1

CHAPTER OUTLINE

CHAPTER GOALS

In this chapter you will learn how to:

1. Describe the three states of matter
2. Classify matter as a pure substance, mixture, element, or compound
3. Report measurements using the metric units of length, mass, and volume
4. Use significant figures
5. Use scientific notation for very large and very small numbers
6. Use conversion factors to convert one unit to another
7. Convert temperature from one scale to another
8. Define density and specific gravity and use density to calculate the mass or volume of a substance

Determining the weight and length of a newborn are common measurements performed by healthcare professionals.

MATTER AND MEASUREMENT

EVERYTHING you touch, feel, or taste is composed of chemicals—that is, **matter**—so an understanding of its composition and properties is crucial to our appreciation of the world around us. Some matter—lakes, trees, sand, and soil—is naturally occurring, while other examples of matter—aspirin, CDs, nylon fabric, plastic syringes, and vaccines—are made by humans. To understand the properties of matter, as well as how one form of matter is converted to another, we must also learn about measurements. Following a recipe, pumping gasoline, and figuring out drug dosages involve manipulating numbers. Thus, Chapter 1 begins our study of chemistry by examining the key concepts of matter and measurement.

1.1 CHEMISTRY—THE SCIENCE OF EVERYDAY EXPERIENCE

What activities might occupy the day of a typical student? You may have done some or all of the following tasks: eaten some meals, drunk coffee or cola, taken a shower with soap, gone to the library to research a paper, taken notes in a class, checked email on a computer, watched some television, ridden a bike or car to a part-time job, taken an aspirin to relieve a headache, and spent some of the evening having snacks and refreshments with friends. Perhaps, without your awareness, your life was touched by chemistry in each of these activities. What, then, is this discipline we call **chemistry?**

- *Chemistry* is the study of matter—its composition, properties, and transformations.

What is **matter?**

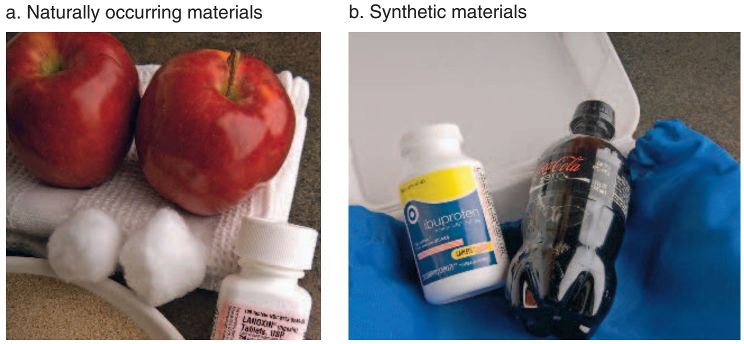
- *Matter* is anything that has mass and takes up volume.

In other words, **chemistry studies anything that we touch, feel, see, smell, or taste,** from simple substances like water or salt, to complex substances like proteins and carbohydrates that combine to form the human body. Some matter—cotton, sand, an apple, and the cardiac drug digoxin—is **naturally occurring,** meaning it is isolated from natural sources. Other substances—nylon, Styrofoam, the plastic used in soft drink bottles, and the pain reliever ibuprofen—are **synthetic,** meaning they are produced by chemists in the laboratory (Figure 1.1).

Sometimes a chemist studies what a substance is made of, while at other times he or she might be interested in its properties. Alternatively, the focus may be how to convert one substance into another (Figure 1.2). While naturally occurring rubber exists as the sticky white liquid **latex,** the laboratory process of vulcanization converts it to the stronger, more elastic material used in tires and other products. Although the anticancer drug **taxol** was first isolated in small quantities from the bark of the Pacific yew tree, stripping the bark killed these rare and magnificent trees. Taxol, sold under the trade name of Paclitaxel, is now synthesized in the lab from a substance in the pine needles of the common English yew tree, making it readily available for many cancer patients.

FIGURE 1.1 Naturally Occurring and Synthetic Materials

a. Naturally occurring materials b. Synthetic materials



Matter occurs in nature or is synthesized in the lab. (a) Sand and apples are two examples of natural materials. Cotton fabric is woven from cotton fiber, obtained from the cotton plant. The drug digoxin, widely prescribed for decades for patients with congestive heart failure, is extracted from the leaves of the woolly foxglove plant. (b) Nylon was the first synthetic fiber made in the laboratory. It quickly replaced the natural fiber silk in parachutes and ladies' stockings. Styrofoam and PET, the plastic used for soft drink bottles, are strong yet lightweight synthetic materials used for food storage. Over-the-counter pain relievers like ibuprofen are synthetic. The starting materials for all of these useful products are obtained from petroleum.

FIGURE 1.2 Transforming Natural Materials into Useful Synthetic Products

(a) Latex, the sticky liquid that oozes from a rubber tree when it is cut, is too soft for most applications. (b) Vulcanization converts latex to the stronger, elastic rubber used in tires and other products. (c) Taxol was first isolated by stripping the bark of the Pacific yew tree, a process that killed these ancient trees. Estimates suggest that sacrificing one 100-year-old tree provided enough taxol for only a single dose for one cancer patient. (d) Taxol, which is active against breast, ovarian, and some lung tumors, is now synthesized in the lab from a substance that occurs in the needles of the common English yew tree.

Chemistry is truly the science of everyday experience. Soaps and detergents, newspapers and CDs, lightweight exercise gear and Gore-Tex outer wear, condoms and oral contraceptives, Tylenol and penicillin—all of these items are products of chemistry. Without a doubt, advances in chemistry have transformed life in modern times.

PROBLEM 1.1

Look around you and identify five objects. Decide if they are composed of natural or synthetic materials.

PROBLEM 1.2

Imagine that your job as a healthcare professional is to take a blood sample from a patient and store it in a small container in a refrigerator until it is picked up for analysis in the hospital lab. You might have to put on gloves and a mask, use a plastic syringe with a metal needle, store the sample in a test tube or vial, and place it in a cold refrigerator. Pick five objects you might encounter during the process and decide if they are made of naturally occurring or synthetic materials.

1.2 STATES OF MATTER

Matter exists in three common states—solid, liquid, and gas.

- A *solid* has a definite volume, and maintains its shape regardless of the container in which it is placed. The particles of a solid lie close together, and are arranged in a regular three-dimensional array.
- A *liquid* has a definite volume, but takes on the shape of the container it occupies. The particles of a liquid are close together, but they can randomly move around, sliding past one another.
- A *gas* has no definite shape or volume. The particles of a gas move randomly and are separated by a distance much larger than their size. The particles of a gas expand to fill the volume and assume the shape of whatever container they are put in.

For example, water exists in its solid state as ice or snow, liquid state as liquid water, and gaseous state as steam or water vapor. Blow-up circles like those in Figure 1.3 will be used commonly in this text to indicate the composition and state of the particles that compose a substance. In this molecular art, different types of particles are shown in color-coded spheres, and the distance between the spheres signals its state—solid, liquid, or gas.

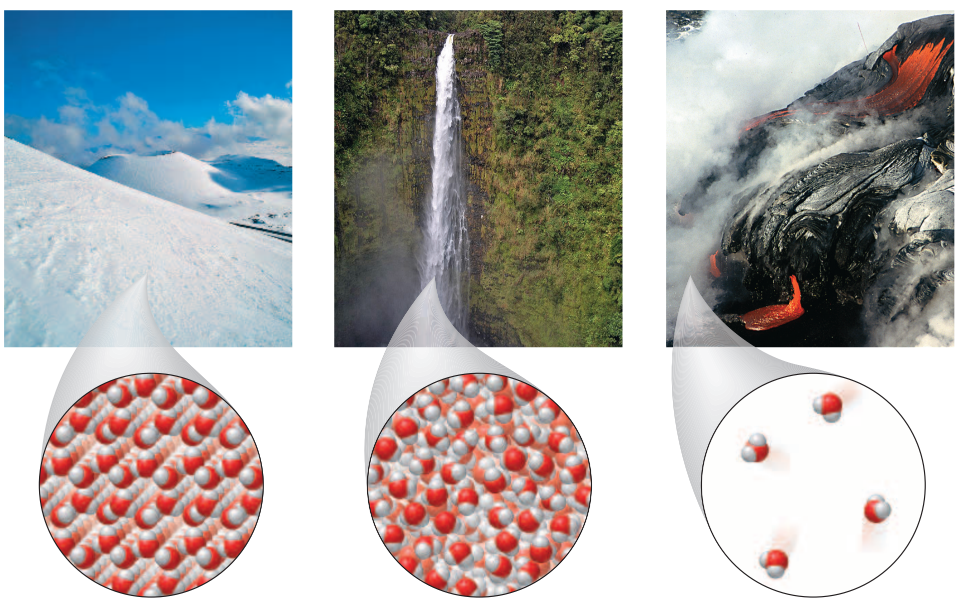
Matter is characterized by its **physical properties** and **chemical properties.**

- ***Physical properties*** **are those that can be observed or measured without changing the composition of the material.**

Common physical properties include melting point (mp), boiling point (bp), solubility, color, and odor. **A *physical change* alters a substance without changing its composition.** The most common physical changes are **changes in state.** Melting an ice cube to form liquid water, and boiling liquid water to form steam are two examples of physical changes. Water is the substance at the beginning and end of both physical changes. More details about physical changes are discussed in Chapter 7.

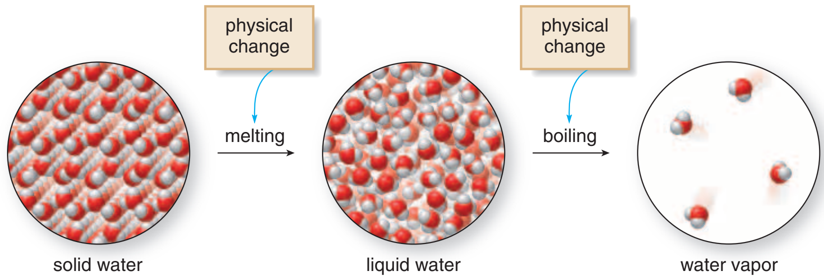
FIGURE 1.3 The Three States of Water—Solid, Liquid, and Gas

a. Solid water | b. Liquid water | c. Gaseous water



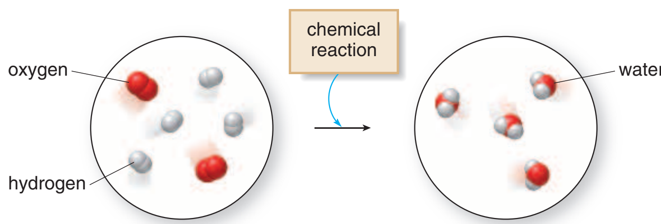
- a. The particles of a solid are close together and highly organized. (Photo: snow-capped Mauna Kea on the Big Island of Hawaii)
- b. The particles of a liquid are close together but more disorganized than the solid. (Photo: Akaka Falls on the Big Island of Hawaii)
- c. The particles of a gas are far apart and disorganized. (Photo: steam formed by a lava flow on the Big Island of Hawaii)

Each red sphere joined to two gray spheres represents a single water particle. In proceeding from left to right, from solid to liquid to gas, the molecular art shows that the level of organization of the water particles decreases. Color-coding and the identity of the spheres within the particles will be addressed in Chapter 2.



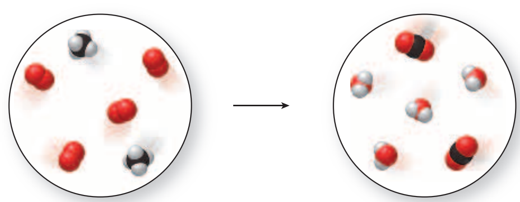
- *Chemical properties* are those that determine how a substance can be converted to another substance.

A *chemical change*, or a *chemical reaction*, converts one material to another. The conversion of hydrogen and oxygen to water is a chemical reaction because the composition of the material is different at the beginning and end of the process. Chemical reactions are discussed in Chapters 5 and 6.



PROBLEM 1.3 Characterize each process as a physical change or a chemical change: (a) making ice cubes; (b) burning natural gas; (c) silver jewelry tarnishing; (d) a pile of snow melting; (e) baking bread.

PROBLEM 1.4 Does the molecular art represent a chemical change or a physical change? Explain your choice.



1.3 CLASSIFICATION OF MATTER

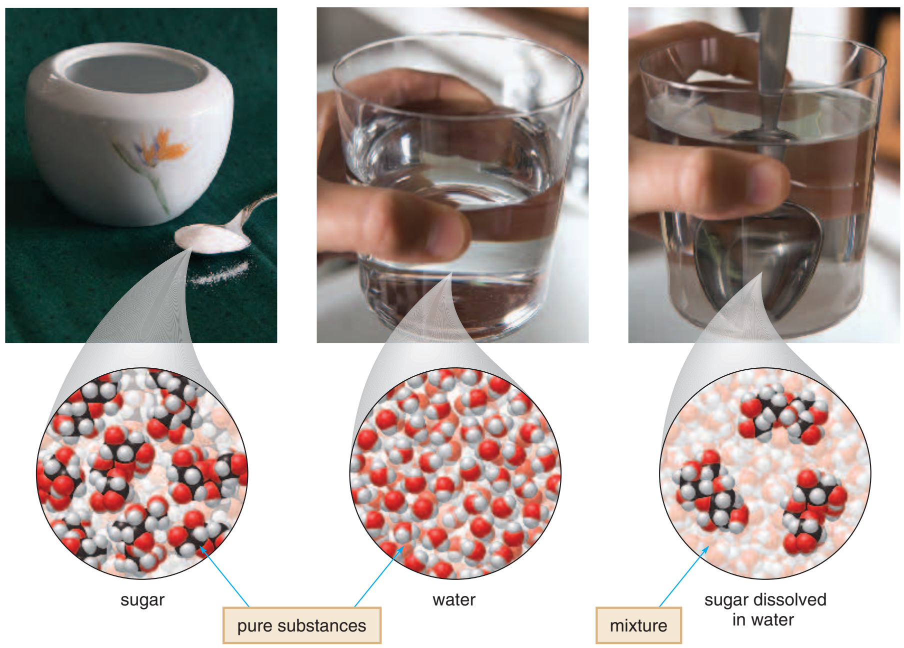
All matter can be classified as either a **pure substance** or a **mixture.**

- A *pure substance* is composed of a single component and has a constant composition, regardless of the sample size and the origin of the sample.

A pure substance, such as water or table sugar, can be characterized by its physical properties, because these properties do not change from sample to sample. **A pure substance cannot be broken down to other pure substances by any physical change.**

- A *mixture* is composed of more than one component. The composition of a mixture can vary depending on the sample.

The physical properties of a mixture may also vary from one sample to another. **A mixture can be separated into its components by physical changes.** Dissolving table sugar in water forms a mixture, whose sweetness depends on the amount of sugar added. If the water is allowed to evaporate from the mixture, pure table sugar and pure water are obtained.



Mixtures can be formed from solids, liquids, and gases, as shown in Figure 1.4. The compressed air breathed by a scuba diver consists mainly of the gases oxygen and nitrogen. A saline solution used in an IV bag contains solid sodium chloride (table salt) dissolved in water. Rubbing alcohol is a mixture composed of two liquids, 2-propanol and water.

A pure substance is classified as either an **element** or a **compound.**

- An *element* is a pure substance that cannot be broken down into simpler substances by a chemical reaction.
- A *compound* is a pure substance formed by chemically combining (joining together) two or more elements.

An alphabetical list of elements is located on the inside front cover of this text. The elements are commonly organized into a periodic table, also shown on the inside front cover, and discussed in much greater detail in Section 2.4.

Nitrogen gas, aluminum foil, and copper wire are all elements. Water is a compound because it is composed of the elements hydrogen and oxygen. Table salt, sodium chloride, is also a compound since it is formed from the elements sodium and chlorine (Figure 1.5). Although only 114 elements are currently known, over 20 million compounds occur naturally or have been synthesized in the laboratory. We will learn much more about elements and compounds in Chapter 2.

Figure 1.6 summarizes the categories into which matter is classified.

PROBLEM 1.5

Classify each item as a pure substance or a mixture: (a) blood; (b) ocean water; (c) a piece of wood; (d) a chunk of ice.

FIGURE 1.4 Three Examples of Mixtures

a. Two gases

b. A solid and a liquid

c. Two liquids

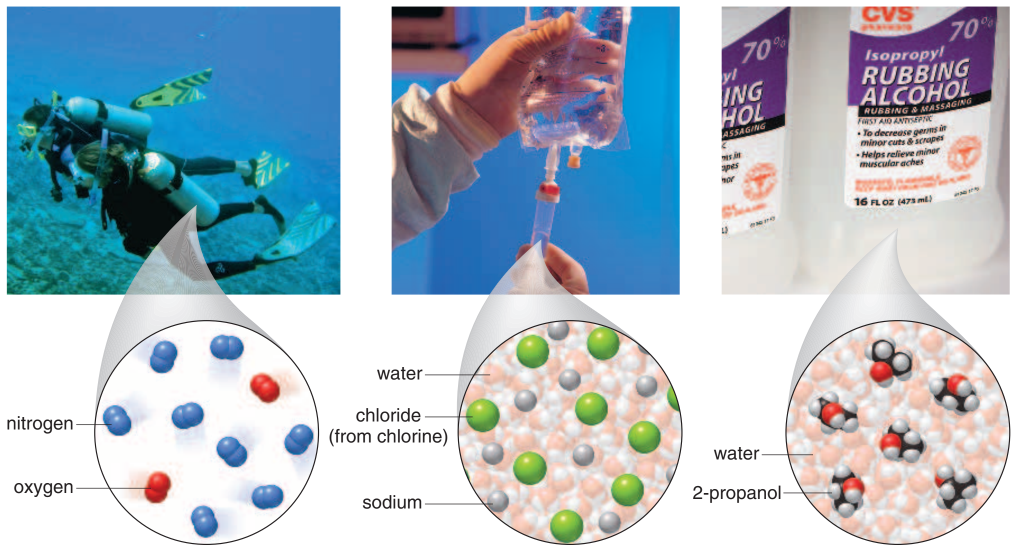


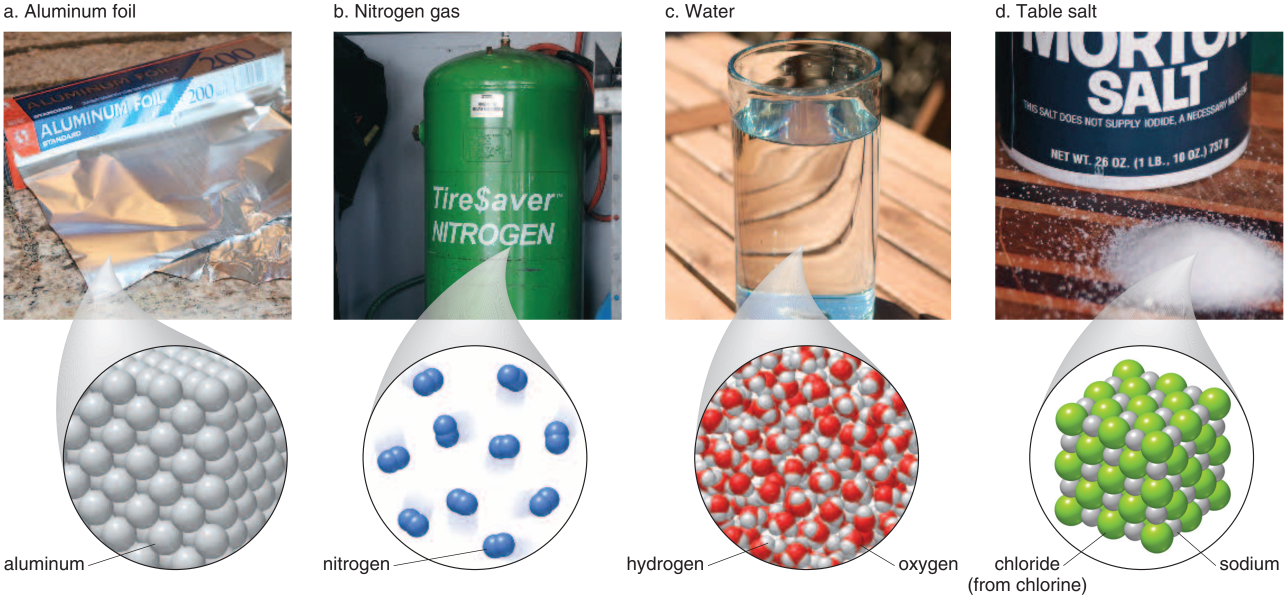
FIGURE 1.5 Elements and Compounds

a. Aluminum foil

b. Nitrogen gas

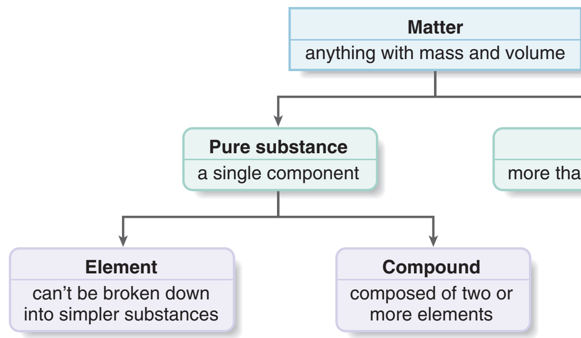
c. Water

d. Table salt



- Aluminum foil and nitrogen gas are elements. Water and table salt are compounds. Color-coding of the spheres used in the molecular art indicates that water is composed of two elements—hydrogen shown as gray spheres, and oxygen shown in red. Likewise, the gray (sodium) and green (chlorine) spheres illustrate that sodium chloride is formed from two elements as well.

FIGURE 1.6 Classification of Matter



PROBLEM 1.6 Classify each item as an element or a compound: (a) the gas inside a helium balloon; (b) table sugar; (c) the rust on an iron nail; (d) aspirin. All elements are listed alphabetically on the inside front cover.

1.4 MEASUREMENT

Any time you check your weight on a scale, measure the ingredients of a recipe, or figure out how far it is from one location to another, you are measuring a quantity. Measurements are routine for healthcare professionals who use weight, blood pressure, pulse, and temperature to chart a patient's progress.

- **Every measurement is composed of a *number* and a *unit.***

Reporting the value of a measurement is meaningless without its unit. For example, if you were told to give a patient an aspirin dosage of 325, does this mean 325 ounces, pounds, grams, milligrams, or tablets? Clearly there is a huge difference among these quantities.

1.4A THE METRIC SYSTEM

In the United States, most measurements are made with the **English system,** using units like miles (mi), gallons (gal), pounds (lb), and so forth. A disadvantage of this system is that the units are not systematically related to each other and require memorization. For example, 1 lb = 16 oz, 1 gal = 4 qt, and 1 mi = 5,280 ft.

In 1960, the **International System of Units** was formally adopted as the uniform system of units for the sciences. **SI units,** as they are called, are based on the metric system, but the system encourages the use of some metric units over others. SI stands for the French words, *Système Internationale.*

Scientists, health professionals, and people in most other countries use the **metric system,** with units like meter (m) for length, gram (g) for mass, and liter (L) for volume. The metric system is slowly gaining popularity in the United States. Although milk is still sold in quart or gallon containers, soft drinks are now sold in one- or two-liter bottles. The weight of packaged foods is often given in both ounces and grams. Distances on many road signs are shown in miles and kilometers. Most measurements in this text will be reported using the metric system, but learning to convert English units to metric units is also a necessary skill that will be illustrated in Section 1.7.

The important features of the metric system are the following:

- **Each type of measurement has a base unit—the meter (m) for length; the gram (g) for mass; the liter (L) for volume; the second (s) for time.**
- **All other units are related to the base unit by powers of 10.**
- **The prefix of the unit name indicates if the unit is larger or smaller than the base unit.**

The base units of the metric system are summarized in Table 1.1, and the most common prefixes used to convert the base units to smaller or larger units are summarized in Table 1.2. **The same prefixes are used for all types of measurement.** For example, the prefix *kilo-* means 1,000 times as large. Thus,

1 **kilo**meter = **1,000** meters or 1 km = 1,000 m
1 **kilo**gram = **1,000** grams or 1 kg = 1,000 g
1 **kilo**liter = **1,000** liters or 1 kL = 1,000 L

The prefix *milli-* means one thousandth as large (1/1,000 or 0.001). Thus,

1 **milli**meter = **0.001** meters or 1 mm = 0.001 m
1 **milli**gram = **0.001** grams or 1 mg = 0.001 g
1 **milli**liter = **0.001** liters or 1 mL = 0.001 L

TABLE 1.1 The Basic Metric Units

Quantity	Metric Base Unit	Symbol
Length	Meter	m
Mass	Gram	g
Volume	Liter	L
Time	Second	s

TABLE 1.2 Common Prefixes Used for Metric Units

Prefix	Symbol	Meaning	Numerical Value[a]	Scientific Notation[b]
Mega-	M	Million	1,000,000.	10^6
Kilo-	k	Thousand	1,000.	10^3
Deci-	d	Tenth	0.1	10^{-1}
Centi-	c	Hundredth	0.01	10^{-2}
Milli-	m	Thousandth	0.001	10^{-3}
Micro-	μ	Millionth	0.000 001	10^{-6}
Nano-	n	Billionth	0.000 000 001	10^{-9}

[a]Numbers that contain five or more digits to the right of the decimal point are written with a small space separating each group of three digits.
[b]How to express numbers in scientific notation is explained in Section 1.6.

The metric symbols are all lower case except for the unit **liter** (L) and the prefix **mega-** (M). Liter is capitalized to distinguish it from the number *one.* Mega is capitalized to distinguish it from the symbol for the prefix *milli-*.

PROBLEM 1.7

What term is used for each of the following units: (a) a million liters; (b) a thousandth of a second; (c) a hundredth of a gram; (d) a tenth of a liter?

PROBLEM 1.8

What is the numerical value of each unit in terms of the base unit? (For example, 1 μL = 0.000 001 L.)

a. 1 ng   b. 1 nm   c. 1 μs   d. 1 ML

1.4B MEASURING LENGTH

The base unit of length in the metric system is the *meter* (m). A meter, 39.4 inches in the English system, is slightly longer than a yard (36 inches). The three most common units derived from a meter are the kilometer (km), centimeter (cm), and millimeter (mm).

$$1{,}000 \text{ m} = 1 \text{ km}$$
$$1 \text{ m} = 100 \text{ cm}$$
$$1 \text{ m} = 1{,}000 \text{ mm}$$

Note how these values are related to those in Table 1.2. Since a centimeter is one *hundredth* of a meter (0.01 m), there are *100* centimeters in a meter.

PROBLEM 1.9

If a nanometer is one billionth of a meter (0.000 000 001 m), how many nanometers are there in one meter?

1.4C MEASURING MASS

Although the terms mass and weight are often used interchangeably, they really have different meanings.

- ***Mass* is a measure of the amount of matter in an object.**
- ***Weight* is the force that matter feels due to gravity.**

The mass of an object is independent of its location. The weight of an object changes slightly with its location on the earth, and drastically when the object is moved from the earth to the moon, where the gravitational pull is only one-sixth that of the earth. Although we often speak of *weighing* an object, we are really *measuring its mass.*

The basic unit of mass in the metric system is the *gram* (g), a small quantity compared to the English pound (1 lb = 454 g). The two most common units derived from a gram are the kilogram (kg) and milligram (mg).

$$1{,}000 \text{ g} = 1 \text{ kg}$$
$$1 \text{ g} = 1{,}000 \text{ mg}$$

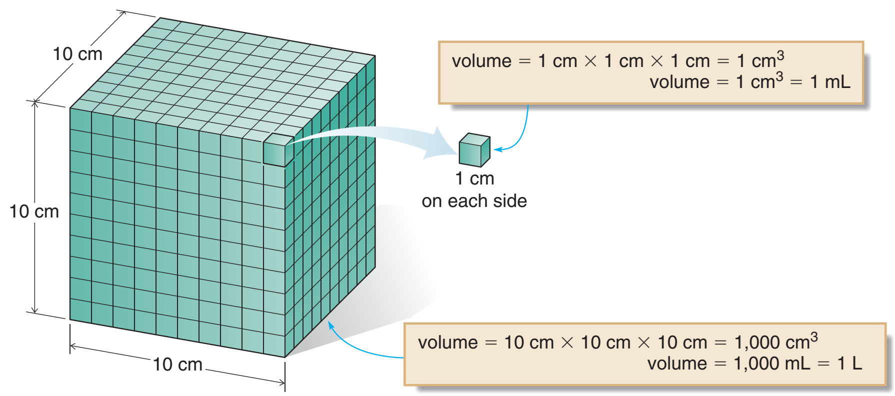
PROBLEM 1.10

If a microgram is one millionth of a gram (0.000 001 g), how many micrograms are there in one gram?

1.4D MEASURING VOLUME

The basic unit of volume in the metric system is the *liter* (L), which is slightly larger than the English quart (1 L = 1.06 qt). One liter is defined as the volume of a cube 10 cm on an edge.

Note the difference between the units **cm** and **cm^3**. The centimeter (cm) is a unit of length. A cubic centimeter (cm^3 or cc) is a unit of volume.



Three common units derived from a liter used in medicine and laboratory research are the deciliter (dL), milliliter (mL), and microliter (μL). **One milliliter is the same as one cubic centimeter (cm^3), which is abbreviated as cc.**

$$1 \text{ L} = 10 \text{ dL}$$
$$1 \text{ L} = 1{,}000 \text{ mL}$$
$$1 \text{ L} = 1{,}000{,}000 \ \mu\text{L}$$
$$1 \text{ mL} = 1 \text{ cm}^3 = 1 \text{ cc}$$

Table 1.3 summarizes common metric units of length, mass, and volume. Table 1.4 lists English units of measurement, as well as their metric equivalents.

PROBLEM 1.11 If a centiliter is one hundredth of a liter (0.01 L), how many centiliters are there in one liter?

TABLE 1.3 Summary of the Common Metric Units of Length, Mass, and Volume

Length	Mass	Volume
1 km = 1,000 m	1 kg = 1,000 g	1 L = 10 dL
1 m = 100 cm	1 g = 1,000 mg	1 L = 1,000 mL
1 m = 1,000 mm	1 mg = 1,000 μg	1 L = 1,000,000 μL
1 cm = 10 mm		1 dL = 100 mL
		1 mL = 1 cm^3 = 1 cc

1.5 SIGNIFICANT FIGURES

Numbers used in chemistry are either **exact** or **inexact.**

- **An *exact* number results from counting objects or is part of a definition.**

Our bodies have 10 fingers, 10 toes, and two kidneys. A meter is composed of 100 centimeters. These numbers are exact because there is no uncertainty associated with them.

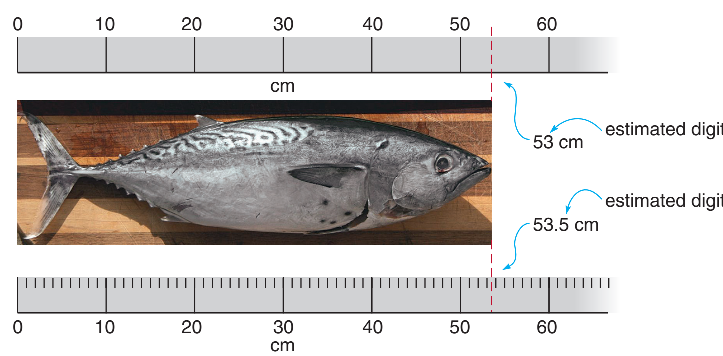
TABLE 1.4 English Units and Their Metric Equivalents

Quantity	English Unit	Metric–English Relationship
Length	1 ft = 12 in.	2.54 cm = 1 in.
	1 yd = 3 ft	1 m = 39.4 in.
	1 mi = 5,280 ft	1 km = 0.621 mi
Mass	1 lb = 16 oz	1 kg = 2.21 lb
	1 ton = 2,000 lb	454 g = 1 lb
		28.4 g = 1 oz
Volume	1 qt = 4 cups	946 mL = 1 qt
	1 qt = 2 pints	1 L = 1.06 qt
	1 qt = 32 fl oz	29.6 mL = 1 fl oz
	1 gal = 4 qt	

Common abbreviations for English units: inch (in.), foot (ft), yard (yd), mile (mi), pound (lb), ounce (oz), gallon (gal), quart (qt), and fluid ounce (fl oz).

- **An *inexact* number results from a measurement or observation and contains some uncertainty.**

Whenever we measure a quantity there is a degree of uncertainty associated with the result. The last number (furthest to the right) is an estimate, and it depends on the type of measuring device we use to obtain it. For example, the length of a fish caught on a recent outing could be reported as 53 cm or 53.5 cm depending on the tape measure used.



- ***Significant figures* are all the digits in a measured number including *one* estimated digit.**

Thus, the length 53 cm has two significant figures, and the length 53.5 cm has three significant figures.

1.5A DETERMINING THE NUMBER OF SIGNIFICANT FIGURES

How many significant figures are contained in a number?

- **All nonzero digits are always significant.**

65.2 g	three significant figures
1,265 m	four significant figures
25 μL	two significant figures
255.345 g	six significant figures

Whether a zero counts as a significant figure depends on its location in the number.

Rules to Determine When a Zero is a Significant Figure

Rule [1] A zero *counts* as a significant figure when it occurs:

- Between two nonzero digits — 29.05 g—four significant figures; 1.0087 mL—five significant figures
- At the end of a number with a decimal point — 25.70 cm—four significant figures; 3.7500 g—five significant figures; 620. lb—three significant figures

Rule [2] A zero does *not* count as a significant figure when it occurs:

- At the beginning of a number — 0.0245 mg—three significant figures; 0.008 mL—one significant figure
- At the end of a number that does not have a decimal point — 2,570 m—three significant figures; 1,245,500 m—five significant figures

In reading a number with a decimal point from left to right, all digits starting with the first nonzero number are significant figures. The number 0.003 450 120 has seven significant figures, shown in red.

SAMPLE PROBLEM 1.1

How many significant figures does each number contain?

a. 34.08 b. 0.0054 c. 260.00 d. 260

ANALYSIS All nonzero digits are significant. A zero is significant only if it occurs between two nonzero digits, or at the end of a number with a decimal point.

SOLUTION Significant figures are shown in red.

a. 34.08 (four) b. 0.0054 (two) c. 260.00 (five) d. 260 (two)

PROBLEM 1.12 How many significant figures does each number contain?

a. 23.45 b. 23.057 c. 230 d. 231.0 e. 0.202 f. 0.003 60 g. 1,245,006 h. 1,200,000

PROBLEM 1.13 How many significant figures does each number contain?

a. 10,040 b. 10,040. c. 1,004.00 d. 1.004 e. 1.0040 f. 0.1004 g. 0.001 004 h. 0.010 040 0

PROBLEM 1.14 Indicate whether each zero in the following numbers is significant.

a. 0.003 04 b. 26,045 c. 1,000,034 d. 0.304 00

1.5B USING SIGNIFICANT FIGURES IN MULTIPLICATION AND DIVISION

We often must perform calculations with numbers that contain a different number of significant figures. The number of significant figures in the answer of a problem depends on the type of mathematical calculation—multiplication (and division) or addition (and subtraction).

- **In multiplication and division, the answer has the same number of significant figures as the original number with the *fewest* significant figures.**

Let's say you drove a car 351.2 miles in 5.5 hours, and you wanted to calculate how many miles per hour you traveled. Entering these numbers on a calculator would give the following result:

four significant figures (351.2 miles)

$$\text{Miles per hour} = \frac{351.2 \text{ miles}}{5.5 \text{ hours}} = 63.854\ 545 \text{ miles per hour}$$

two significant figures (5.5 hours)

The answer must contain only **two** significant figures.

The answer to this problem can have only *two* significant figures, since one of the original numbers (5.5 hours) has only *two* significant figures. To write the answer in proper form, we must **round off the number** to give an answer with only two significant figures. Two rules are used in rounding off numbers.

- **If the first number that must be dropped is 4 or less, drop it and all remaining numbers.**
- **If the first number that must be dropped is 5 or greater, *round the number up* by adding one to the last digit that will be retained.**

In this problem:

These digits must be retained. (63)

first digit to be dropped (8)

63.854 545

These digits must be dropped. (854 545)

- Since the first digit to be dropped is 8 (a 5 or greater), add 1 to the digit to its left.
- The answer 63.854 545 rounded to two digits is **64 miles per hour.**

Table 1.5 gives other examples of rounding off numbers.

TABLE 1.5 Rounding Off Numbers

Original Number	Rounded to	Rounded Number
61.2537	Two places	61
61.2537	Three places	61.3
61.2537	Four places	61.25
61.2537	Five places	61.254

The first number to be dropped is indicated in red in each original number. When this number is 4 or fewer, it and all other digits to its right are dropped. When this number is 5 or greater, 1 is added to the digit to its left.

SAMPLE PROBLEM 1.2

Round off each number to three significant figures.

a. 1.2735 b. 0.002 536 22 c. 3,836.9

ANALYSIS If the answer is to have *three* significant figures, look at the *fourth* number from the left. If this number is 4 or less, drop it and all remaining numbers to the right. If the fourth number from the left is 5 or greater, round the number up by adding one to the third digit.

SOLUTION a. 1.27 b. 0.002 54 c. 3,840 (Omit the decimal point after the 0. The number 3,840. has four significant figures.)

PROBLEM 1.15 Round off each number in Sample Problem 1.2 to two significant figures.

SAMPLE PROBLEM 1.3 Carry out each calculation and give the answer using the proper number of significant figures.

a. 3.81×0.046 b. $120.085/106$

ANALYSIS Since these calculations involve multiplication and division, the answer must have the same number of significant figures as the original number with the fewest number of significant figures.

SOLUTION

a. $3.81 \times 0.046 = 0.1753$

- Since 0.046 has only two significant figures, round the answer to give it two significant figures.

0.1753 — Since this number is 5 (5 or greater), round the 7 to its left up by one.

Answer: 0.18

b. $120.085/106 = 1.132\ 877\ 36$

- Since 106 has three significant figures, round the answer to give it three significant figures.

1.132 877 36 — Since this number is 2 (4 or less), drop it and all numbers to its right.

Answer: 1.13

PROBLEM 1.16 Carry out each calculation and give the answer using the proper number of significant figures.

a. 10.70×3.5 b. $0.206/25{,}993$ c. $1{,}300/41.2$ d. 120.5×26

1.5C USING SIGNIFICANT FIGURES IN ADDITION AND SUBTRACTION

In determining significant figures in addition and subtraction, the decimal place of the last significant digit determines the number of significant figures in the answer.

- **In addition and subtraction, the answer has the same number of decimal places as the original number with the *fewest* decimal places.**

Suppose a baby weighed 3.6 kg at birth and 10.11 kg on his first birthday. To figure out how much weight the baby gained in his first year of life, we subtract these two numbers and report the answer using the proper number of significant figures.

weight at one year = 10.11 kg

weight at birth = 3.6 kg

10.11 kg (two digits after the decimal point)
− 3.6 kg (one digit after the decimal point)
weight gain = 6.51 kg

(last significant digit)

- The answer can have only **one** digit after the decimal point.
- Round 6.51 to 6.5.
- The baby gained 6.5 kg during his first year of life.

Since 3.6 kg has only one significant figure after the decimal point, the answer can have only one significant figure after the decimal point as well.

SAMPLE PROBLEM 1.4

While on a diet, a woman lost 3.52 lb the first week, 2.2 lb the second week, and 0.59 lb the third week. How much weight did she lose in all?

ANALYSIS Add up the amount of weight loss each week to get the total weight loss. When adding, the answer has the same number of decimal places as the original number with the fewest decimal places.

SOLUTION

3.52 lb
2.2 lb ← one digit after the decimal point
0.59 lb
6.31 lb - - - - - → 6.3 lb (round off)
↑ last significant digit

- Since 2.2 lb has only one digit after the decimal point, the answer can have only one digit after the decimal point.
- Round 6.31 to 6.3.
- Total weight loss: 6.3 lb.

PROBLEM 1.17

Carry out each calculation and give the answer using the proper number of significant figures.

a. 27.8 cm + 0.246 cm
b. 102.66 mL + 0.857 mL + 24.0 mL
c. 54.6 mg – 25 mg
d. 2.35 s – 0.266 s

1.6 SCIENTIFIC NOTATION

Hospital laboratory technicians determine thousands of laboratory results each day.

Healthcare professionals and scientists must often deal with very large and very small numbers. For example, the blood platelet count of a healthy adult might be 250,000 platelets per mL. At the other extreme, the level of the female sex hormone estriol during pregnancy might be 0.000 000 250 g per mL of blood plasma. Estriol is secreted by the placenta and its concentration is used as a measure of the health of the fetus.

To write numbers that contain many leading zeros (at the beginning) or trailing zeros (at the end), scientists use **scientific notation.**

- **In scientific notation, a number is written as $y \times 10^x$.**
- **The term *y*, called the coefficient, is a number between 1 and 10.**
- **The value *x* is an exponent, which can be any positive or negative whole number.**

First, let's recall what powers of 10 with *positive* exponents, such as 10^2 or 10^5, mean. These correspond to numbers greater than one, and the positive exponent tells how many zeros are to be written after the number one. Thus, $10^2 = 100$, a number with two zeros after the number one.

The product has two zeros.
$$10^2 = 10 \times 10 = 100$$
The exponent 2 means "multiply two 10s."

The product has five zeros.
$$10^5 = 10 \times 10 \times 10 \times 10 \times 10 = 100{,}000$$
The exponent 5 means "multiply five 10s."

Powers of 10 that contain *negative* exponents, such as 10^{-3}, correspond to numbers less than one. In this case the exponent tells how many places (*not* zeros) are located to the right of the decimal point.

The answer has three places to the right of the decimal point, including the number one.
$$10^{-3} = \frac{1}{10 \times 10 \times 10} = 0.001$$
The exponent −3 means "divide by three 10s."

To write a number in scientific notation, we follow a stepwise procedure.

HOW TO Convert a Standard Number to Scientific Notation

EXAMPLE **Write each number in scientific notation: (a) 2,500; (b) 0.036.**

Step [1] **Move the decimal point to give a number between 1 and 10.**

a. 2500.

Move the decimal point three places to the left to give the number 2.5.

b. 0.036

Move the decimal point two places to the right to give the number 3.6.

Step [2] **Multiply the result by 10^x, where *x* is the number of places the decimal point was moved.**

- If the decimal point is moved to the **left,** *x* is **positive.**
- If the decimal point is moved to the **right,** *x* is **negative.**

a. Since the decimal point was moved three places to the **left,** the exponent is +3, and the coefficient is multiplied by 10^3.

Answer: 2,500 = 2.5×10^3

b. Since the decimal point was moved two places to the **right,** the exponent is –2, and the coefficient is multiplied by 10^{-2}.

Answer: 0.036 = 3.6×10^{-2}

Notice that the number of significant figures in the coefficient in scientific notation must equal the number of significant figures in the original number. Thus, the coefficients for both 2,500 and 0.036 need two significant figures and no more. Table 1.6 shows how several numbers are written in scientific notation.

2,500 = 2.5×10^3 *not* 2.50×10^3 (three significant figures)

two significant figures *not* 2.500×10^3 (four significant figures)

TABLE 1.6 Numbers in Standard Form and Scientific Notation

Number	Scientific Notation
53,400	5.34×10^4
0.005 44	5.44×10^{-3}
3,500,000,000	3.5×10^9
0.000 000 000 123	1.23×10^{-10}
1,000.03	1.00003×10^3

SAMPLE PROBLEM 1.5

Write the recommended daily dietary intake of each nutrient in scientific notation: (a) sodium, 2,400 mg; (b) vitamin B_{12}, 0.000 006 g.

ANALYSIS

Move the decimal point to give a number between 1 and 10. Multiply the number by 10^x, where *x* is the number of places the decimal point was moved. The exponent *x* is (+) when the decimal point moves to the left and (–) when it moves to the right.

SOLUTION

a.

2400. = 2.4×10^3 ← the number of places the decimal point was moved to the left

Move the decimal point three places to the left.

- Write the coefficient as 2.4 (two significant figures), since 2,400 contains two significant figures.

b.

0.000 006 = 6×10^{-6} ← the number of places the decimal point was moved to the right

Move the decimal point six places to the right.

- Write the coefficient as 6 (one significant figure), since 0.000 006 contains one significant figure.

PROBLEM 1.18 Lab results for a routine check-up showed an individual's iron level in the blood to be 0.000 098 g per deciliter, placing it in the normal range. Convert this number to scientific notation.

PROBLEM 1.19 Write each number in scientific notation.

a. 93,200 c. 6,780,000 e. 4,520,000,000,000

b. 0.000 725 d. 0.000 030 f. 0.000 000 000 028

To convert a number in scientific notation to a standard number, reverse the procedure, as shown in Sample Problem 1.6. It is often necessary to add leading or trailing zeros to write the number.

- **When the exponent x is positive, move the decimal point x places to the *right*.**

2.800×10^2 2.800 ---→ 280.0

Move the decimal point to the right two places.

- **When the exponent x is negative, move the decimal point x places to the *left*.**

2.80×10^{-2} 002.80 ---→ 0.0280

Move the decimal point to the left two places.

SAMPLE PROBLEM 1.6 As we will learn in Chapter 4, the element hydrogen is composed of two hydrogen atoms, separated by a distance of 7.4×10^{-11} m. Convert this value to a standard number.

ANALYSIS The exponent in 10^x tells how many places to move the decimal point in the coefficient to generate a standard number. The decimal point goes to the right when x is positive and to the left when x is negative.

SOLUTION 7.4×10^{-11} 000 000 000 07.4 ---→ 0.000 000 000 074 m

Move the decimal point to the left 11 places. **Answer:**

The answer, 0.000 000 000 074, has two significant figures, just like 7.4×10^{-11}.

PROBLEM 1.20 There are 6.02×10^{21} "particles" called molecules (Chapter 4) of aspirin in 1.8 g. Write this number in standard form.

PROBLEM 1.21 Convert each number to its standard form.

a. 6.5×10^3 c. 3.780×10^{-2} e. 2.221×10^6

b. 3.26×10^{-5} d. 1.04×10^8 f. 4.5×10^{-10}

1.7 PROBLEM SOLVING USING THE FACTOR–LABEL METHOD

Often a measurement is recorded in one unit, and then it must be converted to another unit. For example, a patient may weigh 130 lb, but we may need to know her weight in kilograms to calculate a drug dosage. The recommended daily dietary intake of potassium is 3,500 mg, but we may need to know how many grams this corresponds to.

1.7A CONVERSION FACTORS

To convert one unit to another we use one or more **conversion factors.**

original quantity × **conversion factor** = desired quantity

These two quantities are equivalent. Only the units are different.

- **A *conversion factor* is a term that converts a quantity in one unit to a quantity in another unit.**

A conversion factor is formed by taking an equality, such as 2.21 lb = 1 kg, and writing it as a fraction. We can always write a conversion factor in two different ways.

$$\frac{2.21\text{ lb}}{1\text{ kg}} \quad \text{or} \quad \frac{1\text{ kg}}{2.21\text{ lb}}$$

numerator (1 kg), denominator (2.21 lb)

conversion factors for pounds and kilograms

Refer to Tables 1.3 and 1.4 for metric and English units needed in problem solving. Common metric and English units are also listed on the inside back cover.

With pounds and kilograms, either of these values can be written above the division line of the fraction (the numerator) or below the division line (the denominator). The way the conversion factor is written will depend on the problem. Since the values above and below the division line are *equivalent,* a conversion factor always equals one.

SAMPLE PROBLEM 1.7

Write two conversion factors for each pair of units: (a) kilograms and grams; (b) quarts and liters.

ANALYSIS Use the equalities in Tables 1.3 and 1.4 to write a fraction that shows the relationship between the two units.

SOLUTION

a. Conversion factors for kilograms and grams:

$$\frac{1000\text{ g}}{1\text{ kg}} \quad \text{or} \quad \frac{1\text{ kg}}{1000\text{ g}}$$

b. Conversion factors for quarts and liters:

$$\frac{1.06\text{ qt}}{1\text{ L}} \quad \text{or} \quad \frac{1\text{ L}}{1.06\text{ qt}}$$

PROBLEM 1.22

Write two conversion factors for each pair of units.

a. miles and kilometers
b. meters and millimeters
c. grams and pounds
d. milligrams and micrograms

1.7B SOLVING A PROBLEM USING ONE CONVERSION FACTOR

Using conversion factors to convert a quantity in one unit to a quantity in another unit is called the **factor–label method.** In this method, **units are treated like numbers.** As a result, if a unit appears in the numerator in one term and the denominator in another term, the units *cancel*. **The goal in setting up a problem is to make sure *all unwanted units cancel.***

Let's say we want to convert 130 lb to kilograms.

130 lb (original quantity) × **conversion factor** = ? kg (desired quantity)

Two possible conversion factors:

$$\frac{2.21\text{ lb}}{1\text{ kg}} \quad \text{or} \quad \frac{1\text{ kg}}{2.21\text{ lb}}$$

To solve this problem we must use a conversion factor that satisfies two criteria.

- **The conversion factor must relate the two quantities in question—pounds and kilograms.**
- **The conversion factor must cancel out the unwanted unit—pounds.**

How many grams of aspirin are contained in a 325-mg tablet?

This means choosing the conversion factor with the unwanted unit—pounds—*in the denominator* to cancel out pounds in the original quantity. This leaves kilograms as the only remaining unit, and the problem is solved.

conversion factor

$$130\ \cancel{\text{lb}} \times \frac{1\ \text{kg}}{2.21\ \cancel{\text{lb}}} = \boxed{59\ \text{kg}}$$ **answer in kilograms**

Pounds (lb) must be the denominator to cancel the unwanted unit (lb) in the original quantity.

We must use the correct number of significant figures in reporting an answer to each problem. In this case, the value 1 kg is *defined* as 2.21 lb; in other words, 1 kg contains the exact number "1" with *no* uncertainty, so it does not limit the number of digits in the answer. Since 130 lb has two significant figures, the answer is rounded to two significant figures (59 kg).

As problems with units get more complicated, keep in mind the following general steps that are useful for solving any problem using the factor–label method.

HOW TO Solve a Problem Using Conversion Factors

EXAMPLE **How many grams of aspirin are contained in a 325-mg tablet?**

Step [1] **Identify the original quantity and the desired quantity, including units.**

- In this problem the original quantity is reported in milligrams and the desired quantity is in grams.

325 mg	? g
original quantity	desired quantity

Step [2] **Write out the conversion factor(s) needed to solve the problem.**

- We need a conversion factor that relates milligrams and grams (Table 1.3). Since the unwanted unit is in milligrams, **choose the conversion factor that contains milligrams in the denominator so that the *units cancel*.**

Two possible conversion factors: $\frac{1000\ \text{mg}}{1\ \text{g}}$ or $\boxed{\frac{1\ \text{g}}{1000\ \text{mg}}}$ Choose this factor to cancel the unwanted unit, mg.

- Sometimes one conversion factor is all that is needed in a problem. At other times (Section 1.7C) more than one conversion factor is needed.
- If the desired answer has a single unit (grams in this case), **the conversion factor must contain the desired unit in the numerator and the unwanted unit in the denominator.**

Step [3] **Set up and solve the problem.**

- Multiply the original quantity by the conversion factor to obtain the desired quantity.

conversion factor

$$325\ \cancel{\text{mg}} \times \frac{1\ \text{g}}{1000\ \cancel{\text{mg}}} = 0.325\ \text{g of aspirin}$$

original quantity → desired quantity

The number of mg (unwanted unit) cancels.

Step [4] **Write the answer using the correct number of significant figures and check it by estimation.**

- Use the number of significant figures in each inexact (measured) number to determine the number of significant figures in the answer. In this case the answer is limited to three significant figures by the original quantity (325 mg).
- Estimate the answer using a variety of methods. In this case we knew our answer had to be less than one, since it is obtained by dividing 325 by a number larger than itself.

PROBLEM 1.23 The distance between Honolulu, HI, and Los Angeles, CA, is 4,120 km. How many frequent flyer miles will you earn by traveling between the two cities?

PROBLEM 1.24 Convert 25 mL to μL and write the answer in scientific notation.

PROBLEM 1.25 Carry out each of the following conversions.

a. 25 L to dL　　b. 40.0 oz to g　　c. 32 in. to cm　　d. 10 cm to mm

1.7C SOLVING A PROBLEM USING TWO OR MORE CONVERSION FACTORS

Some problems require the use of more than one conversion factor to obtain the desired units in the answer. The same four-step procedure is followed no matter how many conversion factors are needed. Keep in mind:

- **Always arrange the factors so that the denominator in one term cancels the numerator in the preceding term.**

Sample Problem 1.8 illustrates how to solve a problem with two conversion factors.

SAMPLE PROBLEM 1.8 An individual donated 1.0 pint of blood at the local blood bank. How many liters of blood does this correspond to?

ANALYSIS AND SOLUTION

[1] **Identify the original quantity and the desired quantity.**

1.0 pt (original quantity)　　? L (desired quantity)

[2] **Write out the conversion factors.**

- We have no conversion factor that directly relates pints to liters. We do, however, know conversions for pints to quarts, and quarts to liters.

pint–quart conversion: $\frac{2\text{ pt}}{1\text{ qt}}$ or $\frac{1\text{ qt}}{2\text{ pt}}$

quart–liter conversion: $\frac{1.06\text{ qt}}{1\text{ L}}$ or $\frac{1\text{ L}}{1.06\text{ qt}}$

Choose the conversion factors with the unwanted units—pt and qt—in the denominator.

How many liters of blood does this pint of blood contain?

[3] **Solve the problem.**

- To set up the problem so that unwanted units cancel, arrange each term so that the units in the numerator of one term cancel the units of the denominator of the adjacent term. In this problem we need to cancel both pints and quarts to get liters.
- The single desired unit, liters, must be in the **numerator** of one term.

$$1.0\text{ pt} \times \frac{1\text{ qt}}{2\text{ pt}} \times \frac{1\text{ L}}{1.06\text{ qt}} = 0.47\text{ L}$$

Pints cancel. Quarts cancel. Liters do not cancel.

[4] **Check.**

- Since there are two pints in a quart and a quart is about the same size as a liter, one pint should be about half a liter. The answer, 0.47, is just about 0.5.
- Write the answer with two significant figures since one term, 1.0 pt, has two significant figures.

PROBLEM 1.26 Carry out each of the following conversions.

a. 6,250 ft to km   b. 3 cups to L   c. 4.5 ft to cm

PROBLEM 1.27 On a recent road trip, your average speed was 65 miles per hour. What was your average speed in (a) mi/s; (b) m/s?

1.8 FOCUS ON HEALTH & MEDICINE PROBLEM SOLVING USING CLINICAL CONVERSION FACTORS

Sometimes conversion factors don't have to be looked up in a table; they are stated in the problem. If a drug is sold as a 250-mg tablet, this fact becomes a conversion factor relating milligrams to tablets.

$$\frac{250 \text{ mg}}{1 \text{ tablet}} \quad \text{or} \quad \frac{1 \text{ tablet}}{250 \text{ mg}}$$

mg–tablet conversion factors

The active ingredient in Children's Tylenol is acetaminophen.

Alternatively, a drug could be sold as a liquid solution with a specific concentration. For example, Children's Tylenol contains 80 mg of the active ingredient acetaminophen in 2.5 mL. This fact becomes a conversion factor relating milligrams to milliliters.

$$\frac{80 \text{ mg}}{2.5 \text{ mL}} \quad \text{or} \quad \frac{2.5 \text{ mL}}{80 \text{ mg}}$$

mg of acetaminophen–mL conversion factors

Sample Problems 1.9 and 1.10 illustrate how these conversion factors are used in determining drug dosages.

SAMPLE PROBLEM 1.9 A patient is prescribed 1.25 g of amoxicillin, which is available in 250-mg tablets. How many tablets are needed?

ANALYSIS AND SOLUTION

[1] **Identify the original quantity and the desired quantity.**

- We must convert the number of grams of amoxicillin needed to the number of tablets that must be administered.

1.25 g	? tablets
original quantity	desired quantity

[2] **Write out the conversion factors.**

- We have no conversion factor that directly relates grams to tablets. We do know, however, how to relate grams to milligrams, and milligrams to tablets.

g–mg conversion factors

$$\frac{1 \text{ g}}{1000 \text{ mg}} \quad \text{or} \quad \frac{1000 \text{ mg}}{1 \text{ g}}$$

mg–tablet conversion factors

$$\frac{250 \text{ mg}}{1 \text{ tablet}} \quad \text{or} \quad \frac{1 \text{ tablet}}{250 \text{ mg}}$$

Choose the conversion factors with the unwanted units–g and mg–in the denominator.

[3] **Solve the problem.**

- Arrange each term so that the units in the numerator of one term cancel the units in the denominator of the adjacent term. In this problem we need to cancel both grams and milligrams to get tablets.
- The single desired unit, tablets, must be located in the **numerator** of one term.

$$1.25\ \cancel{g} \times \frac{1000\ \cancel{mg}}{1\ \cancel{g}} \times \frac{1\ \text{tablet}}{250\ \cancel{mg}} = 5\ \text{tablets}$$

Grams cancel. Milligrams cancel. Tablets do not cancel.

[4] **Check.**

- The answer of 5 tablets of amoxicillin (not 0.5 or 50) is reasonable. Since the dose in a single tablet (250 mg) is a fraction of a gram, and the required dose is more than a gram, the answer must be greater than one.

SAMPLE PROBLEM 1.10

A dose of 240 mg of acetaminophen is prescribed for a 20-kg child. How many mL of Children's Tylenol (80. mg of acetaminophen per 2.5 mL) are needed?

ANALYSIS AND SOLUTION

[1] **Identify the original quantity and the desired quantity.**

- We must convert the number of milligrams of acetaminophen needed to the number of mL that must be administered.

240 mg	? mL
original quantity	desired quantity

[2] **Write out the conversion factors.**

mg of acetaminophen–mL conversion factors

$$\frac{80.\ \text{mg}}{2.5\ \text{mL}} \quad \text{or} \quad \boxed{\frac{2.5\ \text{mL}}{80.\ \text{mg}}}$$

Choose the conversion factor to cancel mg.

[3] **Solve the problem.**

- Arrange the terms so that the units in the numerator of one term cancel the units of the denominator of the adjacent term. In this problem we need to cancel milligrams to obtain milliliters.
- In this problem we are given a fact we don't need to use—the child weighs 20 kg. We can ignore this quantity in carrying out the calculation.

$$240\ \cancel{mg} \times \frac{2.5\ \text{mL}}{80.\ \cancel{mg}} = 7.5\ \text{mL of Children's Tylenol}$$

Milligrams cancel.

[4] **Check.**

- The answer of 7.5 mL (not 0.75 or 75) is reasonable. Since the required dose is larger than the dose in 2.5 mL, the answer must be larger than 2.5 mL.

PROBLEM 1.28

If one teaspoon contains 5.0 mL, how many teaspoons of Children's Tylenol must be administered in Sample Problem 1.10?

PROBLEM 1.29 A patient is prescribed 0.100 mg of a drug that is available in 25-μg tablets. How many tablets are needed?

PROBLEM 1.30 How many milliliters of Children's Motrin (100 mg of ibuprofen per 5 mL) are needed to give a child a dose of 160 mg?

1.9 TEMPERATURE

Temperature is a measure of how hot or cold an object is. Three temperature scales are used: **Fahrenheit** (most common in the United States), **Celsius** (most commonly used by scientists and countries other than the United States), and **Kelvin** (Figure 1.7).

The Fahrenheit and Celsius scales are both divided into **degrees.** On the Fahrenheit scale, water freezes at 32 °F and boils at 212 °F. On the Celsius scale, water freezes at 0 °C and boils at 100 °C. To convert temperature values from one scale to another, we use two equations, where °C is the Celsius temperature and °F is the Fahrenheit temperature.

To convert from Celsius to Fahrenheit:

$$°F = 1.8(°C) + 32$$

To convert from Fahrenheit to Celsius:

$$°C = \frac{°F - 32}{1.8}$$

The Kelvin scale is divided into **kelvins** (K), not degrees. The only difference between the Kelvin scale and the Celsius scale is the zero point. A temperature of –273 °C corresponds to 0 K. The zero point on the Kelvin scale is called **absolute zero,** the lowest temperature possible. To convert temperature values from Celsius to Kelvin, or vice versa, use two equations.

To convert from Celsius to Kelvin:

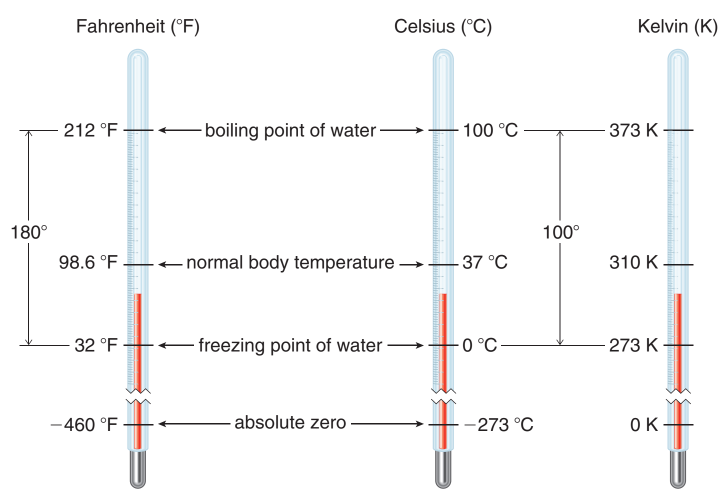
$$K = °C + 273$$

To convert from Kelvin to Celsius:

$$°C = K - 273$$

Although mercury thermometers were used in hospitals to measure temperature for many years, temperature is now more commonly recorded with a digital thermometer. Tympanic thermometers, which use an infrared sensing device placed in the ear, are also routinely used.

FIGURE 1.7 Fahrenheit, Celsius, and Kelvin Temperature Scales Compared



Since the freezing point and boiling point of water span 180° on the Fahrenheit scale, but only 100° on the Celsius scale, a Fahrenheit degree and a Celsius degree differ in size. The Kelvin scale is divided into kelvins (K), not degrees. Since the freezing point and boiling point of water span 100 kelvins, one kelvin is the same size as one Celsius degree.

SAMPLE PROBLEM 1.11

An infant had a temperature of 104 °F. Convert this temperature to both °C and K.

ANALYSIS

First convert the Fahrenheit temperature to degrees Celsius using the equation °C = (°F – 32)/1.8. Then convert the Celsius temperature to kelvins by adding 273.

SOLUTION

[1] **Convert °F to °C:**

$$°C = \frac{°F - 32}{1.8}$$

$$= \frac{104 - 32}{1.8} = 40.\ °C$$

[2] **Convert °C to K:**

$$K = °C + 273$$

$$= 40. + 273 = 313\ K$$

PROBLEM 1.31

When the human body is exposed to extreme cold, hypothermia can result and the body's temperature can drop to 28.5 °C. Convert this temperature to °F and K.

PROBLEM 1.32

Convert each temperature to the requested temperature scale.

a. 20 °C to °F
b. 150 °F to °C
c. 298 K to °F
d. 75 °C to K

1.10 DENSITY AND SPECIFIC GRAVITY

Two additional quantities used to characterize substances are **density** and **specific gravity.**

1.10A DENSITY

***Density* is a physical property that relates the mass of a substance to its volume.** Density is reported in grams per milliliter (g/mL) or grams per cubic centimeter (g/cc).

$$\textbf{density} = \frac{\textbf{mass (g)}}{\textbf{volume (mL or cc)}}$$

The density of a substance depends on temperature. For most substances, the solid state is more dense than the liquid state, and as the temperature increases, the density decreases. This phenomenon occurs because the volume of a sample of a substance generally increases with temperature but the mass is always constant.

Water is an exception to this generalization. Solid water, ice, is *less* dense than liquid water, and from 0 °C to 4 °C, the density of water *increases*. Above 4 °C, water behaves like other liquids and its density decreases. Thus, water's maximum density of 1.000 g/mL occurs at 4 °C. Some representative densities are reported in Table 1.7.

TABLE 1.7 Representative Densities at 25 °C

Substance	Density [g/(mL or cc)]	Substance	Density [g/(mL or cc)]
Oxygen (0 °C)	0.001 43	Urine	1.003–1.030
Gasoline	0.66	Blood plasma	1.03
Ice (0 °C)	0.92	Table sugar	1.59
Water (4 °C)	1.00	Bone	1.80

Although a can of a diet soft drink floats in water because it is less dense, a can of a regular soft drink that contains sugar is more dense than water so it sinks.

The density (not the mass) of a substance determines whether it floats or sinks in a liquid.

- **A less dense substance floats on a more dense liquid.**

Ice floats on water because it is less dense. When petroleum leaks from an oil tanker or gasoline is spilled when fueling a boat, it floats on water because it is less dense. In contrast, a cannonball or torpedo sinks because it is more dense than water.

Knowing the density of a liquid allows us to convert the volume of a substance to its mass, or the mass of a substance to its volume.

To convert volume (mL) to mass (g):

density

$$\cancel{\text{mL}} \times \frac{\text{g}}{\cancel{\text{mL}}} = \text{g}$$

Milliliters cancel.

To convert mass (g) to volume (mL):

inverse of the density

$$\cancel{\text{g}} \times \frac{\text{mL}}{\cancel{\text{g}}} = \text{mL}$$

Grams cancel.

For example, one laboratory synthesis of aspirin uses the liquid acetic acid, which has a density of 1.05 g/mL. If we need 5.0 g for a synthesis, we could use density to convert this mass to a volume that could then be easily measured out using a syringe or pipette.

$$5.0\ \cancel{\text{g}}\ \text{acetic acid} \times \frac{1\ \text{mL}}{1.05\ \cancel{\text{g}}} = 4.8\ \text{mL of acetic acid}$$

Grams cancel.

SAMPLE PROBLEM 1.12 Calculate the mass in grams of 15.0 mL of a saline solution that has a density 1.05 g/mL.

ANALYSIS Use density (g/mL) to interconvert the mass and volume of a liquid.

SOLUTION

density

$$15.0\ \cancel{\text{mL}} \times \frac{1.05\ \text{g}}{1\ \cancel{\text{mL}}} = 15.8\ \text{g of saline solution}$$

Milliliters cancel.

The answer, 15.8 g, is rounded to three significant figures to match the number of significant figures in both factors in the problem.

PROBLEM 1.33 Calculate the mass in grams of 10.0 mL of diethyl ether, an anesthetic that has a density of 0.713 g/mL.

PROBLEM 1.34 (a) Calculate the volume in milliliters of 100. g of coconut oil, which has a density of 0.92 g/mL. (b) How many liters does this correspond to?

PROBLEM 1.35 Ten milliliters of either hexane (density = 0.65 g/mL) or chloroform (density = 1.49 g/mL) was added to a beaker that contains 10 mL of water, forming two layers with water on top. What liquid was added to the beaker?

1.10B SPECIFIC GRAVITY

***Specific gravity* is a quantity that compares the density of a substance with the density of water at the same temperature.**

$$\text{specific gravity} = \frac{\text{density of a substance (g/mL)}}{\text{density of water (g/mL)}}$$

Unlike most other quantities, specific gravity is a quantity without units, since the units in the numerator (g/mL) cancel the units in the denominator (g/mL). Since the density of water is 1.00 g/mL at and around room temperature, **the specific gravity of a substance equals its density, but it contains no units.** For example, if the density of a liquid is 1.5 g/mL at 20 °C, its specific gravity is 1.5.

The specific gravity of urine samples is often measured in a hospital lab. Normal urine has a density in the range of 1.003–1.030 g/mL (Table 1.7), so it has a specific gravity in the range of 1.003–1.030. Consistently high or low values can indicate an imbalance in metabolism. For example, the specific gravity of urine samples from patients with poorly controlled diabetes is abnormally high, because a large amount of glucose is excreted in the urine.

PROBLEM 1.36 (a) If the density of a liquid is 0.80 g/mL, what is its specific gravity? (b) If the specific gravity of a substance is 2.3, what is its density?

CHAPTER HIGHLIGHTS

KEY TERMS

Celsius scale (1.9)
Chemical properties (1.2)
Chemistry (1.1)
Compound (1.3)
Conversion factor (1.7)
Cubic centimeter (1.4)
Density (1.10)
Element (1.3)
English system of measurement (1.4)
Exact number (1.5)
Factor–label method (1.7)
Fahrenheit scale (1.9)
Gas (1.2)
Gram (1.4)
Inexact number (1.5)
Kelvin scale (1.9)
Liquid (1.2)
Liter (1.4)
Mass (1.4)
Matter (1.1)
Meter (1.4)
Metric system (1.4)
Mixture (1.3)
Physical properties (1.2)
Pure substance (1.3)
Scientific notation (1.6)
SI units (1.4)
Significant figures (1.5)
Solid (1.2)
Specific gravity (1.10)
States of matter (1.2)
Temperature (1.9)
Weight (1.4)

KEY CONCEPTS

❶ Describe the three states of matter. (1.1, 1.2)

- Matter is anything that has mass and takes up volume. Matter has three common states:
  - The solid state is composed of highly organized particles that lie close together. A solid has a definite shape and volume.
  - The liquid state is composed of particles that lie close together but are less organized than the solid state. A liquid has a definite volume but not a definite shape.
  - The gas state is composed of highly disorganized particles that lie far apart. A gas has no definite shape or volume.

❷ How is matter classified? (1.3)

- Matter is classified in one of two categories:
  - A pure substance is composed of a single component with a constant composition. A pure substance is either an element, which cannot be broken down into simpler substances by a chemical reaction, or a compound, which is formed by combining two or more elements.
  - A mixture is composed of more than one component and its composition can vary depending on the sample.

❸ What are the key features of the metric system of measurement? (1.4)

- The metric system is a system of measurement in which each type of measurement has a base unit and all other units are related to the base unit by a prefix that indicates if the unit is larger or smaller than the base unit.
- The base units are meter (m) for length, gram (g) for mass, liter (L) for volume, and second (s) for time.

❹ What are significant figures and how are they used in calculations? (1.5)

- Significant figures are all digits in a measured number, including one estimated digit. All nonzero digits are significant. A zero is significant only if it occurs between two nonzero digits, or at the end of a number with a decimal point. A trailing zero in a number without a decimal point is not considered significant.
- In multiplying and dividing with significant figures, the answer has the same number of significant figures as the original number with the fewest significant figures.
- In adding or subtracting with significant figures, the answer has the same number of decimal places as the original number with the fewest decimal places.

❺ What is scientific notation? (1.6)

- Scientific notation is a method of writing a number as $y \times 10^x$, where y is a number between 1 and 10, and x is a positive or negative exponent.
- To convert a standard number to a number in scientific notation, move the decimal point to give a number between 1 and 10. Multiply the result by 10^x, where x is the number of places the decimal point was moved. When the decimal point is moved to the left, x is positive. When the decimal point is moved to the right, x is negative.

❻ How are conversion factors used to convert one unit to another? (1.7, 1.8)

- A conversion factor is a term that converts a quantity in one unit to a quantity in another unit. To use conversion factors to solve a problem, set up the problem with any unwanted unit in the numerator of one term and the denominator of another term, so that unwanted units cancel.

❼ What is temperature and how are the three temperature scales related? (1.9)

- Temperature is a measure of how hot or cold an object is. The Fahrenheit and Celsius temperature scales are divided into degrees. Both the size of the degree and the zero point of these scales differ. The Kelvin scale is divided into kelvins, and one kelvin is the same size as one degree Celsius.

❽ What are density and specific gravity? (1.10)

- Density is a physical property reported in g/mL or g/cc that relates the mass of an object to its volume. A less dense substance floats on top of a more dense liquid.
- Specific gravity is a unitless quantity that relates the density of a substance to the density of water at the same temperature. Since the density of water is 1.00 g/mL at common temperatures, the specific gravity of a substance equals its density, but it contains no units.

PROBLEMS

Selected in-chapter and end-of-chapter problems have brief answers provided in Appendix B.

Matter

1.37 What is the difference between an element and a compound?

1.38 What is the difference between a compound and a mixture?

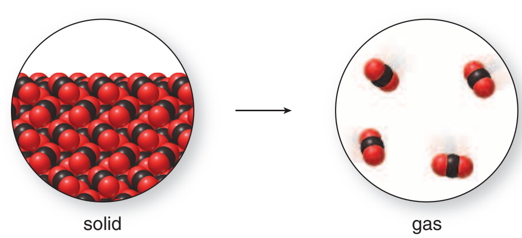
1.39 Describe solids, liquids, and gases in terms of (a) volume (how they fill a container); (b) shape; (c) level of organization of the particles that comprise them; (d) how close the particles that comprise them lie.

1.40 How do physical properties and chemical properties differ?

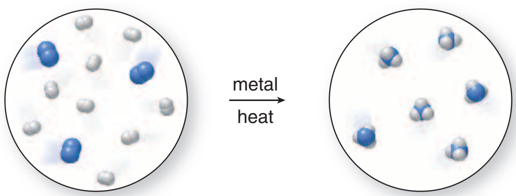
1.41 Classify each process as a chemical or physical change.
a. dissolving calcium chloride in water
b. burning gasoline to power a car
c. heating wax so that it melts

1.42 Classify each process as a chemical or physical change.
a. the condensation of water on the outside of a cold glass
b. mixing a teaspoon of instant coffee with hot water
c. baking a cake

1.43 When a chunk of dry ice (solid carbon dioxide) is placed out in the air, the solid gradually disappears and a gas is formed above the solid. Does the molecular art drawn below indicate that a chemical or physical change has occurred? Explain your choice.



1.44 The inexpensive preparation of nitrogen-containing fertilizers begins with mixing together two elements, hydrogen and nitrogen, at high temperature and pressure in the presence of a metal. Does the molecular art depicted below indicate that a chemical or physical change occurs under these conditions? Explain your choice.



Measurement

1.45 What is the difference between an exact number and an inexact number? Give an example of each type of number.

1.46 Label each quantity as an exact or inexact number.
a. A recipe requires 10 cloves of garlic and two tablespoons of oil.
b. A dog had five puppies whose combined weight was 10 lb.
c. The four bicycles in the family have been ridden for a total of 250 mi.
d. A child fell and had a 4 cm laceration that required 12 stitches.

1.47 Which quantity in each pair is larger?
a. 5 mL or 5 dL c. 5 cm or 5 mm
b. 10 mg or 10 μg d. 10 Ms or 10 ms

1.48 Which quantity in each pair is larger?
a. 10 km or 10 m c. 10 g or 10 μg
b. 10 L or 10 mL d. 10 cm or 10 mm

Significant Figures

1.49 How many significant figures does each number contain?
a. 16.00 c. 0.001 60 e. 1.06 g. 1.060×10^{10}
b. 160 d. 1,600,000 f. 0.1600 h. 1.6×10^{-6}

1.50 How many significant figures does each number contain?
a. 160. c. 0.000 16 e. 1,600. g. 1.600×10^{-10}
b. 160.0 d. 1.60 f. 1.060 h. 1.6×10^{6}

1.51 Round each number to three significant figures.
a. 25,401 c. 0.001 265 982 e. 195.371
b. 1,248,486 d. 0.123 456 f. 196.814

1.52 Round each number in Problem 1.51 to four significant figures.

1.53 Carry out each calculation and report the answer using the proper number of significant figures.
a. 53.6×0.41 c. 65.2/12 e. 694.2×0.2
b. 25.825 – 3.86 d. 41.0 + 9.135 f. 1,045 – 1.26

1.54 Carry out each calculation and report the answer using the proper number of significant figures.
a. $49{,}682 \times 0.80$ c. 1,000/2.34 e. 25,000/0.4356
b. 66.815 + 2.82 d. 21 – 0.88 f. 21.5381 + 26.55

Scientific Notation

1.55 Write each quantity in scientific notation.
a. 1,234 g c. 5,244,000 L e. 44,000 km
b. 0.000 016 2 m d. 0.005 62 g

1.56 Write each quantity in scientific notation.
a. 0.001 25 m c. 54,235.6 m e. 4,440 s
b. 8,100,000,000 lb d. 0.000 001 899 L

1.57 Convert each number to its standard form.
a. 3.4×10^{8} c. 3×10^{2}
b. 5.822×10^{-5} d. 6.86×10^{-8}

1.58 Convert each number to its standard form.
a. 4.02×10^{10} c. 6.86×10^{9}
b. 2.46×10^{-3} d. 1.00×10^{-7}

1.59 Which number in each pair is larger?
a. 4.44×10^{3} or 4.8×10^{2} c. 1.3×10^{8} or 52,300,000
b. 5.6×10^{-6} or 5.6×10^{-5} d. 9.8×10^{-4} or 0.000 089

1.60 Rank the numbers in each group from smallest to largest.
a. 5.06×10^{6}, 7×10^{4}, and 2.5×10^{8}
b. 6.3×10^{-2}, 2.5×10^{-4}, and 8.6×10^{-6}

1.61 Write the recommended daily intake of each nutrient in scientific notation.
a. 0.000 400 g of folate c. 0.000 080 g of vitamin K
b. 0.002 g of copper d. 3,400 mg of chloride

1.62 A blood vessel is 0.40 μm in diameter. (a) Convert this quantity to meters and write the answer in scientific notation. (b) Convert this quantity to inches and write the answer in scientific notation.

1.63 A picosecond is one trillionth of a second (0.000 000 000 001 s). (a) Write this number in scientific notation. (b) How many picoseconds are there in one second? Write this answer in scientific notation.

1.64 Red light has a wavelength of 683 nm. Convert this quantity to meters and write the answer in scientific notation.

Problem Solving and Unit Conversions

1.65 Carry out each of the following conversions.
a. 300 g to mg d. 300 g to oz
b. 2 L to μL e. 2 ft to m
c. 5.0 cm to m f. 3.5 yd to m

1.66 Carry out each of the following conversions.
a. 25 μL to mL d. 300 mL to qt
b. 35 kg to g e. 3 cups to L
c. 2.36 mL to L f. 2.5 tons to kg

1.67 Carry out each of the following conversions.
a. What is the mass in kilograms of an individual who weighs 234 lb?
b. What is the height in centimeters of a child who is 50. in. tall?
c. A patient required 3.0 pt of blood during surgery. How many liters does this correspond to?
d. A patient had a body temperature of 37.7 °C. What is his body temperature in °F?

1.68 Carry out each of the following conversions.
a. What is the mass in pounds of an individual who weighs 53.2 kg?
b. What is the height in inches of a child who is 90. cm tall?
c. How many mL are contained in the 5.0 qt of blood in the human body?
d. A patient had a body temperature of 103.5 °F. What is his body temperature in °C?

1.69 (a) How many milliliters are contained in 1 qt of milk? (b) How many fluid ounces are contained in 1 L of soda?

1.70 Which gasoline is less expensive: gas that sells for $3.00 per gallon or gas that sells for $0.89 per liter?

1.71 The average mass of a human liver is 1.5 kg. Convert this quantity to (a) grams; (b) pounds; (c) ounces.

1.72 The length of a femur (thigh bone) of a patient is 18.2 in. Convert this quantity to (a) meters; (b) centimeters.

Temperature

1.73 Carry out each of the following temperature conversions.
a. An over-the-counter pain reliever melts at 53 °C. Convert this temperature to °F and K.
b. A cake is baked at 350 °F. Convert this temperature to °C and K.

1.74 Methane, the main component of the natural gas used for cooking and heating homes, melts at –183 °C and boils at –162 °C. Convert each temperature to °F and K.

1.75 Which temperature in each pair is higher?
a. –10 °C or 10 °F   b. –50 °C or –50 °F

1.76 Rank the temperatures in each group from lowest to highest.
a. 0 °F, 0 °C, 0 K   b. 100 K, 100 °C, 100 °F

Density and Specific Gravity

1.77 What is the difference between density and specific gravity?

1.78 If you have an equal mass of two different substances (**A** and **B**), but the density of **A** is twice the density of **B,** what can be said about the volumes of **A** and **B?**

1.79 If a urine sample has a mass of 122 g and a volume of 121 mL, what is its density in g/mL?

1.80 The density of sucrose, table sugar, is 1.56 g/cc. What volume (in cubic centimeters) does 20.0 g of sucrose occupy?

1.81 Isooctane is a high-octane component of gasoline. If the density of isooctane is 0.692 g/mL, how much does 220 mL weigh?

1.82 A volume of saline solution weighed 25.6 g at 4 °C. An equal volume of water at the same temperature weighed 24.5 g. What is the density of the saline solution?

1.83 If milk has a density of 1.03 g/mL, what is the mass of 1 qt, reported in kilograms?

1.84 If gasoline has a density of 0.66 g/mL, how many kilograms does 1 gal weigh?

1.85 Which is the upper layer when each of the following liquids is added to water?
a. heptane (density = 0.684 g/mL)
b. olive oil (density = 0.92 g/mL)
c. chloroform (density = 1.49 g/mL)
d. carbon tetrachloride (density = 1.59 g/mL)

1.86 Which of the following solids float on top of water and which sink?
a. aluminum (density = 1.70 g/cc)
b. lead (density = 11.34 g/cc)
c. Styrofoam (density = 0.100 g/cc)
d. maple wood (density = 0.74 g/cc)

1.87 (a) What is the specific gravity of mercury, the liquid used in thermometers, if it has a density of 13.6 g/mL? (b) What is the density of ethanol if it has a specific gravity of 0.789?

1.88 Why is specific gravity a unitless quantity?

General Questions

1.89 What are the advantages of using the metric system of measurement over the English system of measurement?

1.90 When you convert pounds to grams, how do you decide which unit of the conversion factor is located in the numerator?

1.91 Rank the quantities in each group from smallest to largest.
a. 100 μL, 100 dL, and 100 mL
b. 1 dL, 10 mL, and 1,000 μL
c. 10 g, 100 mg, and 0.1 kg
d. 1 km, 100 m, and 1,000 cm

1.92 What is the difference between mass and weight?

Applications

1.93 A lab test showed an individual's cholesterol level to be 186 mg/dL. (a) Convert this quantity to g/dL. (b) Convert this quantity to mg/L.

1.94 Hemoglobin is a protein that transports oxygen from the lungs to the rest of the body. Lab results indicated a patient had a hemoglobin concentration in the blood of 15.5 g/dL, which is in the normal range. (a) Convert the number of grams to milligrams and write the answer in scientific notation. (b) Convert the number of grams to micrograms and write the answer in scientific notation.

1.95 A woman was told to take a dose of 1.5 g of calcium daily. How many 500-mg tablets should she take?

1.96 The recommended daily calcium intake for a woman over 50 years of age is 1,200 mg. If one cup of milk has 306 mg of calcium, how many cups of milk provide this amount of calcium? (b) How many milliliters of milk does this correspond to?

1.97 A medium banana contains 451 mg of the nutrient potassium. How many bananas would you have to eat in one day to obtain the recommended daily intake of 3.5 g of potassium?

1.98 A single 1-oz serving of tortilla chips contains 250 mg of sodium. If an individual ate the entire 13-oz bag, how many grams of sodium would he ingest? If the recommended daily intake of sodium is 2.4 g, does this provide more or less than the recommended daily value, and by how much?

1.99 A bottle of liquid medication contains 300 mL and costs $10.00. (a) If the usual dose is 20. mL, how much does each dose cost? (b) If the usual dose is two tablespoons (1 tablespoon = 15 mL), how much does each dose cost?

1.100 The average nicotine content of a Camel cigarette is 1.93 mg. (a) Convert this quantity to both grams and micrograms. (b) Nicotine patches, which are used to help quit smoking, release nicotine into the body by absorption through the skin. The patches come with different amounts of nicotine. A smoker begins with the amount of nicotine that matches his typical daily intake. The maximum amount of nicotine in one brand of patch supplies a smoker with 21 mg of nicotine per day. If an individual smoked one pack of 20 Camel cigarettes each day, would a smoker get more or less nicotine per day using this patch?

1.101 A chemist synthesized 0.510 kg of aspirin in the lab. If the normal dose of aspirin is two 325-mg tablets, how many doses did she prepare?

1.102 Maalox is the trade name for an antacid and antigas medication used for relief of heartburn, bloating, and acid indigestion. Each 5-mL portion of Maalox contains 400 mg of aluminum hydroxide, 400 mg of magnesium hydroxide, and 40 mg of simethicone. If the recommended dose is two teaspoons four times a day, how many grams of each substance would an individual take in a 24-hour period. (1 teaspoon = 5 mL.)

1.103 Children's Chewable Tylenol contains 80 mg of acetaminophen per tablet. If the recommended dosage is 10 mg/kg, how many tablets are needed for a 42-lb child?

1.104 A patient is prescribed 2.0 g of a medication to be taken four times a day. If the medicine is available in 500-mg tablets, how many tablets are needed in a 24-hour period?

1.105 Children's Liquid Motrin contains 100. mg of the pain reliever ibuprofen per 5 mL. If the dose for a 45-lb child is 1.5 teaspoons, how many grams of ibuprofen would the child receive? (1 teaspoon = 5 mL.)

1.106 Often the specific amount of a drug to be administered must be calculated from a given dose in mg per kilogram of body weight. This assures that individuals who have very different body mass get the proper dose. If the proper dosage of a drug is 2 mg/kg of body weight, how many milligrams would a 110-lb individual need?

1.107 If the proper dose of a medication is 10 μg/kg of body weight, how many milligrams would a 200-lb individual need?

1.108 If a 180-lb patient is prescribed 20 mg of the cholesterol-lowering drug Lipitor daily, what dosage is the patient receiving in mg/kg of his body weight?