Galen (200–130 BCE) recognized that the blood on the left side of the heart somehow gets to the right side. This is a fact. A **fact** is a statement based on direct experience. It is a consistent and reproducible observation. Having observed this fact, Galen then proposed a hypothesis to explain it. A **hypothesis** is a statement that is proposed, without actual proof, to explain the facts and their relationship. Because Galen could not actually see how the blood got from the left side to the right side of the heart, he came up with the hypothesis that tiny holes must be present in the muscular wall that separates the two halves.

Up to this point, a modern scientist and an ancient philosopher would behave the same way. Each would offer a hypothesis to explain the facts. From this point on, however, their methods would differ. To Galen, his explanation sounded right and that was enough to make him believe it, even though he couldn't see any holes. His hypothesis was, in fact, believed by virtually all physicians for more than 1000 years. When we use the scientific method, however, we do not believe a hypothesis just because it sounds right. We test it, using the most rigorous testing we can imagine.

Courtesy of the National Library of Medicine

Galen.

Hypothesis A statement that is proposed, without actual proof, to explain a set of facts and their relationship

William Harvey (1578–1657) tested Galen's hypothesis by dissecting human and animal hearts and blood vessels. He discovered that one-way valves separate the upper chambers of the heart from the lower chambers. He also discovered that the heart is a pump that, by contracting and expanding, pushes the blood out. Harvey's teacher, Fabricius (1537–1619), had previously observed that one-way valves exist in the veins, so that blood in the veins can travel only toward the heart and not the other way.

Harvey put these facts together to come up with a new hypothesis: Blood is pumped by the heart and circulates throughout the body. This was a better hypothesis than Galen's because it fitted the facts more closely. Even so, it was still a hypothesis and, according to the scientific method, had to be tested further. One important test took place in 1661, four years after Harvey died. Harvey had predicted that because there had to be a way for the blood to get from the arteries to the veins, tiny blood vessels must connect them. In 1661 the Italian anatomist Malpighi (1628–1694), using the newly invented microscope, found these tiny vessels, which are now called capillaries.

Malpighi's discovery supported the blood circulation hypothesis by fulfilling Harvey's prediction. When a hypothesis passes the tests, we have more confidence in it and call it a theory. A **theory** is the formulation of an apparent relationship among certain observed phenomena, which has been verified to some extent. In this sense, a theory is the same as a hypothesis except that we have a stronger belief in it because more evidence supports it. No matter how much confidence we have in a theory, however, if we discover new facts that conflict with it or if it does not pass newly devised tests, the theory must be altered or rejected. In the history of science, many firmly established theories have eventually been thrown out because they could not pass new tests.

Theory The formulation of an apparent relationship among certain observed phenomena, which has been verified. A theory explains many interrelated facts and can be used to make predictions about natural phenomena. Examples are Newton's theory of gravitation and the kinetic molecular theory of gases, which we will encounter in Section 6.6. This type of theory is also subject to testing and will be discarded or modified if it is contradicted by new facts.

The scientific method is thus very simple. We don't accept a hypothesis or a theory just because it sounds right. We devise tests, and only if the hypothesis or theory passes the tests do we accept it. The enormous progress made since 1600 in chemistry, biology, and the other sciences is a testimony to the value of the scientific method.

You may get the impression from the preceding discussion that science progresses in one direction: facts first, hypothesis second, theory last. Real life is not so simple, however. Hypotheses and theories call the attention of scientists to discover new facts. An example of this scenario is the discovery of the element germanium. In 1871, Mendeleev's Periodic Table—a graphic description of elements organized by properties—predicted the existence of a new element whose properties would be similar to those of silicon. Mendeleev called this element eka-silicon. In 1886, it was discovered in Germany (hence the name), and its properties were truly similar to those predicted by theory.

On the other hand, many scientific discoveries result from **serendipity,** or chance observation. An example of serendipity occurred in 1926, when James Sumner of Cornell University left an enzyme preparation of jack bean urease in a refrigerator over the weekend. Upon his return, he found that his solution contained crystals that turned out to be a protein. This chance discovery led to the hypothesis that all enzymes are proteins. Of course, serendipity is not enough to move science forward. Scientists must have the creativity and insight to recognize the significance of their observations. Sumner fought for more than 15 years for his hypothesis to gain acceptance because people believed that only small molecules can form crystals. Eventually his view won out, and he was awarded a Nobel Prize in chemistry in 1946.

1.3 How Do Scientists Report Numbers?

Scientists often have to deal with numbers that are very large or very small. For example, an ordinary copper penny (dating from before 1982, when pennies in the United States were still made of copper) contains approximately

$$29{,}500{,}000{,}000{,}000{,}000{,}000{,}000 \text{ atoms of copper}$$

and a single copper atom weighs

$$0.0000000000000000000000000023 \text{ pound}$$

which is equal to

$$0.0000000000000000000000104 \text{ gram}$$

Many years ago, an easy way to handle such large and small numbers was devised. This method, which is called **exponential notation,** is based on powers of 10. In exponential notation, the number of copper atoms in a penny is written

$$2.95 \times 10^{22}$$

and the weight of a single copper atom is written

$$2.3 \times 10^{-25} \text{ pound}$$

which is equal to

$$1.04 \times 10^{-22} \text{ grams}$$

The origin of this shorthand form can be seen in the following examples:

$$100 = 1 \times 10 \times 10 = 1 \times 10^2$$

$$1000 = 1 \times 10 \times 10 \times 10 = 1 \times 10^3$$

What we have just said in the form of an equation is "100 is a one with two zeroes after the one, and 1000 is a one with three zeroes after the one." We can also write

$$1/100 = 1/10 \times 1/10 = 1 \times 10^{-2}$$

$$1/1000 = 1/10 \times 1/10 \times 1/10 = 1 \times 10^{-3}$$

where negative exponents denote numbers less than 1. The exponent in a very large or very small number lets us keep track of the number of zeros. That number can become unwieldy with very large or very small quantities, and it is easy to lose track of a zero. Exponential notation helps us deal with this possible source of mathematical error.

When it comes to measurements, not all the numbers you can generate in your calculator or computer are of equal importance. Only the number of digits that are known with certainty are significant. Suppose that you measured the weight of an object as 3.4 g on a balance that you can read to the nearest 0.1 g. You can report the weight as 3.4 g but not as 3.40 or 3.400 g because you do not know the added zeros with certainty. This becomes even more important when you do calculations using a calculator. For example, you might measure a cube with a ruler and find that each side is 2.9 cm. If you are asked to calculate the volume, you multiply $2.9 \times 2.9 \times 2.9$. The calculator will then give you an answer that is 23.389 cm^3. However, your initial measurements were only good to a tenth of a centimeter, so your final answer cannot be good to a thousandth of a centimeter. As a scientist, it is important to report data that have the correct number of **significant figures.** A detailed account of using significant figures is presented in Appendix II. A discussion of accuracy, precision, and significant figures can be found in laboratory manuals [see Bettelheim and Landesberg, *Laboratory Experiments,* sixth edition (Experiment 2)].

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EXAMPLE 1.1

Multiply:

(a) $(4.73 \times 10^5)(1.37 \times 10^2)$ (b) $(2.7 \times 10^{-4})(5.9 \times 10^8)$

Divide:

(c) $\dfrac{7.08 \times 10^{-8}}{300}$ (d) $\dfrac{5.8 \times 10^{-6}}{6.6 \times 10^{-8}}$ (e) $\dfrac{7.05 \times 10^{-3}}{4.51 \times 10^5}$

Solution

The way to do calculations of this sort is to use a button on scientific calculators that automatically uses exponential notation. The button is usually labeled "E."

(a) Enter 4.73E5, press the multiplication key, enter 1.37E2, and press the "=" key. The answer is 6.48×10^7. The calculator will display this number as 6.48E7. This answer makes sense. We add exponents when we multiply, and the sum of these two exponents is correct $(5 + 2 = 7)$. We also multiply the numbers, 4.73×1.37. This is approximately $4 \times 1.5 = 6$, so 6.48 is also reasonable.

(b) Here we have to deal with a negative exponent, so we use the "+/-" key. Enter 2.7E+/-4, press the multiplication key, enter 5.9E8, and press the "=" key. The calculator will display the answer as 1.593E5. To have the correct number of significant figures, we should report our answer as 1.6E5. This answer makes sense because 2.7 is a little less than 3, and 5.9 is a little less than 6, so we predict a number slightly less than 18; also the algebraic sum of the exponents $(-4 + 8)$ is equal to 4. This gives 16×10^4. In scientific notation, we normally prefer to report numbers between 1 and 10, so we rewrite our answer as 1.6×10^5. We made the first number 10 times smaller, so we increased the exponent by 1 to reflect that change.

(c) Enter 7.08E+/-8, press the division key, enter 300, and press the "=" key. The answer is 2.36×10^{-10}. The calculator will display this number as 2.36E − 10. We subtract exponents when we divide, and we can also write 300 as 3.00×10^2.

(d) Enter 5.8E+/-6, press the division key, enter 6.6E+/-8, and press the "=" key. The calculator will display the answer as 87.878787878788. We report this answer as 88 to get the right number of significant figures. This answer makes sense. When we divide 5.8 by 6.6, we get a number slightly less than 1. When we subtract the exponents algebraically $(-6 - [-8])$, we get 2. This means that the answer is slightly less than 1×10^2, or slightly less than 100.

(e) Enter 7.05E+/-3, press the division key, enter 4.51E5, and press the "=" key. The calculator displays the answer as 1.5632E-8, which, to the correct number of significant figures, is 1.56×10^{-8}. The algebraic subtraction of exponents is $-3 - 5 = -8$.

Problem 1.1

Multiply:

(a) $(6.49 \times 10^{7})(7.22 \times 10^{-3})$ (b) $(3.4 \times 10^{-5})(8.2 \times 10^{-11})$

Divide:

(a) $\dfrac{6.02 \times 10^{23}}{3.10 \times 10^{5}}$ (b) $\dfrac{3.14}{2.30 \times 10^{-5}}$

The label on this bottle of water shows the metric size (one liter) and the eqivalent in quarts.

1.4 How Do We Make Measurements?

In our daily lives we are constantly making measurements. We measure ingredients for recipes, driving distances, gallons of gasoline, weights of fruits and vegetables, and the timing of TV programs. Doctors and nurses measure pulse rates, blood pressures, temperatures, and drug dosages. Chemistry, like other sciences, is based on measurements.

A measurement consists of two parts: a number and a unit. A number without a unit is usually meaningless. If you were told that a person's weight is 57, the information would be of very little use. Is it 57 pounds, which would indicate that the person is very likely a child or a midget, or 57 kilograms, which is the weight of an average woman or a small man? Or is it perhaps some other unit? Because so many units exist, a number by itself is not enough; the unit must also be stated.

In the United States, most measurements are made with the English system of units: pounds, miles, gallons, and so on. In most other parts of the world, however, few people could tell you what a pound or an inch is. Most countries use the **metric system,** a system that originated in France about 1800 and that has since spread throughout the world. Even in the United States, metric measurements are slowly being introduced (Figure 1.2). For

Metric system A system of units of measurement in which the divisions to subunits are made by a power of 10

Figure 1.2 Road sign in Massachusetts showing metric equivalents of mileage.

example, many soft drinks and most alcoholic beverages now come in metric sizes. Scientists in the United States have been using metric units all along.

Around 1960, international scientific organizations adopted another system, called the **International System of Units** (abbreviated **SI**). The SI is based on the metric system and uses some of the metric units. The main difference is that the SI is more restrictive: It discourages the use of certain metric units and favors others. Although the SI has advantages over the older metric system, it also has significant disadvantages. For this reason U.S. chemists have been very slow to adopt it. At this time, approximately 40 years after its introduction, not many U.S. chemists use the entire SI, although some of its preferred units are gaining ground.

In this book we will use the metric system (Table 1.1). Occasionally we will mention the preferred SI unit.

Table 1.1 Base Units in the Metric System

Length	meter (m)
Volume	liter (L)
Mass	gram (g)
Time	second (s)
Temperature	°Celsius (°C)
Energy	calorie (cal)
Amount of substance	mole (mol)

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A. Length

The key to the metric system (and the SI) is that there is one base unit for each kind of measurement and that other units are related to the base unit only by powers of 10. As an example, let us look at measurements of length. In the English system we have the inch, the foot, the yard, and the mile (not to mention such older units as the league, furlong, ell, and rod). If you want to convert one unit to another unit, you must memorize or look up these conversion factors:

$$5280 \text{ feet} = 1 \text{ mile}$$

$$1760 \text{ yards} = 1 \text{ mile}$$

$$3 \text{ feet} = 1 \text{ yard}$$

$$12 \text{ inches} = 1 \text{ foot}$$

All this is unnecessary in the metric system (and the SI). In both systems the base unit of length is the **meter (m).** To convert to larger or smaller units we do not use arbitrary numbers like 12, 3, and 1760, but only 10, 100, 1/100, 1/10, or other powers of 10. This means that *to convert from one metric or SI unit to another, we only have to move the decimal point.* Furthermore, the other units are named by putting prefixes in front of "meter," and *these prefixes are the same throughout the metric system and the SI.* Table 1.2 lists the most important of these prefixes. If we put some of these prefixes in front of "meter," we have

$$1 \text{ kilometer (km)} = 1000 \text{ meters (m)}$$

$$1 \text{ centimeter (cm)} = 0.01 \text{ meter}$$

$$1 \text{ nanometer (nm)} = 10^{-9} \text{ meter}$$

Table 1.2 The Most Common Metric Prefixes

Prefix	Symbol	Value
giga	G	10^9 = 1,000,000,000 (one billion)
mega	M	10^6 = 1,000,000 (one million)
kilo	k	10^3 = 1000 (one thousand)
deci	d	10^{-1} = 0.1 (one-tenth)
centi	c	10^{-2} = 0.01 (one-hundredth)
milli	m	10^{-3} = 0.001 (one-thousandth)
micro	μ	10^{-6} = 0.000001 (one-millionth)
nano	n	10^{-9} = 0.000000001 (one-billionth)

Table 1.3 Some Conversion Factors Between the English and Metric Systems

Length	Mass	Volume
1 in. = 2.54 cm	1 oz = 28.35 g	1 qt = 0.946 L
1 m = 39.37 in	1 lb = 453.6 g	1 gal = 3.785 L
1 mile = 1.609 km	1 kg = 2.205 lb	1 L = 33.81 fl oz
	1 g = 15.43 grains	1 fl oz = 29.57 mL
		1 L = 1.057 qt

For people who have grown up using English units, it is helpful to have some idea of the size of metric units. Table 1.3 shows some conversion factors.

Some of these conversions are difficult enough that you will probably not remember them and must, therefore, look them up when you need them. Some are easier. For example, a meter is about the same as a yard. A kilogram is a little over two pounds. There are almost four liters in a gallon. These conversions may be important to you someday. For example, if you rent a car in Europe, the price of gas listed on the sign at the gas station will be in Euros per liter. When you realize that you are spending a dollar per liter and you know that there are almost four liters to a gallon, you will realize why so many people take the bus or a train instead.

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Hypodermic syringe. Note that the volumes are indicated in mL.

B. Volume

Volume is space. The volume of a liquid, solid, or gas is the space occupied by that substance. The base unit of volume in the metric system is the **liter (L).** This unit is a little larger than a quart (Table 1.3). The only other common metric unit for volume is the milliliter (mL), which is equal to 10^{-3} L.

$$1 \text{ mL} = 0.001 \text{ L}$$
$$1000 \text{ mL} = 1 \text{ L}$$

One milliliter is exactly equal to one cubic centimeter (cc or cm^3):

$$1 \text{ mL} = 1 \text{ cc}$$

Thus there are 1000 cc in 1 L.

C. Mass

Mass is the quantity of matter in an object. The base unit of mass in the metric system is the **gram (g).** As always in the metric system, larger and smaller units are indicated by prefixes. The ones in common use are

$$1 \text{ kilogram (kg)} = 1000 \text{ g}$$
$$1 \text{ milligram (mg)} = 0.001 \text{ g}$$

The gram is a small unit; there are 453.6 g in one pound (Table 1.3).

We use a device called a balance to measure mass. Figure 1.3 shows two types of laboratory balances.

There is a fundamental difference between mass and weight. Mass is independent of location. The mass of a stone, for example, is the same whether we measure it at sea level, on top of a mountain, or in the depths of a mine. In contrast, weight is not independent of location. **Weight** is the

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Figure 1.3 Two laboratory balances.

force a mass experiences under the pull of gravity. This point was dramatically demonstrated when the astronauts walked on the surface of the Moon. The Moon, being a smaller body than the Earth, exerts a weaker gravitational pull. Consequently, even though the astronauts wore space suits and equipment that would be heavy on Earth, they felt lighter on the Moon and could execute great leaps and bounces during their walks.

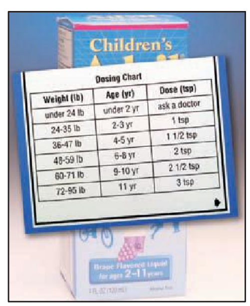
Although mass and weight are different concepts, they are related to each other by the force of gravity. We frequently use the words interchangeably because we weigh objects by comparing their masses to standard refer-

Chemical Connections 1A

Drug Dosage and Body Mass

In many cases, drug dosages are prescribed on the basis of body mass. For example, the recommended dosage of a drug may be 3 mg of drug for each kilogram of body weight. In this case, a 50 kg (110 lb) woman would receive 150 mg and an 82 kg (180 lb) man would get 246 mg. This adjustment is especially important for children, because a dose suitable for an adult will generally be too much for a child, who has much less body mass. For this reason, manufacturers package and sell smaller doses of certain drugs, such as aspirin, for children.

Drug dosage may also vary with age. Occasionally, when an elderly patient has an impaired kidney or liver function, the clearance of a drug from the body is delayed, and the drug may stay in the body longer than is normal. This persistence can cause dizziness, vertigo, and migraine-like headaches, resulting in falls and broken bones. Such delayed clearance must be monitored, and the drug dosage adjusted accordingly.



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This package of Advil has a chart showing the proper doses for children of a given weight.

ence masses (weights) on a balance, and the gravitational pull is the same on the unknown object and on the standard masses. Because the force of gravity is essentially constant, mass is always directly proportional to weight.

D. Time

Time is the one quantity for which the units are the same in all systems: English, metric, and SI. The base unit is the **second (s):**

$$60 \text{ s} = 1 \text{ min}$$

$$60 \text{ min} = 1 \text{ h}$$

E. Temperature

Most people in the United States are familiar with the Fahrenheit scale of temperature. The metric system uses the centigrade, or **Celsius,** scale. In this scale, the boiling point of water is set at 100°C and the freezing point at 0°C. We can convert from one scale to the other by using the following formulas:

$$°F = \frac{9}{5}°C + 32$$

$$°C = \frac{5}{9}(°F - 32)$$

The 32 in these equations is a defined number and is, therefore, treated as if it had an infinite number of zeros following the decimal point. (See Appendix II.)

EXAMPLE 1.2

Normal body temperature is 98.6°F. Convert this temperature to Celsius.

Solution

$$°C = \frac{5}{9}(98.6 - 32) = \frac{5}{9}(66.6) = 37.0°C$$

Problem 1.2

Convert:
(a) 64.0°C to Fahrenheit (b) 47°F to Celsius

Figure 1.4 shows the relationship between the Fahrenheit and Celsius scales.

A third temperature scale is the **Kelvin (K)** scale, also called the absolute scale. The size of a Kelvin degree is the same as that of a Celsius degree; the only difference is the zero point. The temperature −273°C is taken as the zero point on the Kelvin scale. This makes conversions between Kelvin and Celsius very easy. To go from Celsius to Kelvin, just *add* 273; to go from Kelvin to Celsius, *subtract* 273:

$$K = °C + 273$$

$$°C = K - 273$$

Figure 1.4 also shows the relationship between the Kelvin and Celsius scales. Note that we don't use the degree symbol in the Kelvin scale: 100°C equals 373 K, not 373°K.

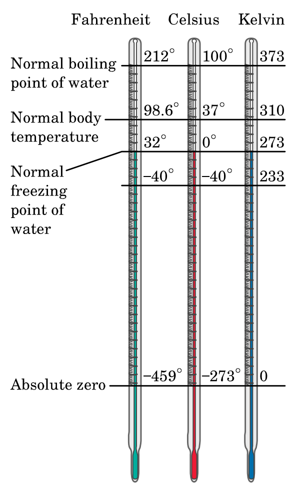

Figure 1.4 Three temperature scales.

Why was −273°C chosen as the zero point on the Kelvin scale? The reason is that *−273°C, or 0 K, is the lowest possible temperature*. Because of this, 0 K is called **absolute zero.** Temperature reflects how fast molecules move. The more slowly they move, the colder it gets. At absolute zero, molecules stop moving altogether. Therefore, the temperature cannot get any lower. For some purposes it is convenient to have a scale that begins at the lowest possible temperature; the Kelvin scale fulfills this need. The Kelvin is the SI unit.

It is very important to have a "gut feeling" about the relative sizes of the units in the metric system. Often, while doing calculations, the only thing that might offer a clue that you have made an error is your understanding of the sizes of the units. For example, if you are calculating the amount of a chemical that is dissolved in water and you come up with an answer of 254 kg/mL, does your answer make sense? If you have no intuitive feeling about the size of a kilogram or a milliliter, you will not know. If you realize that a milliliter is about the volume of a thimble and that a standard bag of sugar might weigh 2 kg, then you will realize that there is no way to pack 254 kg into a thimble of water, and you will know that you made a mistake.

1.5 What Is a Handy Way to Convert from One Unit to Another?

Factor-label method A procedure in which the equations are set up so that all the unwanted units cancel and only the desired units remain

Conversion factor A ratio of two different units

We frequently need to convert a measurement from one unit to another. The best and most foolproof way to do this is the **factor-label method.** In this method we follow the rule that *when multiplying numbers we also multiply units, and when dividing numbers we also divide units.*

For conversions between one unit and another, it is always possible to set up two fractions, called **conversion factors.** Suppose we wish to convert the weight of an object from 381 grams to pounds. We are converting the units, but we are not changing the object itself. We want a ratio that reflects the change in units. In Table 1.3, we see that there are 453.6 grams in 1 pound. That is, the amount of matter in 453.6 grams is the same as the amount in 1 pound. In that sense, it is a one-to-one ratio, even though the units are not numerically the same. The conversion factors between grams and pounds therefore are

$$\frac{1\text{ lb}}{453.6\text{ g}} \quad \text{and} \quad \frac{453.6\text{ g}}{1\text{ lb}}$$

To convert 381 grams to pounds, we must multiply by the proper conversion factor—but which one? Let us try both and see what happens.

First let us multiply by 1 lb/453.6 g:

$$381\text{ }\cancel{\text{g}} \times \frac{1\text{ lb}}{453.6\text{ }\cancel{\text{g}}} = 0.840\text{ lb}$$

Following the procedure of multiplying and dividing units when we multiply and divide numbers, we find that dividing grams by grams cancels out the grams. We are left with pounds, which is the answer we want. Thus 1 lb/453.6 g is the correct conversion factor because it converts grams to pounds.

Suppose we had done it the other way, multiplying by 453.6 g/1 lb:

$$381\text{ g} \times \frac{453.6\text{ g}}{1\text{ lb}} = 173{,}000\,\frac{\text{g}^2}{\text{lb}}$$

When we multiply grams by grams, we get g^2 (grams squared). Dividing by pounds gives g^2/lb. This is not the unit we want, so we used the incorrect conversion factor.

HOW TO . . .

Do Unit Conversions by the Factor-Label Method

One of the most useful ways of approaching conversions is to ask three questions:

- What information am I given? This is the starting point.
- What do I want to know? This is the answer that you want to find.
- What is the connection between the first two? This is the conversion factor. Of course, more than one conversion factor may be needed for some problems.

Let's look at how to apply these principles to a conversion from pounds to kilograms. Suppose we want to know the weight in kilograms of a woman who weighs 125 lb. We see in Table 1.3 that there are 2.205 lb in 1 kg. Note that we are starting out with pounds and we want an answer in kilograms.

$$125\ \cancel{\text{lb}} \times \frac{1\ \text{kg}}{2.205\ \cancel{\text{lb}}} = 56.7\ \text{kg}$$

- The weight in pounds is the starting point. We were given that information.
- We wanted to know the weight in kilograms. That was the desired answer, and we found the number of kilograms.
- The connection between the two is the conversion factor in which the unit of the desired answer is in the numerator of the fraction, rather than the denominator. It is not simply a mechanical procedure to set up the equation so that units cancel; it is a first step to understanding the underlying reasoning behind the factor-label method. If you set up the equation to give the desired unit as the answer, you have made the connection properly.

If you apply this kind of reasoning, you can always pick the right conversion factor. Given the choice between

$$\frac{2.205\ \text{lb}}{1\ \text{kg}} \quad \text{and} \quad \frac{1\ \text{kg}}{2.205\ \text{lb}}$$

you know that the second conversion factor will give an answer in kilograms, so you use it. When you check the answer, you see that it is reasonable. You expect a number that is about one half of 125, which is 62.5. The actual answer, 56.7, is close to that value. The number of pounds and the number of kilograms are not the same, but they represent the same weight. That fact makes the use of conversion factors logically valid; the factor-label method uses the connection to obtain a numerical answer.

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Click *Mastering the Essentials* to practice using the **Factor-Label Method**

The advantage of the factor-label method is that it lets us know when we have made an incorrect calculation. *If the units of the answer are not the ones we are looking for, the calculation must be wrong.* Incidentally, this principle works not only in unit conversions but in all problems where we make calculations using measured numbers.

The factor-label method gives the correct mathematical solution for a problem. However, it is a mechanical technique and does not require you to think through the problem. Thus it may not provide a deeper understanding. For this reason and also to check your answer (because it is easy to make mistakes in arithmetic—for example, by punching the wrong numbers into a calculator), you should always ask yourself if the answer you have obtained is reasonable. For example, the question might ask the mass of a single oxygen atom. If your answer comes out 8.5×10^6 g, it is not reasonable. A single atom cannot weigh more than you do! In such a case, you have obviously made a mistake and should take another look to see where you went wrong. Of course, everyone makes mistakes at times, but if you check you can at least determine whether your answer is reasonable. If it is not, you will immediately know that you have made a mistake and can then correct it.

Checking whether an answer is reasonable gives you a deeper understanding of the problem because it forces you to think through the relationship between the question and the answer. The concepts and the mathematical relationships in these problems go hand in hand. Mastery of the mathematical skills makes the concepts clearer, and insight into the concepts suggests ways to approach the mathematics. We will now give a few examples of unit conversions and then test the answers to see whether they are reasonable. To save space, we will practice this technique mostly in this chapter, but you should use a similar approach in all later chapters.

In unit conversion problems, you should always check two things. First, the numeric factor by which you multiply tells you whether the answer will be larger or smaller than the number being converted. Second, the factor tells you how much greater or smaller than the number you start with your answer should be. For example, if 100 kg is converted to pounds and there are 2.205 lb in 1 kg, then an answer of about 200 is reasonable—but an answer of 0.2 or 2000 is not.

EXAMPLE 1.3

The distance between Rome and Milan (the largest cities in Italy) is 358 miles. How many kilometers separate the two?

Solution

We want to convert miles to kilometers. Table 1.3 shows that 1 mi = 1.609 km. From this we get two conversion factors:

$$\frac{1 \text{ mi}}{1.609 \text{ km}} \quad \text{and} \quad \frac{1.609 \text{ km}}{1 \text{ mi}}$$

Which should we use? We use the one that gives the answer in kilometers:

$$358 \text{ mi} \times \text{conversion factor} = ? \text{ km}$$

This means that the miles must cancel, so the conversion factor 1.609 km/1 mi is appropriate.

$$358 \cancel{\text{mi}} \times \frac{1.609 \text{ km}}{1 \cancel{\text{mi}}} = 576 \text{ km}$$

Is this answer reasonable? We want to convert a given distance in miles to the same distance in kilometers. The conversion factor in Table 1.3 tells us that in a given distance the number of kilometers is larger than the number of miles. How much larger? The actual number is 1.609, which is approximately 1.5 times larger. Thus we expect that the answer in kilometers will be about 1.5 times greater than the number given in miles. The number given in miles is 358, which, *for the purpose of checking whether our answer is reasonable,* we can round off to, say, 400. Multiplying this number by 1.5 gives an approximate answer of 600 km. Our actual answer, 576 km, was of the same order of magnitude as the estimated answer, so we can say that it is reasonable. If the estimated answer had been 6 km, or 60 km, or 6000 km, we would suspect that we had made a mistake in calculating the actual answer.

Problem 1.3

How many kilograms are in 241 lb? Check your answer to see if it is reasonable.

EXAMPLE 1.4

The label on a container of olive oil says 1.844 gal. How many milliliters does the container hold?

Solution

Table 1.3 shows no factor for converting gallons to milliliters, but it does show that 1 gal = 3.785 L. Because we know that 1000 mL = 1 L, we can solve this problem by multiplying by two conversion factors, making certain that all units cancel except milliliters:

$$1.844\ \cancel{\text{gal}} \times \frac{3.785\ \cancel{\text{L}}}{1\ \cancel{\text{gal}}} \times \frac{1000\ \text{mL}}{1\ \cancel{\text{L}}} = 6980\ \text{mL}$$

Is this answer reasonable? The conversion factor in Table 1.3 tells us that there are more liters in a given volume than gallons. How much more? Approximately four times more. We also know that any volume in milliliters is 1000 times larger than the same volume in liters. Thus we expect that the volume expressed in milliliters will be 4 × 1000, or 4000, times more than the volume given in gallons. The estimated volume in milliliters will be approximately 1.8 × 4000, or 7000 mL. But we also expect that the actual answer should be somewhat less than the estimated figure because we overestimated the conversion factor (4 rather than 3.785). Thus the answer, 6980 mL, is quite reasonable. Note that the answer is given to four significant figures.

Problem 1.4

Calculate the number of kilometers in 8.55 miles. Check your answer to see whether it is reasonable.

EXAMPLE 1.5

The maximum speed limit on many roads in the United States is 65 mi/h. How many meters per second (m/s) is this speed?

Solution

Here we have essentially a double conversion problem: We must convert miles to meters and hours to seconds. We use as many conversion factors as necessary, always making sure that we use them in such a way that the proper units cancel:

$$65\ \frac{\cancel{\text{mi}}}{\cancel{\text{h}}} \times \frac{1.609\ \cancel{\text{km}}}{1\ \cancel{\text{mi}}} \times \frac{1000\ \text{m}}{1\ \cancel{\text{km}}} \times \frac{1\ \cancel{\text{h}}}{60\ \cancel{\text{min}}} \times \frac{1\ \cancel{\text{min}}}{60\ \text{s}} = 29\ \frac{\text{m}}{\text{s}}$$

Is this answer reasonable? To estimate the 65 mi/h speed in meters per second, we must first establish the relationship between miles and meters. As in Example 1.3, we know that there are more kilometers than miles in a given distance. How much more? As there are approximately 1.5 km in 1 mi, there must be approximately 1500 times more meters. We also know that in 1 hour there are $60 \times 60 = 3600$ seconds. The ratio of meters to seconds will be approximately 1500/3600, which is about one half. Therefore, we estimate that the speed in meters per second will be about one half of that in miles per hour or 32 m/s. Once again, the actual answer, 29 m/s, is not far from the estimate of 32 m/s, so the answer is reasonable.

As shown in these examples, when canceling units we do not cancel the numbers. The numbers are multiplied and divided in the ordinary way.

Estimating the answer is a good thing to do when working any mathematical problem, not just unit conversions.

Problem 1.5

Convert the speed of sound, 332 m/s to mi/h. Check your answer to see whether it is reasonable.

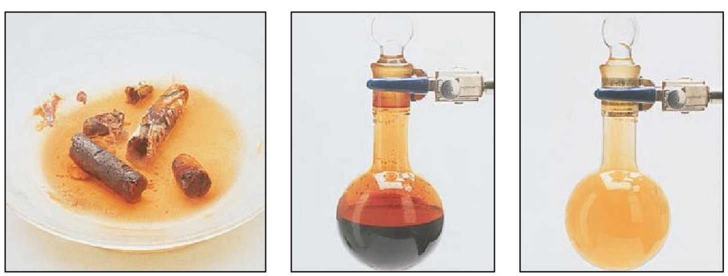
GOB ChemistryNow™

Click *Coached Problems* to practice doing unit conversions by the **Factor-Label Method**

1.6 What Are the States of Matter?

Matter can exist in three states: gas, liquid, and solid. **Gases** have no definite shape or volume. They expand to fill whatever container they are put into. On the other hand, they are highly compressible and can be forced into small containers. **Liquids** also have no definite shape, but they do have a definite volume that remains the same when they are poured from one container to another. Liquids are only slightly compressible. **Solids** have definite shapes and definite volumes. They are essentially incompressible.

Whether a substance is a gas, a liquid, or a solid depends on its temperature and pressure. On a cold winter day, a puddle of liquid water turns to ice; it becomes a solid. If we heat water in an open pot at sea level, the liquid boils at 100°C; it becomes a gas—we call it steam. If we heated the same pot of water on top of Mount Everest, it would boil at about 70°C due to the reduced atmospheric pressure. Most substances can exist in the three states: They are gases at high temperature, liquids at a lower temperature, and solids when their temperature becomes low enough. Figure 1.5 shows a single substance in the three different states.



(*a*) (*b*) (*c*)

Charles D. Winters

Figure 1.5 The three states of matter for bromine: (*a*) bromine as a solid, (*b*) bromine as a liquid, and (*c*) bromine as a gas.

The chemical identity of a substance does not change when it is converted from one state to another. Water is still water whether it is in the form of ice, steam, or liquid water. We discuss the three states of matter, and the changes between one state and another, at greater length in Chapter 6.

1.7 What Are Density and Specific Gravity?

A. Density

One of the many pollution problems that the world faces is the spillage of petroleum into the oceans from oil tankers or from offshore drilling. When oil spills into the ocean, it floats on top of the water. The oil doesn't sink because it is not soluble in water and because water has a higher density than oil. When two liquids are mixed (assuming that one does not dissolve in the other), the one of lower density floats on top (Figure 1.6).

The **density** of any substance is defined as its *mass per unit volume*. Not only do all liquids have a density, but so do all solids and gases. Density is calculated by dividing the mass of a substance by its volume:

$$d = \frac{m}{V} \quad d = \text{density}, \quad m = \text{mass}, \quad V = \text{volume}$$

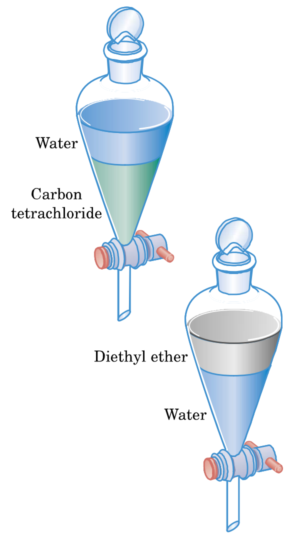


Figure 1.6 Two separatory funnels containing water and another liquid. The density of carbon tetrachloride is 2.961 g/mL, that of water is 1.00 g/mL, and that of diethyl ether is 0.713 g/mL. In each case the liquid with the lower density is on top.

EXAMPLE 1.6

If 73.2 mL of a liquid has a mass of 61.5 g, what is its density in g/mL?

Solution

$$d = \frac{m}{V} = \frac{61.5\ \text{g}}{73.2\ \text{mL}} = 0.840\ \frac{\text{g}}{\text{mL}}$$

Problem 1.6

The density of titanium is 4.54 g/mL. What is the mass, in grams, of 17.3 mL of titanium? Check your answer to see whether it is reasonable.

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The spillage of more than 10 million gallons of petroleum in Prince William Sound, Alaska, in March 1989 caused a great deal of environmental damage.

EXAMPLE 1.7

The density of iron is 7.86 g/cm^3. What is the volume in milliliters of an irregularly shaped piece of iron that has a mass of 524 g?

Solution

Here we are given the mass and the density. In this type of problem, it is useful to derive a conversion factor from the density. Since 1 cm^3 is exactly 1 mL, we know that the density is 7.86 g/mL. This means that 1 mL of iron has a mass of 7.86 g. From this we can get two conversion factors:

$$\frac{1\ \text{mL}}{7.86\ \text{g}} \quad \text{and} \quad \frac{7.86\ \text{g}}{1\ \text{mL}}$$

As usual, we multiply the mass by whichever conversion factor results in the cancellation of all but the correct unit:

$$524\ \cancel{g} \times \frac{1\ \text{mL}}{7.86\ \cancel{g}} = 66.7\ \text{mL}$$

Is this answer reasonable? The density of 7.86 g/mL tells us that the volume in milliliters of any piece of iron is always less than its mass in grams. How much less? Approximately eight times less. Thus we expect the volume to be approximately 500/8 = 63 mL. As the actual answer is 66.7 mL, it is reasonable.

Problem 1.7

An unknown substance has a mass of 56.8 g and occupies a volume of 23.4 mL. What is its density in g/mL? Check your answer to see whether it is reasonable.

The density of any liquid or solid is a physical property that is constant, which means that it always has the same value at a given temperature. We use physical properties to help identify a substance. For example, the density of chloroform (a liquid formerly used as an inhalation anesthetic) is 1.483 g/mL at 20°C. If we want to find out if an unknown liquid is chloroform, one thing we might do is measure its density at 20°C. If the density is, say, 1.355 g/mL, we know the liquid isn't chloroform. If the density is 1.483 g/mL, we cannot be sure the liquid is chloroform, because other liquids might also have this density, but we can then measure other physical properties (the boiling point, for example). If all the physical properties we measure match those of chloroform, we can be reasonably sure the liquid is chloroform.

We have said that the density of a pure liquid or solid is a constant at a given temperature. Density does change when the temperature changes. Almost always, density decreases with increasing temperature. This is true because mass does not change when a substance is heated, but volume almost always increases because atoms and molecules tend to get farther apart as the temperature increases. Since $d = m/V$, if m stays the same and V gets larger, d must get smaller.

The most common liquid, water, provides a partial exception to this rule. As the temperature increases from 4°C to 100°C, the density of water does decrease, but from 0°C to 4°C, the density increases. That is, water has its maximum density at 4°C. This anomaly and its consequences are due to the unique structure of water and will be discussed in Chemical Connections 6E.

B. Specific Gravity

Because density is equal to mass divided by volume, it always has units, most commonly g/mL or g/cc or (g/L for gases). **Specific gravity** is numerically the same as density, but it has no units (it is dimensionless). The reason is that specific gravity is defined as a comparison of the density of a substance with the density of water, which is taken as a standard. For example, the density of copper at 20°C is 8.92 g/mL. The density of water at the same temperature is 1.00 g/mL. Therefore, copper is 8.92 times as dense as water, and its specific gravity at 20°C is 8.92. Because water is taken as the standard and because the density of water is 1.00 g/mL at 20°C, the specific gravity of any substance is always numerically equal to its density, provided that the density is measured in g/mL or g/cc.

Specific gravity is often measured by a hydrometer. This simple device consists of a weighted glass bulb that is inserted into a liquid and allowed to float. The stem of the hydrometer has markings, and the specific gravity is read where the meniscus (the curved surface of the liquid) hits the marking. The specific gravity of the acid in your car battery and that of a urine sample in a clinical laboratory are measured by hydrometers. A hydrometer measuring a urine sample is also called urinometer (Figure 1.7). Normal urine can vary in specific gravity from about 1.010 to 1.030. Patients with diabetes mellitus have an abnormally high specific gravity of their urine samples, while those with some forms of kidney disease have an abnormally low specific gravity.

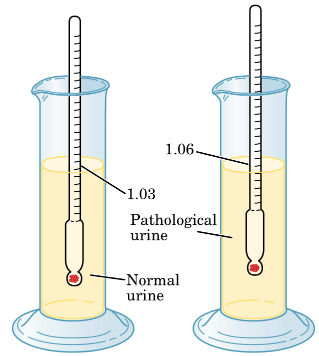

Figure 1.7 Urinometer.

EXAMPLE 1.8

The density of ethanol at 20°C is 0.789 g/mL. What is its specific gravity?

Solution

$$\text{Specific gravity} = \frac{0.789\ \text{g/mL}}{1.00\ \text{g/mL}} = 0.789$$

Problem 1.8

The specific gravity of a urine sample is 1.016. What is its density, in g/mL?

1.8 How Do We Describe the Various Forms of Energy?

Energy is defined as the capacity to do work. It can be described as being either kinetic energy or potential energy.

Kinetic energy (KE) is the energy of motion. Any object that is moving possesses kinetic energy. We can calculate how much energy by the formula $KE = 1/2mv^2$, where m is the mass of the object and v is its velocity. This means that kinetic energy increases (1) when an object moves faster and (2) when a heavier object is moving. When a truck and a bicycle are moving at the same velocity, the truck has more kinetic energy.

Potential energy is stored energy. The potential energy possessed by an object arises from its capacity to move or to cause motion. For example, body weight in the up position on a seesaw contains potential energy—it is capable of doing work. If given a slight push, it will move down. The potential energy of the body in the up position is converted to the kinetic energy of the body in the down position moving to the up position. Work is done against gravity in the process. Figure 1.8 shows another way in which potential energy is converted to kinetic energy.

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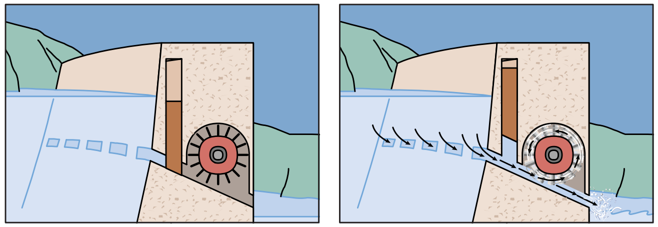
Potential energy is stored in this drawn bow and becomes kinetic energy in the arrow when released.

An important principle in nature is that things have a tendency to seek their lowest possible potential energy. We all know that water always flows downhill and not uphill.

Several forms of energy exist. The most important are (1) mechanical energy, light, heat, and electrical energy, which are examples of kinetic energy possessed by all moving objects, whether elephants or molecules or electrons; and (2) chemical energy and nuclear energy, which are examples of potential energy or stored energy. In chemistry the most important form



Potential energy ready to do work | Kinetic energy doing work

Figure 1.8 The water held back by the dam possesses potential energy, which is converted to kinetic energy when the water is released.

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An example of energy conversion. Light energy from the sun is converted to electrical energy by solar cells. The electricity runs a refrigerator on the back of the camel, keeping the vaccines cool so that they can be delivered to remote locations.

of potential energy is chemical energy—the energy stored within chemical substances and given off when they take part in a chemical reaction. For example, a log possesses chemical energy. When the log is ignited in a fireplace, the chemical energy (potential) of the wood is turned into energy in the form of heat and light. Specifically, the potential energy has been transformed into thermal energy (heat makes molecules move faster) and the radiant energy of light.

The various forms of energy can be converted from one to another. In fact, we make such conversions all the time. A power plant operates either on the chemical energy derived from burning fuel or on nuclear energy. This energy is converted to heat, which is converted to the electricity that is sent over transmission wires into houses and factories. Here we convert the electricity to light, heat (in an electrical heater, for example), or mechanical energy (in the motors of refrigerators, vacuum cleaners, and other devices).

Although one form of energy can be converted to another, the *total amount* of energy in any system does not change. *Energy can be neither created nor destroyed.* This statement is called the **law of conservation of energy.***

1.9 How Do We Describe Heat and the Ways in Which It Is Transferred?

A. Heat and Temperature

One form of energy that is particularly important in chemistry is **heat.** This is the form of energy that most frequently accompanies chemical reactions. Heat is not the same as temperature, however. Heat is a form of energy, but temperature is not.

The difference between heat and temperature can be seen in the following example. If we have two beakers, one containing 100 mL of water and the other containing 1 L of water at the same temperature, the heat content of the water in the larger beaker is ten times that of the water in the smaller beaker, even though the temperature is the same in both. If you were to dip your hand accidentally into a liter of boiling water, you would be much more severely burned than if only one drop fell on your hand. Even though the water is at the same temperature in both cases, the liter of boiling water has much more heat.

*This statement is not completely true. As discussed in Sections 3.8 and 3.9, it is possible to convert matter to energy, and vice versa. Therefore, a more correct statement would be *matter-energy can be neither created nor destroyed.* However, the law of conservation of energy is valid for most purposes and is highly useful.

Chemical Connections 1B

Hypothermia and Hyperthermia

The human body cannot tolerate temperatures that are too low. A person outside in very cold weather (say, −20°F [−29°C]) who is not protected by heavy clothing will eventually freeze to death because the body loses heat. Normal body temperature is 37°C. When the outside temperature is lower than that, heat flows out of the body. When the air temperature is moderate (10°C to 25°C), this poses no problem and is, in fact, necessary because the body produces more heat than it needs and must lose some. At extremely low temperatures, however, too much heat is lost and body temperature drops, a condition called **hypothermia.** A drop in body temperature of 1 or 2°C causes shivering, which is the body's attempt to increase its temperature by the heat generated through muscular action. An even greater drop results in unconsciousness and eventually death.

The opposite condition is **hyperthermia.** It can be caused either by high outside temperatures or by the body itself when an individual develops a high fever. A sustained body temperature as high as 41.7°C (107°F) is usually fatal.

As we saw in Section 1.4, temperature is measured in degrees. Heat can be measured in various units, the most common of which is the **calorie,** which is defined as the amount of heat necessary to raise the temperature of 1 g of liquid water by 1°C. This is a small unit, and chemists more often use the kilocalorie (kcal):

$$1 \text{ kcal} = 1000 \text{ cal}$$

Nutritionists use the word "Calorie" (with a capital "C") to mean the same thing as "kilocalorie"; that is, 1 Cal = 1000 cal = 1 kcal. The calorie is not part of the SI. The official SI unit for heat is the **joule (J),** which is about one-fourth as big as the calorie:

$$1 \text{ cal} = 4.184 \text{ J}$$

B. Specific Heat

As we noted, it takes 1 cal to raise the temperature of 1 g of liquid water by 1°C. **Specific heat (SH)** is the amount of heat necessary to raise the temperature of 1 g of any substance by 1°C. Each substance has its own specific heat, which is a physical property of that substance, like density or melting point. Table 1.4 lists specific heats for a few common substances. For example, the specific heat of iron is 0.11 cal/g · °C. Therefore, if we had 1 g of iron at 20°C, it would require only 0.11 cal to increase the temperature to 21°C. Under the same conditions, aluminum would require twice as much heat. Thus cooking in an aluminum pan of the same weight as an iron pan would require more heat than cooking in the iron pan. Note from Table 1.4 that ice and steam do not have the same specific heat as liquid water.

Table 1.4 Specific Heats for Some Common Substances

Substance	Specific Heat (cal/g · °C)	Substance	Specific Heat (cal/g · °C)
Water	1.00	Wood (typical)	0.42
Ice	0.48	Glass (typical)	0.22
Steam	0.48	Rock (typical)	0.20
Iron	0.11	Ethanol	0.59
Aluminum	0.22	Methanol	0.61
Copper	0.092	Ether	0.56
Lead	0.038	Carbon tetrachloride	0.21

Chemical Connections 1C

Cold Compresses, Waterbeds, and Lakes

The high specific heat of water is useful in cold compresses and makes them last a long time. For example, consider two patients with cold compresses: one compress made by soaking a towel in water and the other made by soaking a towel in ethanol. Both are at 0°C. Each gram of water in the water compress requires 25 cal to make the temperature of the compress rise to 25°C (after which it must be changed). Because the specific heat of ethanol is 0.59 cal/g · °C (see Table 1.4), each gram of ethanol requires only 15 cal to reach 25°C. If the two patients give off heat at the same rate, the ethanol compress is less effective because it will reach 25°C a good deal sooner than the water compress and will need to be changed sooner.

The high specific heat of water also means that it takes a great deal of heat to increase its temperature. That is why it takes a long time to get a pot of water to boil. Anyone who has a waterbed (300 gallons) knows that it takes days for the heater to bring the bed up to the desired temperature. It is particularly annoying when an overnight guest tries to adjust the temperature of your waterbed because the guest will probably have left before the change is noticed, but then you will have to set it back to your favorite temperature. This same effect in reverse explains why the outside temperature can be below zero (°C) for weeks before a lake will freeze. Large bodies of water do not change temperature very quickly.

It is easy to make calculations involving specific heats. The equation is

$$\text{Amount of heat} = \text{specific heat} \times \text{mass} \times \text{change in temperature}$$

$$\text{Amount of heat} = \text{SH} \times m \times \Delta T$$

where ΔT is the change in temperature.

We can also write this equation as

$$\text{Amount of heat} = \text{SH} \times m \times (T_2 - T_1)$$

where T_2 is the final temperature and T_1 is the initial temperature in °C.

EXAMPLE 1.9

How many calories are required to heat 352 g of water from 23°C to 95°C?

Solution

$$\text{Amount of heat} = \text{SH} \times m \times \Delta T$$

$$\begin{aligned}\text{Amount of heat} &= \text{SH} \times m \times (T_2 - T_1)\\ &= \frac{1.00\ \text{cal}}{\cancel{g} \cdot \cancel{°C}} \times 352\ \cancel{g} \times (95 - 23)\cancel{°C}\\ &= 2.5 \times 10^4\ \text{cal}\end{aligned}$$

Is this answer reasonable? Each gram of water requires one calorie to raise its temperature by one degree. We have approximately 350 g of water. To raise its temperature by one degree would therefore require approximately 350 calories. But we are raising the temperature not by one degree but by approximately 70 degrees (from 23 to 95). Thus the total number of calories will be approximately 70 × 350 = 24,500 cal, which is close to the calculated answer. (Even though we were asked for the answer in calories, we should note that it will be more convenient to convert to 25 kcal. We are going to see that conversion from time to time.)

Problem 1.9

How many calories are required to heat 731 g of water from 8°C to 74°C? Check your answer to see whether it is reasonable.

EXAMPLE 1.10

If we add 450 cal of heat to 37 g of ethanol at 20°C, what is the final temperature?

Solution

The specific heat of ethanol is 0.59 cal/g · °C (see Table 1.4).

$$\text{Amount of heat} = \text{SH} \times m \times \Delta T$$

$$\text{Amount of heat} = \text{SH} \times m \times (T_2 - T_1)$$

$$450 \text{ cal} = 0.59 \text{ cal/g} \cdot {}^\circ\text{C} \times 37 \text{ g} \times (T_2 - T_1)$$

We can show the units in fraction form by rewriting this equation.

$$450 \text{ cal} = 0.59 \frac{\text{cal}}{\text{g} \cdot {}^\circ\text{C}} \times 37 \text{ g} \times (T_2 - T_1)$$

$$(T_2 - T_1) = \frac{\text{amount of heat}}{\text{SH} \times m}$$

$$(T_2 - T_1) = \frac{450 \cancel{\text{cal}}}{\left[\dfrac{0.59 \cancel{\text{cal}} \times 37 \cancel{\text{g}}}{\cancel{\text{g}} \cdot {}^\circ\text{C}}\right]} = \frac{21}{1/{}^\circ\text{C}} = 21^\circ\text{C}$$

(Note that we have the reciprocal of temperature in the denominator, which gives us temperature in the numerator. The answer has units of degrees Celsius.) Since the starting temperature is 20°C, the final temperature is 41°C.

Is this answer reasonable? The specific heat of ethanol is 0.59 cal/g · °C. This value is close to 0.5, meaning that about half a calorie will raise the temperature of 1 g by 1°C. However, 37 g of ethanol needs approximately 40 times as many calories for a rise, and $40 \times \frac{1}{2} = 20$ calories. We are adding 450 calories, which is about 20 times as much. Thus we expect the temperature to rise by about 20°C, from 20°C to 40°C. The actual answer, 41°C, is quite reasonable.

Problem 1.10

A 100 g piece of iron at 25°C is heated by adding 230 cal. What will be the final temperature? Check your answer to see whether it is reasonable.

EXAMPLE 1.11

We heat 50.0 g of an unknown substance by adding 205 cal, and its temperature rises by 7.0°C. What is its specific heat? Using Table 1.4, identify the substance.

Solution

$$SH = \frac{\text{Amount of heat}}{m \times (\Delta T)}$$

$$SH = \frac{\text{Amount of heat}}{m \times (T_2 - T_1)}$$

$$SH = \frac{205 \text{ cal}}{50.0 \text{ g} \times 7.0°\text{C}} = 0.59 \text{ cal/g} \cdot °\text{C}$$

The substance in Table 1.4 having a specific heat of 0.59 cal/g · °C is ethanol.

Is this answer reasonable? If we had water instead of an unknown substance with SH = 1 cal/g · °C, raising the temperature of 50.0 g by 7.0°C would require 50 × 7.0 = 350 cal. But we added only approximately 200 cal. Therefore, the SH of the unknown substance must be less than 1.0. How much less? Approximately 200/350 = 0.6. The actual answer, 0.59 cal/g · °C, is quite reasonable.

Problem 1.11

It required 88.2 cal to heat 13.4 g of an unknown substance from 23°C to 176°C. What is the specific heat of the unknown substance? Check your answer to see whether it is reasonable.

SUMMARY OF KEY QUESTIONS

SECTION 1.1 Why Do We Call Chemistry the Study of Matter?

- **Chemistry** is the science that deals with the structure of matter and the changes it can undergo. In a **chemical change** or **chemical reaction,** substances are used up and others are formed.
- Chemistry is also the study of energy changes during chemical reactions. In **physical changes** substances do not change their identity.

SECTION 1.2 What Is the Scientific Method?

- The **scientific method** is a tool used in science and medicine. The heart of the scientific method is the testing of **hypotheses** and **theories** by collecting facts.

SECTION 1.3 How Do Scientists Report Numbers?

- Because we frequently use very large or very small numbers, we use powers of 10 to express these numbers more conveniently, a method called **exponential notation.**
- With exponential notation, we no longer have to keep track of so many zeros, and we have the added convenience of being able to see which digits convey information (**significant figures**) and which merely indicate the position of the decimal point.

SECTION 1.4 How Do We Make Measurements?

- In chemistry we use the **metric system** for measurements.
- The base units are the meter for length, the liter for volume, the gram for mass, the second for time, and the calorie for heat. Other units are indicated by prefixes that represent powers of 10. Temperature is measured in degrees Celsius or in kelvins.

SECTION 1.5 What Is a Handy Way to Convert from One Unit to Another?

- Conversions from one unit to another are best done by the **factor-label method,** in which units are multiplied and divided.

SECTION 1.6 What Are the States of Matter?

- There are three states of matter: **solid, liquid,** and **gas.**

SECTION 1.7 What Are Density and Specific Gravity?

- **Density** is mass per unit volume. **Specific gravity** is density relative to water and thus has no units. Density usually decreases with increasing temperature.

SECTION 1.8 How Do We Describe the Various Forms of Energy?

- **Kinetic energy** is energy of motion; **potential energy** is stored energy. Energy can be neither created nor destroyed, but it can be converted from one form to another.

SECTION 1.9 How Do We Describe Heat and the Ways in Which It Is Transferred?

- **Heat** is a form of energy and is measured in calories. A calorie is the amount of heat necessary to raise the temperature of 1 g of liquid water by 1°C.
- Every substance has a **specific heat,** which is a physical constant. The specific heat is the number of calories required to raise the temperature of 1 g of a substance by 1°C.

PROBLEMS

GOB Chemistry Now™

Assess your understanding of this chapter's topics with additional quizzing and conceptual-based problems at **http://now.brookscole.com/gob8** or on the CD.

A blue problem number indicates an applied problem.

■ denotes problems that are available on the GOB ChemistryNow website or CD and are assignable in OWL.

SECTION 1.1 Why Do We Call Chemistry the Study of Matter?

1.12 The life expectancy of a citizen in the United States is 76 years. Eighty years ago it was 56 years. In your opinion, what was the major contributor to this spectacular increase in life expectancy? Explain your answer.

1.13 Define the following terms:
(a) Matter (b) Chemistry

SECTION 1.2 What Is the Scientific Method?

1.14 ■ In Table 1.4 you find four metals (iron, aluminum, copper, and lead) and three organic compounds (ethanol, methanol, and ether).What kind of hypothesis would you suggest about the specific heats of these chemicals?

1.15 ■ In a newspaper, you read that Dr. X claimed that he has found a new remedy to cure diabetes. The remedy is an extract of carrots. How would you classify this claim: (a) fact, (b) theory, (c) hypothesis, or (d) hoax? Explain your choice of answer.

1.16 ■ Classify each of the following as a chemical or physical change:
(a) Burning gasoline
(b) Making ice cubes
(c) Boiling oil
(d) Melting lead
(e) Rusting iron
(f) Making ammonia from nitrogen and hydrogen
(g) Digesting food

SECTION 1.3 How Do Scientists Report Numbers?

Exponential Notation

1.17 Write in exponential notation:
(a) 0.351 (b) 602.1 (c) 0.000128 (d) 628122

1.18 Write out in full:
(a) 4.03×10^5 (b) 3.2×10^3
(c) 7.13×10^{-5} (d) 5.55×10^{-10}

1.19 Multiply:
(a) $(2.16 \times 10^5)(3.08 \times 10^{12})$
(b) $(1.6 \times 10^{-8})(7.2 \times 10^8)$
(c) $(5.87 \times 10^{10})(6.62 \times 10^{-27})$
(d) $(5.2 \times 10^{-9})(6.8 \times 10^{-15})$

1.20 Divide:
(a) $\dfrac{6.02 \times 10^{23}}{2.87 \times 10^{10}}$ (b) $\dfrac{3.14}{2.93 \times 10^{-4}}$
(c) $\dfrac{5.86 \times 10^{-9}}{2.00 \times 10^{3}}$ (d) $\dfrac{7.8 \times 10^{-12}}{9.3 \times 10^{-14}}$
(e) $\dfrac{6.83 \times 10^{-12}}{5.02 \times 10^{14}}$

1.21 Add:
(a) $(7.9 \times 10^4) + (5.2 \times 10^4)$
(b) $(8.73 \times 10^4) + (6.7 \times 10^3)$
(c) $(3.63 \times 10^{-4}) + (4.776 \times 10^{-3})$

1.22 Subtract:
(a) $(8.50 \times 10^3) - (7.61 \times 10^2)$
(b) $(9.120 \times 10^{-2}) - 3.12 \times 10^{-3})$
(c) $(1.3045 \times 10^2) - (2.3 \times 10^{-1})$

1.23 ■ Solve:
$$\frac{(3.14 \times 10^3) \times (7.80 \times 10^5)}{(5.50 \times 10^2)}$$

1.24 ■ Solve:
$$\frac{(9.52 \times 10^4) \times (2.77 \times 10^{-5})}{(1.39 \times 10^7) \times (5.83 \times 10^2)}$$

Significant Figures

1.25 How many significant figures are in the following:
(a) 0.012 (b) 0.10203
(c) 36.042 (d) 8401.0
(e) 32100 (f) 0.0402
(g) 0.000012

1.26 How many significant figures are in the following:
(a) 5.71×10^{13} (b) 4.4×10^{5}
(c) 3×10^{-6} (d) 4.000×10^{-11}
(e) 5.5550×10^{-3}

1.27 Round off to two significant figures:
(a) 91.621 (b) 7.329
(c) 0.677 (d) 0.003249
(e) 5.88

1.28 ■ Multiply these numbers, using the correct number of significant figures in your answer:
(a) 3630.15×6.8
(b) 512×0.0081
(c) $5.79 \times 1.85825 \times 1.4381$

1.29 Divide these numbers, using the correct number of significant figures in your answer:
(a) $\frac{3.185}{2.08}$ (b) $\frac{6.5}{3.0012}$ (c) $\frac{0.0035}{7.348}$

1.30 Add these groups of measured numbers using the correct number of significant figures in your answer:

(a)
$$\begin{array}{r} 37.4083 \\ 5.404 \\ 10916.3 \\ 3.94 \\ 0.0006 \\ \hline \end{array}$$

(b)
$$\begin{array}{r} 84 \\ 8.215 \\ 0.01 \\ 151.7 \\ \hline \end{array}$$

(c)
$$\begin{array}{r} 51.51 \\ 100.27 \\ 16.878 \\ 3.6817 \\ \hline \end{array}$$

SECTION 1.4 How Do We Make Measurements?

1.31 In the SI system, the second is the base unit of time. We talk about atomic events that occur in picoseconds (10^{-12} s) or even in femtoseconds (10^{-15} s). But we don't talk about megaseconds or kiloseconds; the old standards of minutes, hours, and days prevail. How many minutes and hours are 20 kiloseconds?

1.32 ■ How many grams are in the following:
(a) 1 kg (b) 1 mg

1.33 Estimate without actually calculating which one is the shorter distance:
(a) 20 mm or 0.3 m
(b) 1 inch or 30 mm
(c) 2000 m or 1 mile

1.34 For each of these, tell which answer is closest:
(a) A baseball bat has a length of 100 mm or 100 cm or 100 m
(b) A glass of milk holds 23 cc or 230 mL or 23 L
(c) A man weighs 75 mg or 75 g or 75 kg
(d) A tablespoon contains 15 mL or 150 mL or 1.5 L
(e) A paper clip weighs 50 mg or 50 g or 50 kg
(f) Your hand has a width of 100 mm or 100 cm or 100 m
(g) An audiocassette weighs 40 mg or 40 g or 40 kg

1.35 You are taken for a helicopter ride in Hawaii from Kona (sea level) to the top of the volcano Mauna Kea. Which property of your body would change during the helicopter ride:
(a) height (b) weight (c) volume (d) mass

1.36 Convert to Celsius and to Kelvin:
(a) 320°F (b) 212°F (c) 0°F (d) −250°F

1.37 Convert to Fahrenheit and to Kelvin:
(a) 25°C (b) 40°C (c) 250°C (d) −273°C

SECTION 1.5 What Is a Handy Way to Convert from One Unit to Another?

1.38 Make the following conversions (conversion factors are given in Table 1.3):
(a) 42.6 kg to lb (b) 1.62 lb to g
(c) 34 in. to cm (d) 37.2 km to mi
(e) 2.73 gal to L (f) 62 g to oz
(g) 33.61 qt to L (h) 43.7 L to gal
(i) 1.1 mi to km (j) 34.9 mL to fl oz

1.39 Make the following metric conversions:
(a) 96.4 mL to L (b) 275 mm to cm
(c) 45.7 kg to g (d) 475 cm to m
(e) 21.64 cc to mL (f) 3.29 L to cc
(g) 0.044 L to mL (h) 711 g to kg
(i) 63.7 mL to cc (j) 0.073 kg to mg
(k) 83.4 m to mm (l) 361 mg to g

1.40 You drive in Canada where the distances are marked in kilometers. The sign says you are 80 km from Ottawa. You are traveling at a speed of 75 mi/h. Would you reach Ottawa within one hour, after one hour, or later than that?

1.41 The speed limit in some European cities is 80 km/h. How many miles per hour is this?

1.42 Your car gets 25.00 miles on a gallon of gas. What would be your car's fuel efficiency in km/L?

SECTION 1.6 What Are the States of Matter?

1.43 Which states of matter have a definite volume?

1.44 Will most substances be solids, liquids, or gases at low temperatures?

1.45 ■ Does the chemical nature of a substance change when it melts from a solid to a liquid?

SECTION 1.7 What Are Density and Specific Gravity?

1.46 The volume of a rock weighing 1.075 kg is 334.5 mL. What is the density of the rock in g/mL? Express it to three significant figures.

1.47 The density of manganese is 7.21 g/mL, that of calcium chloride is 2.15 g/mL, and that of sodium acetate is 1.528 g/mL. You place these three solids in a liquid, in which they are not soluble. The liquid has a density of 2.15 g/mL. Which will sink to the bottom, which will stay on the top, and which will stay in the middle of the liquid?

1.48 The density of titanium is 4.54 g/mL. What is the volume, in milliliters, of 163 g of titanium?

1.49 A 335.0 cc sample of urine has a mass of 342.6 g. What is the density, in g/mL, to three decimal places?

1.50 The density of methanol at 20°C is 0.791 g/mL. What is the mass, in grams, of a 280 mL sample?

1.51 The density of dichloromethane, a liquid insoluble in water, is 1.33 g/cc. If dichloromethane and water are placed in a separatory funnel, which will be the upper layer?

1.52 A sample of 10.00 g of oxygen has a volume of 6702 mL. The same weight of carbon dioxide occupies 5058 mL.

(a) What is the density of each gas in g/L?

(b) Carbon dioxide is used as a fire extinguisher to cut off the fire's supply of oxygen. Do the densities of these two gases explain the fire-extinguishing ability of carbon dioxide?

1.53 Crystals of a material are suspended in the middle of a cup of water at 2°C. This means that the densities of the crystal and of the water are the same. How might you enable the crystals to rise to the surface of the water so that you can harvest them?

SECTION 1.8 How Do We Describe the Various Forms of Energy?

1.54 On many country roads you see telephones powered by a solar panel. What principle is at work in these devices?

1.55 While you drive your car, your battery is charged. How would you describe this process in terms of kinetic and potential energy?

SECTION 1.9 How Do We Describe Heat and the Ways in Which It Is Transferred?

1.56 ■ How many calories are required to heat the following (specific heats are given in Table 1.4)?

(a) 52.7 g of aluminum from 100°C to 285°C

(b) 93.6 g of methanol from −35°C to 55°C

(c) 3.4 kg of lead from −33°C to 730°C

(d) 71.4 g of ice from −77°C to −5°C

1.57 If 168 g of an unknown liquid requires 2750 cal of heat to raise its temperature from 26°C to 74°C, what is the specific heat of the liquid?

1.58 The specific heat of steam is 0.48 cal/g · °C. How many kilocalories are needed to raise the temperature of 10.5 kg steam from 120°C to 150°C?

Chemical Connections

1.59 (Chemical Connections 1A) If the recommended dose of a drug is 445 mg for a 180 lb man, what would be a suitable dose for a 135 lb man?

1.60 (Chemical Connections 1A) The average lethal dose of heroin is 1.52 mg/kg of body weight. Estimate how many grams of heroin would be lethal for a 200 lb man.

1.61 (Chemical Connections 1B) How does the body react to hypothermia?

1.62 (Chemical Connections 1B) Low temperatures often cause people to shiver. What is the function of this involuntary body action?

1.63 (Chemical Connections 1C) Which would make a more efficient cold compress, ethanol or methanol? (Refer to Table 1.4.)

Additional Problems

1.64 The meter is a measure of length. Tell what each of the following units measures:

(a) cm^3	(b) mL	(c) kg	(d) cal
(e) g/cc	(f) joule	(g) °C	(h) cm/s

1.65 A brain weighing 1.0 lb occupies a volume of 620 mL. What is the specific gravity of the brain?

1.66 ■ If the density of air is 1.25×10^{-3} g/cc, what is the mass in kilograms of the air in a room that is 5.3 m long, 4.2 m wide, and 2.0 m high?

1.67 Classify these as kinetic or potential energy:

(a) Water held by a dam

(b) A speeding train

(c) A book on its edge before falling

(d) A falling book

(e) Electric current in a light bulb

1.68 The kinetic energy possessed by an object with a mass of 1 g moving with a velocity of 1 cm/s is called 1 erg. What is the kinetic energy, in ergs, of an athlete with a mass of 127 lb running at a velocity of 14.7 mi/h?

1.69 A European car advertises an efficiency of 22 km/L, while an American car claims an economy of 30 mi/gal. Which car is more efficient?

1.70 In Potsdam, New York, you can buy gas for US$2.02/gal. In Montreal, Canada, you pay US$0.60/L. (Currency conversions are outside the scope of this text, so you are not asked to do them here.) Which is the better buy? Is your calculation reasonable?

1.71 Shivering is the body's response to increase the body temperature. What kind of energy is generated by shivering?

1.72 When the astronauts walked on the Moon, they could make giant leaps in spite of their heavy gear.

(a) Why were their weights on the Moon so small?

(b) Were their masses different on the Moon than on the Earth?

1.73 Which of the following is the largest mass and which is the smallest?

(a) 41 g (b) 3×10^3 mg

(c) 8.2×10^6 μg (d) 4.1310×10^{-8} kg

1.74 Which quantity is bigger in each of the following pairs:

(a) 1 gigaton : 10 megaton

(b) 10 micrometer : 1 millimeter

(c) 10 centigram : 200 milligram

1.75 In Japan, high-speed "bullet trains" move with an average speed of 220 km/h. If Dallas and Los Angeles were connected by such a train, how long would it take to travel nonstop between these cities (a distance of 1490 miles)?

1.76 The specific heats of some elements at 25°C are as follows: aluminum = 0.215 cal/g · °C; carbon (graphite) = 0.170 cal/g · °C; iron = 0.107 cal/g · °C; mercury = 0.0331 cal/g · °C.

(a) Which element would require the smallest amount of heat to raise the temperature of 100 g of the element by 10°C?

(b) If the same amount of heat needed to raise the temperature of 1 g of aluminum by 25°C were applied to 1 g of mercury, by how many degrees would its temperature be raised?

(c) If a certain amount of heat is used to raise the temperature of 1.6 g of iron by 10°C, the temperature of 1 g of which element would also be raised by 10°C, using the same amount of heat?

1.77 Water that contains deuterium rather than ordinary hydrogen (see Chapter 3) is called heavy water. The specific heat of heavy water at 25°C is 4.217 J/g · °C. Which requires more energy to raise the temperature of 10.0 g by 10°C, water or heavy water?

1.78 One quart of milk costs 80 cents and one liter costs 86 cents. Which is the better buy?

1.79 Consider butter, density 0.860 g/mL, and sand, density 2.28 g/mL.

(a) If 1.00 mL of butter is thoroughly mixed with 1.00 mL of sand, what is the density of the mixture?

(b) What would be the density of the mixture if 1.00 g of the same butter were mixed with 1.00 g of the same sand?

1.80 Which speed is the fastest?

(a) 70 mi/h (b) 140 km/h

(c) 4.5 km/s (d) 48 mi/min

1.81 In calculating the specific heat of a substance, the following data are used: mass = 92.15 g; heat = 3.200 kcal; rise in temperature = 45°C. How many significant figures should you report in calculating the specific heat?

1.82 A solar cell generates 500 kilojoules of energy per hour. To keep a refrigerator at 4°C, one needs 250 kcal/h. Can the solar cell supply sufficient energy per hour to maintain the temperature of the refrigerator?

1.83 The specific heat of urea is 1.339 J/g · °C . If one adds 60.0 J of heat to 10.0 g of urea at 20°C, what would be the final temperature?

Special Categories

Three special categories of problems—Tying It Together, Looking Ahead, and Challenge Problems—will appear from time to time at the ends of chapters. Not every chapter will have these problems, but they will appear to make specific points.

Tying It Together

1.84 Heats of reaction are frequently measured by monitoring the change in temperature of a water bath in which the reaction mixture is immersed. A water bath used for this purpose contains 2.000 L of water. In the course of the reaction, the temperature of the water rose 4.85°C. How many calories were liberated by the reaction? (You will need to use what you know about unit conversions and apply that information to what you know about energy and heat.)

1.85 You have samples of urea (a solid at room temperature) and pure ethyl alcohol (a liquid at room temperature). Which technique or techniques would you use to measure the amount of each substance?

Looking Ahead

1.86 You have a sample of material used in folk medicine. Suggest the approach you would use to determine whether this material contains an effective substance for treating disease. If you do find a new and effective substance, can you think of a way to determine the amount present in your sample? (Pharmaceutical companies have used this approach to produce many common medications.)

1.87 Many substances that are involved in chemical reactions in the human body (and in all organisms) contain carbon, hydrogen, oxygen, and nitrogen arranged in specific patterns. Would you expect new medications to have features in common with these substances, or would you expect them to be drastically different? What are the reasons for your answer?

Challenge Problems

1.88 If 2 kg of a given reactant is consumed in the reaction described in Problem 1.84, how many calories are liberated for each kilogram?

1.89 You have a water sample that contains a contaminant you want to remove. You know that the contaminant is much more soluble in diethyl ether than it is in water. You have a separatory funnel available. Propose a way to remove the contaminant.