CHAPTER OUTLINE

CHAPTER TWO

Measurements in Chemistry

Measurements can never be exact; there is always some degree of uncertainty.

It would be extremely difficult for a carpenter to build cabinets without being able to use hammers, saws, and drills. They are the tools of a carpenter's trade. Chemists also have "tools of the trade." The tool they use most is called *measurement*. Understanding measurement is indispensable in the study of chemistry. Questions such as "How much … ?," "How long … ?," and "How many … ?" simply cannot be answered without resorting to measurements. This chapter will help you learn what you need to know to deal properly with measurement. Much of the material in the chapter is mathematical. This is necessary; measurements require the use of numbers.

2.1 Measurement Systems

Learning Focus

Understand why scientists prefer the metric system of units over the English system of units when making measurements.

We all make measurements on a routine basis. For example, measurements are involved in following a recipe for making brownies, in determining our height and weight, and in fueling a car with gasoline. **Measurement** *is the determination of the dimensions, capacity, quantity, or extent of something.* In chemical laboratories, the most common types of measurements are those of mass, volume, length, time, temperature, pressure, and concentration.

Two systems of measurement are in use in the United States at present: (1) the English system of units, and (2) the metric system of units. Common measurements of commerce, such as those used in a grocery store, are made in the *English system*. The units of this system include the inch, foot, pound, quart, and gallon. The *metric system* is used in scientific work. The units of this system include the gram, meter, and liter.

▸ The word *metric* is derived from the Greek word *metron,* which means "measure."

Figure 2.1 Metric system units are becoming increasingly evident on highway signs and consumer products.

The United States is in the process of voluntary conversion to the metric system fo measurements of commerce. Metric system units now appear on numerous consumer products (Figure 2.1). Soft drinks now come in 1-, 2- and 3-liter containers. Road signs in some states display distances in both miles and kilometers. Canned and packaged goods such as cereals and mixes on grocery store shelves now have the masses of their contents listed in grams as well as in pounds and ounces.

Interrelationships between units of the same type, such as volume or length, are less complicated in the metric system than in the English system. Within the metric system conversion from one unit size to another can be accomplished simply by multiplying o dividing by units of 10, because the metric system is a decimal unit system—that is, it i based on multiples of 10. The metric system is simply more convenient to use.

Practice Questions and Problems

2.1 List the most common types of measurements made in chemical laboratories.

2.2 What is the main reason why scientists prefer to use the metric measurement syster instead of the English measurement system?

2.2 Metric System Units

Learning Focus

Recognize units of the metric system by name and abbreviation, and know the numerical meanings associated with various metric system prefixes.

In the metric system, there is one base unit for each type of measurement (length, mas volume, and so on). The names of fractional parts of the base unit and multiples of th base unit are constructed by adding prefixes to the base unit. These prefixes indicate th size of the unit relative to the base unit. Table 2.1 lists common metric system prefixe along with their symbols or abbreviations and mathematical meanings. The prefixes i color are the ones most frequently used.

The meaning of a metric system prefix is independent of the base unit it modifies ar always remains constant. For example, the prefix *kilo-* always means 1000; a *kilo*secor is 1000 seconds, a *kilo*watt is 1000 watts, and a *kilo*calorie is 1000 calories. Similarly, th prefix *nano-* always means one-billionth; a *nano*meter is one-billionth of a meter, *nano*gram is one-billionth of a gram, and a *nano*liter is one-billionth of a liter.

▶ The use of numerical prefixes should not be new to you. Consider the use of the prefix *tri-* in the words *tri*angle, *tri*cycle, *trio*, *tri*nity, and *tri*ple. Each of these words conveys the idea of three of something. The metric system prefixes are used in the same way.

Metric Length Units

The **meter** (m) *is the base unit of length in the metric system.* It is about the same si as the English yard; 1 meter equals 1.09 yards (Figure 2.2a). The prefixes listed in Tab 2.1 enable us to derive other units of length from the meter. The kilometer (km) is 10 times larger than the meter; the centimeter (cm) and millimeter (mm) are, respective one-hundredth and one-thousandth of a meter. Most laboratory length measurements a made in centimeters rather than meters because of the meter's relatively large size.

▶ *Length* is measured by determining the distance between two points.

Table 2.1
Common Metric System Prefixes with Their Symbols and Mathematical Meanings

	Prefix[a]	Symbol	Mathematical Meaning[b]
Multiples	giga-	G	1,000,000,000 (10^9, billion)
	mega-	M	1,000,000 (10^6, million)
	kilo-	k	1000 (10^3, thousand)
Fractional parts	deci-	d	0.1 (10^{-1}, one-tenth)
	centi-	c	0.01 (10^{-2}, one-hundredth)
	milli-	m	0.001 (10^{-3}, one-thousandth)
	micro-	μ (Greek mu)	0.000001 (10^{-6}, one-millionth)
	nano-	n	0.000000001 (10^{-9}, one-billionth)
	pico-	p	0.000000000001 (10^{-12}, one-trillionth)

[a] Other prefixes also are available but are less commonly used.
[b] The power-of-10 notation for denoting numbers is considered in Section 2.5.

Metric Mass Units

▶ *Mass* is measured by determining the amount of matter in an object.

The **gram** (g) *is the base unit of mass in the metric system.* It is a very small unit compared with the English ounce and pound (Figure 2.2b). It takes approximately 28 grams to equal 1 ounce and nearly 454 grams to equal 1 pound. Both grams and milligrams (mg) are commonly used in the laboratory, where the kilogram (kg) is generally too large.

The terms *mass* and *weight* are often used interchangeably in measurement discussions; technically, however, they have different meanings. **Mass** *is a measure of the total quantity of matter in an object.* **Weight** *is a measure of the force exerted on an object by the pull of gravity.*

▶ Students often erroneously think that the terms *mass* and *weight* have the same meaning. *Mass* is a measure of the amount of material present in a sample. *Weight* is a measure of the force exerted on an object by the pull of gravity.

The mass of a substance is a constant; the weight of an object varies with the object's geographical location. For example, matter at the equator weighs less than it would at the North Pole because the pull of gravity is less at the equator. Because Earth is not a perfect sphere, but bulges at the equator, the magnitude of gravitational attraction is less at the equator. An object would weigh less on the moon than on Earth because of the smaller size of the moon and the correspondingly lower gravitational attraction. Quantitatively, a 22.0-lb mass weighing 22.0 lb at Earth's North Pole would weigh 21.9 lb at Earth's equator and only 3.7 lb on the moon. In outer space, an astronaut may be weightless but never massless. In fact, he or she has the same mass in space as on Earth.

(a) Length	(b) Mass	(c) Volume
A meter is slightly larger than a yard. 1 meter = 1.09 yards. A baseball bat is about 1 meter long.	A gram is a small unit compared to a pound. 1 gram = 1/454 pound. Two pennies, five paperclips, and a marble have masses of about 5, 2, and 5 grams, respectively.	A liter is slightly larger than a quart. 1 liter = 1.06 quarts. Most beverages are now sold by the liter rather than by the quart.

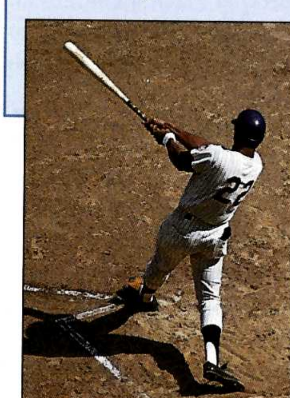


Figure 2.2 Comparisons of the base metric system units of length (meter), mass (gram), and volume (liter) with common objects.

Metric Volume Units

▶ Volume is measured by determining the amount of space occupied by a three-dimensional object.

The **liter** (L) *is the base unit of volume in the metric system.* The abbreviation for liter is a capital L rather than a lower-case l because a lower-case l is easily confused with the number 1. A liter is a volume equal to that occupied by a cube that is 10 centimeters on each side. Because the volume of a cube is calculated by multiplying length times width times height (which are all the same for a cube), we have

$$\begin{aligned} 1 \text{ liter} &= \text{volume of a cube with edge } 10 \text{ cm} \\ &= 10 \text{ cm} \times 10 \text{ cm} \times 10 \text{ cm} \\ &= 1000 \text{ cm}^3 \end{aligned}$$

A liter is also equal to 1000 milliliters; the prefix *milli-* means one-thousandth. Therefore,

$$1000 \text{ mL} = 1000 \text{ cm}^3$$

Dividing both sides of this equation by 1000 reveals that

$$1 \text{ mL} = 1 \text{ cm}^3$$

▶ Another abbreviation for the unit cubic centimeter, used in medical situations, is cc.

$$1 \text{ cm}^3 = 1 \text{ cc}$$

Consequently, the units mL and cm^3 are the same. In practice, mL is used for volumes of liquids and gases, cm^3 for volumes of solids. Figure 2.3 shows the relationship between 1 mL (1 cm^3) and its parent unit, the liter, in terms of cubic measurements.

A liter and a quart have approximately the same volume; 1 liter equals 1.06 quarts (Figure 2.2c). The milliliter and deciliter (dL) are commonly used in the laboratory. Deciliter units are routinely encountered in clinical laboratory reports detailing the composition of body fluids. A deciliter is equal to 100 mL (0.100 L).

Example 2.1 **Mass Units and Measurement**

Complete the following table.

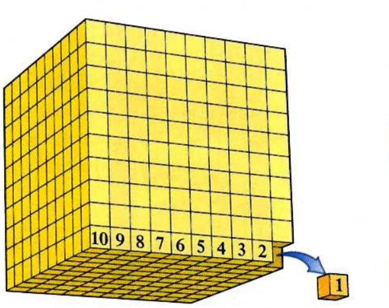
Type of measurement	Name of metric unit	Metric unit abbreviation
mass	________	kg
________	meter	________
________	cubic centimeter	________
________	________	mL

Solution

mass	kilogram	kg
length	meter	m
volume	cubic centimeter	cm^3
volume	milliliter	mL

Figure 2.3 A cube 10 cm on a side has a volume of 1000 cm^3, which is equal to 1 L. A cube 1 cm on a side has a volume of 1 cm^3, which is equal to 1 mL.

Total volume of large cube = 1000 cm^3 = 1 L



1 cm^3 = 1 mL

Practice Questions and Problems

2.3 Write the name of the metric system prefix associated with each of the following mathematical meanings.
a. 10^3 b. 10^{-3} c. 10^{-6} d. 1/10

2.4 Provide the full unit name for each of the following abbreviations, and indicate what is being measured (time, mass, etc.)
a. cm b. kL c. mg d. ng

2.5 Arrange each of the following sets of units in order of increasing size (from smallest to largest).
a. Milligram, centigram, nanogram
b. Gigameter, megameter, kilometer
c. Microliter, deciliter, picoliter
d. Milligram, kilogram, microgram

Learning Focus

Classify a number as exact or inexact on the basis of the context of its use.

2.3 Exact and Inexact Numbers

In scientific work, numbers are grouped in two categories: *exact numbers* and *inexact numbers*. An **exact number** *has a value that has no uncertainty associated with it—that is, it is known exactly.* Exact numbers occur in definitions (for example, there are exactly 12 objects in a dozen, not 12.01 or 12.02); in counting (for example, there can be 7 people in a room, but never 6.99 or 7.03); and in simple fractions (for example, 1/3, 3/5, or 5/9).

An **inexact number** *has a value that has a degree of uncertainty associated with it.* Inexact numbers result anytime a measurement is made. It is impossible to make an *exact* measurement; some uncertainty will always be present. Flaws in construction of a measuring device, improper calibration of an instrument, and the skill (or lack of skill) of a person using a measuring device all contribute to error (uncertainty).

Practice Questions and Problems

2.6 Indicate whether the number in each of the following statements is an exact or an inexact number.
a. A classroom contains 32 chairs.
b. There are 60 seconds in a minute.
c. A bowl of cherries weighs 3.2 pounds.
d. A newspaper article contains 323 words.

2.7 Indicate whether each of the following quantities would involve an exact number or an inexact number.
a. The length of a swimming pool
b. The number of gummi bears in a bag
c. The number of inches in a foot
d. The surface area of a living room rug

2.8 A person is told that there are 12 inches in a foot and also that a piece of rope is 12 inches long. What is the fundamental difference between the value of 12 in these two pieces of information?

Learning Focus

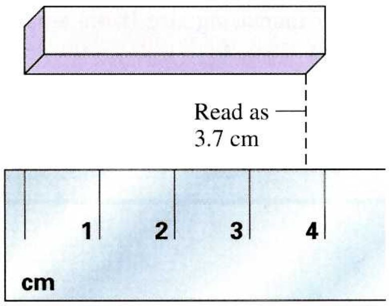
Know the fundamental rule for recording measurements, and be able to determine the number of significant figures in a given measurement.

2.4 Uncertainty in Measurement and Significant Figures

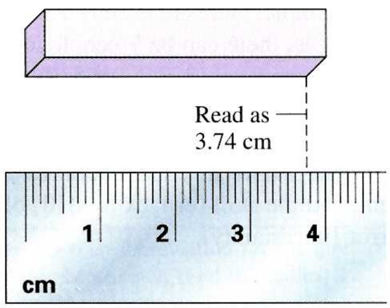
As noted in the previous section, because of the limitations of the measuring device and the limited powers of observation of the individual making the measurement, every measurement carries a degree of uncertainty or error. Even when very elaborate measuring devices are used, some degree of uncertainty is always present.

Origin of Measurement Uncertainty

To illustrate how measurement uncertainty arises, let us consider how two different rulers, shown in Figure 2.4, are used to measure a given length. Using ruler A, we can say with certainty that the length of the rod is between 3 and 4 centimeters. We can further say that the actual length is closer to 4 centimeters and estimate it to be 3.7 centimeters. Ruler B has more subdivisions on its scale than ruler A. It is marked off in tenths of a centimeter instead of in centimeters. Using ruler B, we can definitely say that the length of the rod is between 3.7 and 3.8 centimeters and can estimate it to be 3.74 centimeters.



Ruler A



Ruler B

Figure 2.4 The scale on a measuring device determines the magnitude of the uncertainty for the recorded measurement. Measurements made with ruler A will have greater uncertainty than those made with ruler B.

Note how both length measurements (ruler A and ruler B) contain some digits (all those except the last one) that are exactly known and one digit (the last one) that is estimated. It is this last digit, the estimated one, that produces uncertainty in a measurement. Note also that the uncertainty in the second length measurement is less than that in the first one—an uncertainty in the hundredths place compared with an uncertainty in the tenths place. We say that the second measurement is more *precise* than the first one; that is, it has less uncertainty than the first measurement.

Only one estimated digit is ever recorded as part of a measurement. It would be incorrect for a scientist to report that the length of the metal rod in Figure 2.4 is 3.745 centimeters as read by using ruler B. The value 3.745 contains two estimated digits, the 4 and the 5, and indicates a measurement of precision greater than what is actually obtainable with that particular measuring device. Again, only one estimated digit is ever recorded as part of a measurement.

Because measurements are never exact, two types of information must be conveyed whenever a numerical value for a measurement is recorded: (1) the magnitude of the measurement and (2) the uncertainty of the measurement. The magnitude is indicated by the digit values. Uncertainty is indicated by the number of significant figures recorded. **Significant figures** *are the digits in any measurement that are known with certainty plus one digit that is uncertain.*

▶ The term *significant figures* is often verbalized in shortened form as "sig figs."

Guidelines for Determining Significant Figures

Recognizing the number of significant figures in a measured quantity is easy for measurements we make ourselves, because we know the type of instrument we are using and its limitations. However, when someone else makes the measurement, such information is often not available. In such cases, we follow a set of guidelines for determining the number of significant figures in a measured quantity.

1. In any measurement, all nonzero digits are significant.
2. *Zeros* may or may not be significant because zeros can be used in two ways: (1) to position a decimal point, and (2) to indicate a measured value. Zeros that perform the first function are not significant, and zeros that perform the second function are significant. When zeros are present in a measured number, we follow these rules:
   a. *Leading zeros,* those at the beginning of a number, are never significant.

      0.0141 has three significant figures.
      0.0000000048 has two significant figures.

   b. *Confined zeros,* those between nonzero digits, are always significant.

      3.063 has four significant figures.
      0.001004 has four significant figures.

   c. *Trailing zeros,* those at the end of a number, are significant if a decimal point is present in the number.

      56.00 has four significant figures.
      0.05050 has four significant figures.

   d. *Trailing zeros,* those at the end of a number, are not significant if the number lacks an explicitly shown decimal point.

      59,000,000 has two significant figures.
      6010 has three significant figures.

Practice Questions and Problems

2.9 Why are measured numbers restricted to a specific number of significant figures?

2.10 Indicate to what decimal position readings should be recorded (nearest 0.1, 0.01, etc.) for measurements made with the following devices.

a. A thermometer with a smallest scale marking of 1°C
b. A graduated cylinder with a smallest scale marking of 0.1 mL
c. A volumetric device with a smallest scale marking of 10 mL
d. A ruler with a smallest scale marking of 1 mm

2.11 Which is the estimated digit in each of the following numbers?
a. 2.31 cm b. 1.0 mL c. 25°C d. 1.23 g

2.12 Determine the number of significant figures in each of the following measured values.
a. 6.000 b. 0.0032 c. 0.01001 d. 65,400 e. 766.010 f. 0.03050

2.13 In which of the following pairs of numbers do both members of the pair contain the same number of significant figures?
a. 11.01 and 11.00 b. 2002 and 2020
c. 0.000066 and 660,000 d. 0.05700 and 0.05070

Learning Focus

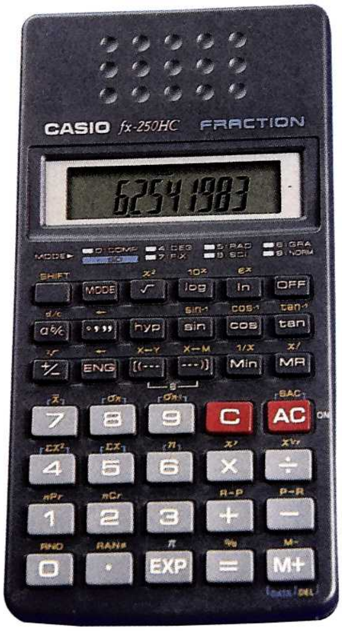
Be able to adjust calculated answers obtained using measurements to the correct number of significant figures.

2.5 Significant Figures and Mathematical Operations

When measurements are added, subtracted, multiplied, or divided, consideration must be given to the number of significant figures in the computed result. Mathematical operations should not increase (or decrease) the precision of experimental measurements.

Hand-held electronic calculators generally "complicate" preciseness considerations because they are not programmed to take significant figures into account. Consequently, the digital readouts display more digits than are warranted (Figure 2.5). It is a mistake to record these extra digits, because they are not significant figures and hence are meaningless.

Figure 2.5 The digital readout on an electronic calculator usually shows more digits than are needed—and more than are acceptable. Calculators are not programmed to account for significant figures.



Rounding Off Numbers

When we obtain calculator answers that contain too many digits, it is necessary to drop the nonsignificant digits, a process that is called rounding off. **Rounding off** *is the process of deleting unwanted (nonsignificant) digits from calculated numbers*. There are two rules for rounding off numbers.

1. *If the first digit to be deleted is 4 or less, simply drop it and all the following digits*. For example, the number 3.724567 becomes 3.72 when rounded to three significant figures.
2. *If the first digit to be deleted is 5 or greater, that digit and all that follow are dropped and the last retained digit is increased by one*. The number 5.00673 becomes 5.01 when rounded to three significant figures.

Operational Rules

Significant-figure considerations in mathematical operations that involve measured numbers are governed by two rules, one for multiplication and division and one for addition and subtraction.

1. *In multiplication and division, the number of significant figures in the answer is the same as the number of significant figures in the measurement that contains the fewest significant figures*. For example,

Four significant figures ↓ Three significant figures ↓

$$6.038 \times 2.57 = 15.51766 \quad \text{(calculator answer)}$$
$$= 15.5 \quad \text{(correct answer)}$$

↑ Three significant figures

The calculator answer is rounded to three significant figures because the measurement with the fewest significant figures (2.57) contains only three significant figures.

2. *In addition and subtraction, the answer has no more digits to the right of the decimal point than are found in the measurement with the fewest digits to the right of the decimal point.* For example,

$$\begin{array}{rl} 9.333 & \leftarrow \text{Uncertain digit (thousandths)} \\ +\ 1.4 & \leftarrow \text{Uncertain digit (tenths)} \\ \hline 10.733 & \text{(calculator answer)} \\ 10.7 & \text{(correct answer)} \end{array}$$

↑ Uncertain digit (tenths)

The calculator answer is rounded to the tenths place because the uncertainty in the number 1.4 is in the tenths place.

▶ Concisely stated, the significant-figure operational rules are
× or ÷: Keep smallest number of significant figures in answer.
+ or −: Keep smallest number of decimal places in answer.

Note the contrast between the rule for multiplication and division and the rule for addition and subtraction. In multiplication and division, significant figures are counted; in addition and subtraction, decimal places are counted. It is possible to gain or lose significant figures during addition or subtraction, but *never* during multiplication or division. In our previous sample addition problem, one of the input numbers (1.4) has two significant figures and the correct answer (10.7) has three significant figures. This is allowable in addition (and subtraction) because we are counting decimal places, not significant figures.

Example 2.2 **Expressing Answers to the Proper Number of Significant Figures**

Perform the following computations, expressing your answers to the proper number of significant figures.

a. 6.7321×0.0021 **b.** $\frac{16{,}340}{23.42}$ **c.** 6.000×4.000

d. $8.3 + 1.2 + 1.7$ **e.** $3.07 \times (17.6 - 13.73)$

Solution

a. The calculator answer to this problem is

$$6.7321 \times 0.0021 = 0.01413741$$

The input number with the least number of significant figures is 0.0021.

$$6.7321 \times 0.0021$$

↗ Five significant figures ↖ Two significant figures

Thus the calculator answer must be rounded to two significant figures.

0.01413741 (Calculator answer) becomes 0.014 (Correct answer)

b. The calculator answer to this problem is

$$\frac{16{,}340}{23.42} = 697.69427$$

Both input numbers contain four significant figures. Thus the correct answer will also contain four significant figures.

697.69427 (Calculator answer) becomes 697.7 (Correct answer)

c. The calculator answer to this problem is

$$6.000 \times 4.000 = 24$$

Both input numbers contain four significant figures. Thus the correct answer must also contain four significant figures:

24	becomes	24.00
Calculator answer		Correct answer

Note that here the calculator answer had too few significant figures. Most calculators cut off zeros after the decimal point even if these zeros are significant. Using too few significant figures in an answer is just as wrong as using too many.

d. The calculator answer to this problem is

$$8.3 + 1.2 + 1.7 = 11.2$$

All three input numbers have uncertainty in the tenths place. Thus the last retained digit in the correct answer will be that of tenths. (In this particular problem, the calculator answer and the correct answer are the same, a situation that does not occur very often.)

e. This problem involves the use of both multiplication and subtraction significant-figure rules. We do the subtraction first.

$$17.6 - 13.73 = 3.87 \quad \text{(calculator answer)}$$
$$= 3.9 \quad \text{(correct answer)}$$

This answer must be rounded to tenths because the input number 17.6 involves only tenths. We now do the multiplication.

$$3.07 \times 3.9 = 11.973 \quad \text{(calculator answer)}$$
$$= 12 \quad \text{(correct answer)}$$

The number 3.9 limits the answer to two significant figures.

Because exact numbers (Section 2.3) have no uncertainty associated with them, they possess an unlimited number of significant figures. Therefore, such numbers never limit the number of significant figures in a computational answer.

The Chemistry at a Glance on the following page summarizes the rules that govern which digits in a measurement are significant.

Practice Questions and Problems

2.14 Round off each of the following numbers to the number of significant figures indicated in parentheses.

a. 0.350763 (three) b. 653,899 (four)
c. 22.55555 (five) d. 0.277654 (four)

2.15 Without actually solving the problem, indicate the number of significant figures that should be present in the answers to the following multiplication and division problems.

a. $10.300 \times 0.30 \times 0.300$ b. $3300 \times 3330 \times 333.0$ c. $\dfrac{6.0}{33.0}$ d. $\dfrac{6.000}{33}$

2.16 Carry out the following multiplications and divisions, expressing your answer to the correct number of significant figures. Assume that all numbers are measured numbers.

a. $2.0000 \times 2.00 \times 0.0020$ b. 4.1567×0.00345

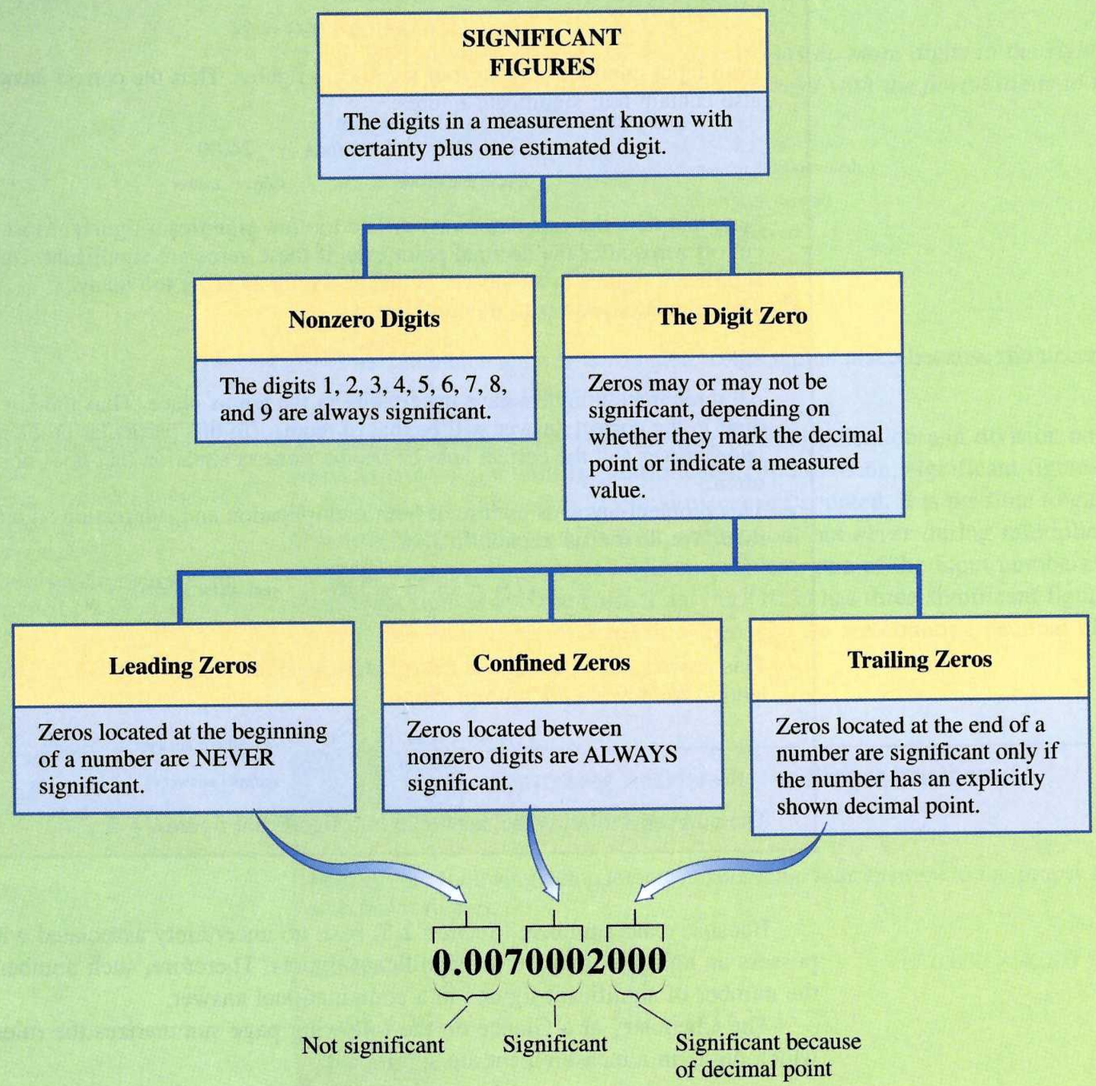
c. $\dfrac{533{,}000}{465{,}300}$ d. $\dfrac{4.670 \times 3.00}{2.450}$

2.17 Carry out the following additions and subtractions, expressing your answer to the correct number of significant figures.

a. $12 + 23 + 127$ b. $3.111 + 3.11 + 3.1$
c. $1237.6 + 23 + 0.12$ d. $43.65 - 23.7$

Chemistry *at a Glance*

Significant Figures



Learning Focus

Be able to convert numbers from decimal notation to scientific notation.

2.6 Scientific Notation

Up to this point in the chapter, we have expressed all numbers in decimal notation, the everyday method for expressing numbers. Such notation becomes cumbersome for very large and very small numbers (which occur frequently in scientific work). For example, in one drop of blood, which is 92% water by mass, there are approximately

1,600,000,000,000,000,000,000 molecules

of water, each of which has a mass of

0.000000000000000000000030 gram

Recording such large and small numbers is not only time-consuming but also open to error; often, too many or too few zeros are recorded. Also, it is impossible to multiply or divide such numbers with most calculators because they can't accept that many digits. (Most calculators accept either 8 or 10 digits.)

▶ Scientific notation is also called exponential notation.

A method called *scientific notation* exists for expressing multidigit numbers involving many zeros in compact form. **Scientific notation** *is a system in which an ordinary decimal number is expressed as the product of a number between 1 and 10 and 10 raised to a power.* The ordinary decimal number is called a *coefficient* and is written first. The number 10 raised to a power is called an *exponential term.* The coefficient is always multiplied by the exponential term.

Coefficient — Exponent

$$1.07 \times 10^{4}$$

Multiplication sign — Exponential term

The two previously cited numbers that deal with molecules of water are expressed in scientific notation as

$$1.6 \times 10^{21} \text{ molecules}$$

and

$$3.0 \times 10^{-22} \text{ gram}$$

Obviously, scientific notation is a much more concise way of expressing numbers. Such scientific notation is compatible with most calculators.

Converting from Decimal to Scientific Notation

The procedure for converting a number from decimal notation to scientific notation has two parts.

1. *The decimal point in the decimal number is moved to the position behind the first nonzero digit.*
2. *The exponent for the exponential term is equal to the number of places the decimal point has been moved.* The exponent is positive if the original decimal number is 10 or greater and is negative if the original decimal number is less than 1. For numbers between 1 and 10, the exponent is zero.

The following two examples illustrate the use of these procedures:

$$93{,}000{,}000 = 9.3 \times 10^{7}$$

Decimal point is moved 7 places

$$0.0000037 = 3.7 \times 10^{-6}$$

Decimal point is moved 6 places

Significant Figures and Scientific Notation

How do significant-figure considerations affect scientific notation? The answer is simple. *Only significant figures become part of the coefficient.* The numbers 63, 63.0, and 63.00, which respectively have two, three, and four significant figures, when converted to scientific notation become, respectively,

▶ The decimal and scientific notation forms of a number *always* contain the same number of significant figures.

6.3×10^{1} (two significant figures)

6.30×10^{1} (three significant figures)

6.300×10^{1} (four significant figures)

Practice Questions and Problems

2.18 Express the following numbers in scientific notation.
a. 120.7 b. 0.0034 c. 231.00 d. 23,000

2.19 Which number in each pair of numbers is the larger of the two?
a. 1.0×10^{-3} or 1.0×10^{-6} b. 1.0×10^{3} or 1.0×10^{-2}
c. 6.3×10^{4} or 2.3×10^{4} d. 6.3×10^{-4} or 1.2×10^{-4}

2.20 How many significant figures are present in each of the following measured numbers?
a. 1.0×10^{2} b. 5.34×10^{5} c. 5.34×10^{-4} d. 6.000×10^{3}

Learning Focus

Understand what conversion factors are, and be able to use them and dimensional analysis to change from one unit to another.

2.7 Conversion Factors and Dimensional Analysis

With both the English unit and metric unit systems in common use in the United States, we often must change measurements from one system to their equivalent in the other system. The mathematical tool we use to accomplish this task is a general method of problem solving called *dimensional analysis*. Central to the use of dimensional analysis is the concept of conversion factors. A **conversion factor** *is a ratio that specifies how one unit of measurement is related to another.*

Conversion factors are derived from equations (equalities) that relate units. Consider the quantities "1 minute" and "60 seconds," both of which describe the same amount of time. We may write an equation describing this fact.

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

This fixed relationship is the basis for the construction of a pair of conversion factors that relate seconds and minutes.

$$\frac{1 \text{ min}}{60 \text{ sec}} \quad \text{and} \quad \frac{60 \text{ sec}}{1 \text{ min}}$$

These two quantities are the same.

Note that conversion factors always come in pairs, one member of the pair being the reciprocal of the other. Also note that the numerator and the denominator of a conversion factor always describe the same amount of whatever we are considering. One minute and 60 seconds denote the same amount of time.

Conversion Factors Within a System of Units

Most students are familiar with and have memorized numerous conversion factors within the English system of measurement (English-to-English conversion factors). Some of these factors, with only one member of a conversion factor pair being listed, are

▶ In order to avoid confusion with the word *in*, the abbreviation for inches, in., includes a period. This is the only unit abbreviation in which a period appears.

$$\frac{12 \text{ in.}}{1 \text{ ft}} \quad \frac{3 \text{ ft}}{1 \text{ yd}} \quad \frac{4 \text{ qt}}{1 \text{ gal}} \quad \frac{16 \text{ oz}}{1 \text{ lb}}$$

Such conversion factors contain an unlimited number of significant figures because the numbers within them arise from definitions.

Metric-to-metric conversion factors are similar to English-to-English conversion factors in that they arise from definitions. Individual conversion factors are derived from the meanings of the metric system prefixes (Section 2.2). For example, the set of conversion factors involving kilometer and meter come from the equality

$$1 \text{ kilometer} = 10^3 \text{ meters}$$

and those relating microgram and gram come from the equality

$$1 \text{ microgram} = 10^{-6} \text{ gram}$$

The two pairs of conversion factors are

▶ In order to obtain metric-to-metric conversion factors, you need to know the meaning of the metric system prefixes in terms of powers of 10 (see Table 2.1).

$$\frac{10^3 \text{ m}}{1 \text{ km}} \quad \text{and} \quad \frac{1 \text{ km}}{10^3 \text{ m}} \qquad \frac{1 \ \mu\text{g}}{10^{-6} \text{ g}} \quad \text{and} \quad \frac{10^{-6} \text{ g}}{1 \ \mu\text{g}}$$

Note that the numerical equivalent of the prefix is always associated with the base (unprefixed) unit in a metric-to-metric conversion factor.

The number 1 always goes with the *prefixed* unit.

$$\frac{1 \text{ mL}}{10^{-3} \text{ L}}$$

The power of 10 always goes with the *unprefixed* unit.

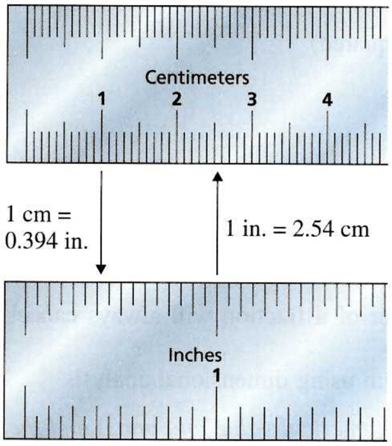


Figure 2.6 It is experimentally determined that 1 inch equals 2.54 centimeters or 1 centimeter equals 0.394 inch.

Conversion Factors Between Systems of Units

Conversion factors that relate metric units to English units and vice versa are not exact defined quantities because they involve two different systems of measurement. The numbers associated with these conversion factors must be determined experimentally (see Figure 2.6). Table 2.2 lists commonly encountered relationships between metric system and English system units. These few conversion factors are sufficient to solve most of the problems that we will encounter.

Metric-to-English conversion factors can be specified to differing numbers of significant figures. For example,

$$1.00\ \text{lb} = 454\ \text{g}$$
$$1.000\ \text{lb} = 453.6\ \text{g}$$
$$1.0000\ \text{lb} = 453.59\ \text{g}$$

In a problem-solving context, which "version" of a conversion factor is used depends on how many significant figures there are in the other numbers of the problem. Conversion factors should never limit the number of significant figures in the answer to a problem. The conversion factors in Table 2.2 are given to three significant figures, which is sufficient for the applications we will make of them.

Dimensional Analysis

Dimensional analysis *is a general problem-solving method in which the units associated with numbers are used as a guide in setting up calculations.* In this method, units are treated in the same way as numbers; that is, they can be multiplied, divided, or canceled. For example, just as

$$5 \times 5 = 5^2 \qquad \text{(5 squared)}$$

Table 2.2
Equalities and Conversion Factors That Relate the English and Metric Systems of Measurement

	Metric to English	English to Metric
Length		
1.00 inch = 2.54 centimeters	$\frac{1.00\ \text{in.}}{2.54\ \text{cm}}$	$\frac{2.54\ \text{cm}}{1.00\ \text{in.}}$
1.00 meter = 39.4 inches	$\frac{39.4\ \text{in.}}{1.00\ \text{m}}$	$\frac{1.00\ \text{m}}{39.4\ \text{in.}}$
1.00 kilometer = 0.621 mile	$\frac{0.621\ \text{mi}}{1.00\ \text{km}}$	$\frac{1.00\ \text{km}}{0.621\ \text{mi}}$
Mass		
1.00 pound = 454 grams	$\frac{1.00\ \text{lb}}{454\ \text{g}}$	$\frac{454\ \text{g}}{1.00\ \text{lb}}$
1.00 kilogram = 2.20 pounds	$\frac{2.20\ \text{lb}}{1.00\ \text{kg}}$	$\frac{1.00\ \text{kg}}{2.20\ \text{lb}}$
1.00 ounce = 28.3 grams	$\frac{1.00\ \text{oz}}{28.3\ \text{g}}$	$\frac{28.3\ \text{g}}{1.00\ \text{oz}}$
Volume		
1.00 quart = 0.946 liter	$\frac{1.00\ \text{qt}}{0.946\ \text{L}}$	$\frac{0.946\ \text{L}}{1.00\ \text{qt}}$
1.00 liter = 0.265 gallon	$\frac{0.265\ \text{gal}}{1.00\ \text{L}}$	$\frac{1.00\ \text{L}}{0.265\ \text{gal}}$
1.00 milliliter = 0.034 fluid ounce	$\frac{0.034\ \text{fl oz}}{1.00\ \text{mL}}$	$\frac{1.00\ \text{mL}}{0.034\ \text{fl oz}}$

we have

$$\text{cm} \times \text{cm} = \text{cm}^2 \quad \text{(cm squared)}$$

Also, just as the 3s cancel in the expression

$$\frac{\cancel{3} \times 5 \times 7}{\cancel{3} \times 2}$$

the centimeters cancel in the expression

$$\frac{(\cancel{\text{cm}}) \times (\text{in.})}{(\cancel{\text{cm}})}$$

"Like units" found in the numerator and denominator of a fraction will always cancel, just as like numbers do.

The following steps show how to set up a problem using dimensional analysis.

Step 1: *Identify the known or given quantity (both numerical value and units) and the units of the new quantity to be determined.*

This information will always be found in the statement of the problem. Write an equation with the given quantity on the left and the units of the desired quantity on the right.

Step 2: *Multiply the given quantity by one or more conversion factors in such a manner that the unwanted (original) units are canceled, leaving only the desired units.*

The general format for the multiplication is

$$\text{(Information given)} \times \text{(conversion factors)} = \text{(information sought)}$$

The number of conversion factors depends on the individual problem.

Step 3: *Perform the mathematical operations indicated by the conversion factor setup.*

When performing the calculation, double-check to make sure that all units except the desired set have canceled.

Example 2.3 Unit Conversions Within the Metric System

A standard aspirin tablet contains 324 mg of aspirin. How many grams of aspirin are in a standard aspirin tablet?

Solution

Step 1: The given quantity is 324 mg, the mass of aspirin in the tablet. The unit of the desired quantity is grams.

$$324 \text{ mg} = ?\text{ g}$$

Step 2: Only one conversion factor will be needed to convert from milligrams to grams, one that relates milligrams to grams. The two forms of this conversion factor are

$$\frac{1 \text{ mg}}{10^{-3} \text{ g}} \quad \text{and} \quad \frac{10^{-3} \text{ g}}{1 \text{ mg}}$$

The second factor is used because it allows for cancellation of the milligram units, leaving us with grams as the new units.

$$324 \cancel{\text{mg}} \times \left(\frac{10^{-3} \text{ g}}{1 \cancel{\text{mg}}} \right) = ?\text{ g}$$

Step 3: Combining numerical terms as indicated generates the final answer.

$$\left(324 \times \frac{10^{-3}}{1}\right) \text{g} = 0.324 \text{ g}$$

Number from first factor — Numbers from second factor

Note that 10^{-3} is equal to 0.001 and that 324 times 0.001 is equal to 0.324.

The answer is given to three significant figures because the given quantity in the problem, 324 mg, has three significant figures. The conversion factor used arises from a definition and thus does not limit significant figures in any way.

Example 2.4 Unit Conversions Between the Metric and English Systems

Capillaries, the microscopic vessels that carry blood from small arteries to small veins, are on the average only 1 mm long. What is the average length of a capillary in inches?

Solution

Step 1: The given quantity is 1 mm, and the units of the desired quantity are inches.

$$1 \text{ mm} = ? \text{ in.}$$

Step 2: The conversion factor needed for a one-step solution, millimeters to inches, is not given in Table 2.2. However, a related conversion factor, meters to inches, is given. Therefore, we first convert millimeters to meters and then use the meters-to-inches conversion factor in Table 2.2.

$$\text{mm} \longrightarrow \text{m} \longrightarrow \text{in.}$$

The correct conversion factor setup is

$$1 \cancel{\text{mm}} \times \left(\frac{10^{-3} \cancel{\text{m}}}{1 \cancel{\text{mm}}}\right) \times \left(\frac{39.4 \text{ in.}}{1.00 \cancel{\text{m}}}\right) = ? \text{ in.}$$

All of the units except for inches cancel, which is what is needed. The information for the middle conversion factor was obtained from the meaning of the prefix *milli-*.

This setup illustrates the fact that sometimes the given units must be changed to intermediate units before common conversion factors, such as those found in Table 2.2, are applicable.

Step 3: Collecting the numerical factors and performing the indicated math gives

$$\left(\frac{1 \times 10^{-3} \times 39.4}{1 \times 1.00}\right) \text{in.} = 0.0394 \text{ in.} \quad \text{(calculator answer)}$$

$$= 0.04 \text{ in.} \quad \text{(correct answer)}$$

The calculator answer must be rounded to one significant figure because 1 mm, the given quantity, contains only one significant figure.

The next Chemistry at a Glance reviews what we have learned about conversion factors.

Practice Questions and Problems

2.21 Give both forms of the conversion factor that you would use to relate the following sets of units to each other.

a. Gram and kilogram b. Meter and nanometer
c. Liter and milliliter d. Centigram and gram

Chemistry *at a Glance*

Conversion Factors

Characteristics of Conversion Factors

- Ratios that specify how units are related to each other
- Derived from equations that relate units

  1 minute = 60 seconds
- Come in pairs, one member of the pair being the reciprocal of the other

  $$\frac{1\ \text{min}}{60\ \text{sec}} \quad \text{and} \quad \frac{60\ \text{sec}}{1\ \text{min}}$$
- Conversion factors originate from two types of relationships:

  (1) defined relationships and (2) measured relationships

Conversion Factors from DEFINED Relationships

- All English-to-English and metric-to-metric conversion factors
- Such conversion factors have an unlimited number of significant figures

  12 inches = 1 foot (exactly)
  4 quarts = 1 gallon (exactly)
  1 kilogram = 10^3 grams (exactly)
- Metric-to-metric conversion factors are derived using the meaning of the metric system prefixes

Conversion Factors from MEASURED Relationships

- All English-to-metric and metric-to-English conversion factors
- Such conversion factors have a specific number of significant figures, depending on the precision of the defining relationship

  1.00 lb = 454 g (three sig figs)
  1.000 lb = 453.6 g (four sig figs)
  1.0000 lb = 453.59 g (five sig figs)

Prefixes that INCREASE Base Unit Size

kilo- 10^3
mega- 10^6
giga- 10^9

Prefixes that DECREASE Base Unit Size

deci- 10^{-1}
centi- 10^{-2}
milli- 10^{-3}
micro- 10^{-6}
nano- 10^{-9}

2.22 Give both forms of the conversion factor that you would use to relate the following sets of units to each other.
a. Inch and centimeter b. Gram and pound
c. Liter and quart d. Meter and inches

2.23 Write the equality from which each of the following conversion factors is obtained.
a. 60 sec/1 min b. 12 in./1 ft c. 2.54 cm/1.00 in. d. 454 g/1.00 lb

2.24 Using dimensional analysis and a single conversion factor, solve each of the following problems.
a. 160,000 centimeters = ? meters

b. 24 nanometers = ? meters
c. 0.0030 kilometer = ? meters
d. 3.00 millimeters = ? meters

2.25 Using dimensional analysis and a single conversion factor, solve each of the following problems.
a. 6.4 grams = ? pounds
b. 6.4 pounds = ? grams
c. 53 centimeters = ? inches
d. 3.5 quarts = ? liters

2.26 Using dimensional analysis and two conversion factors, solve each of the following problems.
a. 2.0 meters = ? feet
b. 3.2 yards = ? centimeters
c. 3.2 feet = ? meters
d. 4.3 centimeters = ? feet

2.27 The human stomach produces approximately 2500 mL of gastric juice per day. What is the volume, in liters, of gastric juice produced?

2.28 The mass of premature babies is often determined in grams. If a premature baby weighs 1550 g, what is its mass in pounds?

2.8 Density

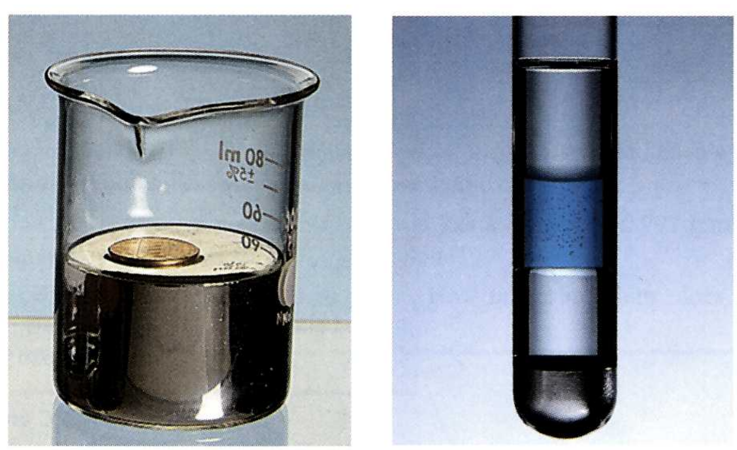
Learning Focus

Calculate the density of a substance, and be able to use density as a conversion factor to calculate the mass or volume of a substance.

Density *is the ratio of the mass of an object to the volume occupied by that object.*

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

People often speak of a substance as being heavier or lighter than another substance. What they actually mean is that the two substances have different densities; a specific volume of one substance is heavier or lighter than the same volume of the second substance (Figure 2.7). Equal masses of substances with different densities will occupy different volumes; the volume contrast is often very striking (see Figure 2.8).



(a) (b)

Figure 2.7 (a) The penny is less dense than the mercury it floats on. (b) Liquids that do not dissolve in one another and that have different densities float on one another, forming layers. The top layer is gasoline, with a density of about 0.8 g/mL. Next is water (plus food coloring), with a density of 1.0 g/mL. The next layer is carbon tetrachloride, with a density of 1.6 g/mL. The bottom layer is mercury, with a density of 13.6 g/mL.

Figure 2.8 Both of these items have a mass of 23 grams, but very different volumes. The volume differences result from the two items having different densities.

A correct density expression includes a number, a mass unit, and a volume unit. Although any mass and volume units can be used, densities are usually expressed in grams per cubic centimeter (g/cm^3) for solids, grams per milliliter (g/mL) for liquids, and grams per liter (g/L) for gases. Table 2.3 gives density values for a number of substances. Note that temperature must be specified with density values, because substances expand and contract with changes in temperature. For the same reason, the pressure of gases is also given with their density values.

Example 2.5 Calculating Density

A student determines that the mass of a 20.0-mL sample of olive oil is 18.4 g. What is the density of the olive oil in grams per milliliter?

Solution

To calculate density, we substitute the given mass and volume values into the defining formula for density.

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{18.4\ \text{g}}{20.0\ \text{mL}} = \underset{\text{Calculator answer}}{0.92\ \frac{\text{g}}{\text{mL}}} = \underset{\text{Correct answer}}{0.920\ \frac{\text{g}}{\text{mL}}}$$

Because both input numbers contain three significant figures, the density is specified to three significant figures.

▶ Density may be used as a conversion factor to convert from mass to volume or vice versa.

Density can be used as a conversion factor that relates the volume of a substance to its mass. This use of density enables us to calculate the volume of a substance if we know its mass. Conversely, the mass can be calculated if the volume is known.

Density conversion factors, like all other conversion factors, have two reciprocal forms. For a density of 1.03 g/mL, the two conversion factor forms are

$$\frac{1.03\ \text{g}}{1\ \text{mL}} \quad \text{and} \quad \frac{1\ \text{mL}}{1.03\ \text{g}}$$

Table 2.3
Densities of Selected Substances

Solids (25°C)			
gold	19.3 g/cm^3	table salt	2.16 g/cm^3
lead	11.3 g/cm^3	bone	1.7–2.0 g/cm^3
copper	8.93 g/cm^3	table sugar	1.59 g/cm^3
aluminum	2.70 g/cm^3	wood, pine	0.30–0.50 g/cm^3
Liquids (25°C)			
mercury	13.55 g/mL	water	0.997 g/mL
milk	1.028–1.035 g/mL	olive oil	0.92 g/mL
blood plasma	1.027 g/mL	ethyl alcohol	0.79 g/mL
urine	1.003–1.030 g/mL	gasoline	0.56 g/mL
Gases (25°C and 1 atmosphere pressure)			
chlorine	3.17 g/L	nitrogen	1.25 g/L
carbon dioxide	1.96 g/L	methane	0.66 g/L
oxygen	1.42 g/L	hydrogen	0.08 g/L
air (dry)	1.29 g/L		

Example 2.6 **Converting from Mass to Volume by Using Density as a Conversion Factor**

Blood plasma has a density of 1.027 g/mL at 25°C. What volume, in milliliters, does 125 g of plasma occupy?

Solution

Step 1: The given quantity is 125 g of blood plasma. The units of the desired quantity are milliliters. Thus our starting point is

$$125 \text{ g} = ? \text{ mL}$$

Step 2: The conversion from grams to milliliters can be accomplished in one step because the given density, used as a conversion factor, directly relates grams to milliliters. Of the two conversion factor forms

$$\frac{1.027 \text{ g}}{1 \text{ mL}} \quad \text{and} \quad \frac{1 \text{ mL}}{1.027 \text{ g}}$$

we will use the latter, because it allows for cancellation of gram units, leaving milliliters.

$$125 \cancel{\text{g}} \times \left(\frac{1 \text{ mL}}{1.027 \cancel{\text{g}}}\right) = ? \text{ mL}$$

Step 3: Doing the necessary arithmetic gives us our answer:

$$\left(\frac{125 \times 1}{1.027}\right) \text{mL} = 121.71372 \text{ mL} \quad \text{(calculator answer)}$$

$$= 122 \text{ mL} \quad \text{(correct answer)}$$

Even though the given density contained four significant figures, the correct answer is limited to three significant figures. This is because the other given number, the mass of blood plasma, had only three significant figures.

At room temperature and pressure, most of the naturally occurring elements—over 85% of them—are solids. Only 2 elements are liquids: Br and Hg. There are 10 gaseous elements: H, He, N, O, F, Ne, Cl, Ar, Kr, and Xe. A consideration of element densities shows that osmium is the most dense of the "solid-state" elements, mercury the most dense of the "liquid-state" elements, and xenon the most dense of the "gaseous-state" elements. Chemical Portraits 3 profiles these three "most dense" elements.

Practice Questions and Problems

2.29 A sample of mercury is found to have a mass of 524.5 g and to have a volume of 38.72 cm^3. What is its density in grams per cubic centimeter?

2.30 Acetone, the solvent in nail polish remover, has a density of 0.791 g/mL. What is the volume, in milliliters, of 20.0 g of acetone?

2.31 Nickel metal has a density of 8.90 g/cm^3. How much, in grams, does 15 cm^3 of nickel metal weigh?

Chemical Portraits 3 Solid, Liquid, and Gaseous "Most Dense" Elements

Osmium (Os)

Profile: With a density of 22.6 g/cm^3, osmium is the densest natural element. This hard, brittle, bluish-white element is used in small amounts to harden metal alloys used for items such as phonograph needles, fountain pen tips, and electrical contacts.

Biochemical considerations: In bulk form, pure Os is unaffected by air or water. However, powdered Os slowly reacts with air to produce OsO_4, a strong smelling volatile compound capable of causing lung and skin problems. This "OsO_4 problem" makes the use of pure Os metal impractical. Its use as an alloying agent is, however, safe. The name *osmium* comes from the Greek word "osme" which means "smell."

The metal osmium is obtained as a "by-product" from the processing of what types of ores?

Mercury (Hg)

Profile: With a density of 13.6 g/cm^3, mercury is the densest "liquid-state" element. Its chief source is the ore *cinnabar* (HgS). A semi-solid mixture of silver, tin, and mercury, which hardens upon standing, has been used widely to make dental fillings. The hardening process produces the compounds Ag_5Hg_8 and Sn_7Hg_8, nontoxic forms of mercury.

Biochemical considerations: When Hg is "spilled" from a broken thermometer, it forms little balls that roll around without adhering to anything. Such spills represent a danger since Hg vapor originating from the "balls" is toxic when inhaled. Spill clean-up involves sprinkling the area with sufur, to form HgS, a nonvolatile Hg-compound that can be picked up with a vacuum system.

What is the antidote for mercury poisoning associated with ingestion of mercury-containing substances?

Xenon (Xe)

Profile: With a density of 5.89 g/cm^3—approximately 5 times that of air—Xe is the densest "gaseous-state" element. The least abundant of the natural elements, Xe occurs in trace amounts in the atmosphere, which is its sole source. Air liquefaction is required for its isolation. Chemically, Xe is an extremely *unreactive* element, reacting only with fluorine, the most reactive of all elements.

Uses: Halogen automobile headlights contain Xe and it is also used in high-intensity photographic flash tubes. The passing of an electric charge through Xe gas produces a brilliant white "lightning-like" flash. Xe in a vacuum tube produces a beautiful blue glow when excited by an electric discharge.

Why are Xe-lamps more economical to operate than standard light bulbs?

See the text web site at **http://chemistry.college.hmco.com/students** for answers to the above questions and for further information.

Learning Focus

List and use the relationships among the Fahrenheit, Celsius, and Kelvin temperature scales, and list the relationships among the heat energy units joules, calories, and Calories.

2.9 Temperature Scales and Heat Energy

Heat is a form of energy. Temperature is an indicator of the tendency of heat energy to be transferred. Heat energy flows from objects of higher temperature to objects of lower temperature.

Three different temperature scales are in common use: Celsius, Kelvin, and Fahrenheit (Figure 2.9). Both the Celsius and the Kelvin scales are part of the metric measurement system; the Fahrenheit scale belongs to the English measurement system. Degrees of different size and different reference points are what produce the various temperature scales.

The *Celsius scale* is the scale most commonly encountered in scientific work. The normal boiling and freezing points of water serve as reference points on this scale, the former having a value of 100° and the latter 0°. Thus there are 100 "degree intervals" between the two reference points.

The *Kelvin scale* is a close relative of the Celsius scale. Both have the same size of degree, and the number of degrees between the freezing and boiling points of water is the same. The two scales differ only in the numbers assigned to the reference points. On the Kelvin scale, the boiling point of water is 373 kelvins (K) and the freezing point of water is 273 K. The choice of these reference points makes all temperature readings on the Kelvin scale positive values. Note that the degree sign (°) is not used with the Kelvin scale. For example, we say that an object has a temperature of 350 K (*not* 350°K).

▶ Zero on the Kelvin scale is known as *absolute zero*. It corresponds to the lowest temperature allowed by nature. How fast particles (molecules) move depends on temperature. The colder it gets, the more slowly they move. At absolute zero, movement stops. Scientists in laboratories have been able to attain temperatures as low as 0.0001 K, but a temperature of 0 K is impossible.

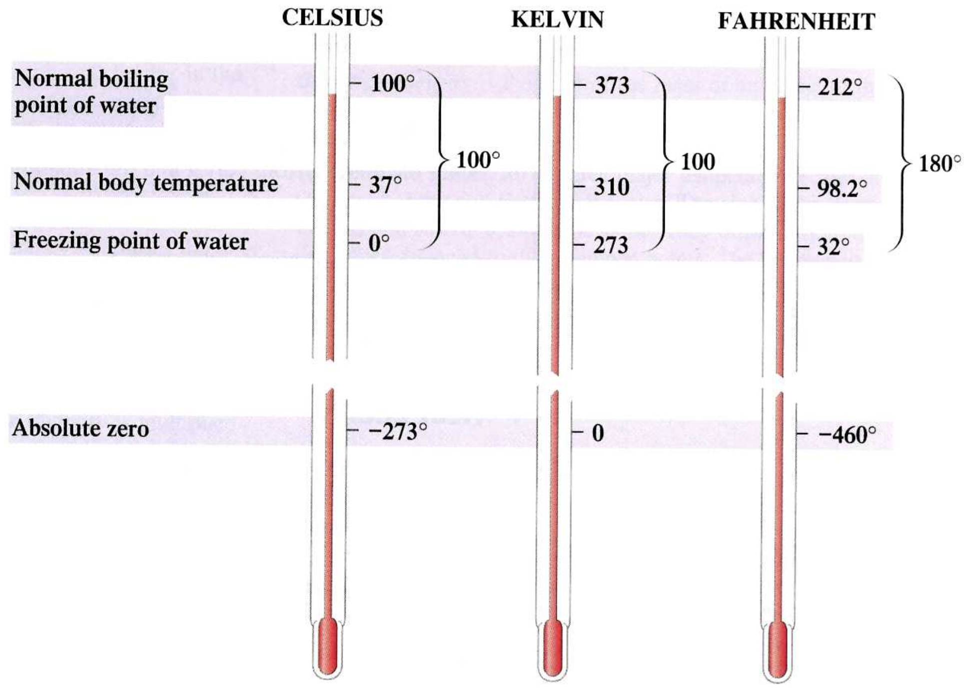


Figure 2.9 The relationships among the Celsius, Kelvin, and Fahrenheit temperature scales are determined by the degree sizes and the reference point values.

The *Fahrenheit scale* has a smaller degree size than the other two temperature scales. On this scale, there are 180 degrees between the freezing and boiling points of water as contrasted to 100 degrees on the other two scales. Thus the Celsius (and Kelvin) degree size is almost two times ($\frac{9}{5}$) larger than the Fahrenheit degree. Reference points on the Fahrenheit scale are 32° for the freezing point of water and 212° for the normal boiling point of water.

Conversions Between Temperature Scales

Because the size of the degree is the same, the relationship between the Kelvin and Celsius scales is very simple. No conversion factors are needed; all that is required is an adjustment for the differing numerical scale values. The adjustment factor is 273, the number of degrees by which the two scales are offset from one another.

$$K = °C + 273$$
$$°C = K - 273$$

The relationship between the Fahrenheit and Celsius scales can also be stated in an equation format.

$$°F = \frac{9}{5}(°C) + 32 \quad \text{or} \quad °C = \frac{5}{9}(°F - 32)$$

Example 2.7 **Converting from One Temperature Scale to Another**

Body temperature for a person with a high fever is found to be 104°F. To what is this temperature equivalent on the following scales?

a. Celsius scale **b.** Kelvin scale

Solution

a. We substitute 104° for °F in the equation

$$°C = \frac{5}{9}\,(°F - 32)$$

Then solving for °C gives

$$°C = \frac{5}{9}(104 - 32) = \frac{5}{9}(72) = 40°$$

b. Using the answer from part **a** and the equation

$$K = °C + 273$$

we get, by substitution,

$$K = 40° + 273 = 313$$

Heat Energy

▶ In discussions involving nutrition, the energy content of foods, and dietary tables, the term *Calorie* (spelled with a capital C) is used. The dietetic Calorie is actually 1 kilocalorie (1000 calories). The statement that an oatmeal raisin cookie contains 60 Calories means that 60 kcal (60,000 cal) of energy is released when the cookie is metabolized (undergoes chemical change) within the body.

The form of energy most often required for or released by chemical reactions and physical changes is *heat energy*. A commonly used unit for the measurement of heat energy is the calorie. A **calorie** (cal) *is the amount of heat energy needed to raise the temperature of 1 gram of water by 1 degree Celsius.* For large amounts of heat energy, the measurement is usually expressed in kilocalories.

$$1 \text{ kilocalorie} = 1000 \text{ calories}$$

Another unit for heat energy that is used with increasing frequency is the joule (J). The relationship between the joule (which rhymes with *pool*) and the calorie is

$$1 \text{ calorie} = 4.184 \text{ joules}$$

Heat energy values in calories can be converted to joules by using the conversion factor

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

Practice Questions and Problems

2.32 An oven for baking pizza operates at approximately 525°F. What is this temperature in degrees Celsius?

2.33 A comfortable temperature for bathtub water is 35°C. What is this temperature in degrees Fahrenheit?

2.34 The body temperature for a hypothermia victim is found to have dropped to 29°C. What is this temperature on the Kelvin scale?

2.35 Which quantity of heat energy in each of the following pairs of heat energy values is the larger?
a. 2.0 joules or 2.0 calories
b. 1.0 kilocalorie or 92 calories
c. 100 Calories or 100 calories
d. 2.3 Calories or 1000 kilocalories

CONCEPTS TO REMEMBER

The metric system. The metric system, the measurement system preferred by scientists, is a decimal system in which larger and smaller units of a quantity are related by factors of 10. Prefixes are used to designate relationships between the basic unit and larger or smaller units of a quantity. Units in the metric system include the gram (mass), liter (volume), and meter (length).

Exact and inexact numbers. Numbers are of two kinds: exact and inexact. An exact number has a value that has no uncertainty associated with it. Exact numbers occur in definitions, in counting, and in simple fractions. An inexact number has a value that has a degree of uncertainty associated with it. Inexact numbers are generated anytime a measurement is made.

Significant figures. Significant figures in a measurement are those digits that are certain, plus a last digit that has been estimated. The maximum number of significant figures possible in a measurement is determined by the design of the measuring device.

Calculations and significant figures. Calculations should never improve (or decrease) the precision of experimental measurements. In multiplication and division, the number of significant figures in the answer is the same as that in the measurement containing the fewest significant figures. In addition and subtraction, the answer has no more digits to the right of the decimal point than are found in the measurement with the fewest digits to the right of the decimal point.

Scientific notation. Scientific notation is a system for writing decimal numbers in a more compact form that greatly simplifies the mathematical operations of multiplication and division. In this system, numbers are expressed as the product of a number between 1 and 10 and 10 raised to a power.

Dimensional analysis. Dimensional analysis is a general problem-solving method in which the units associated with numbers are used as a guide in setting up calculations. A given quantity is multiplied by one or more conversion factors in such a manner that the unwanted (original) units are canceled, leaving only the desired units.

Density. Density is the ratio of the mass of an object to the volume occupied by that object. A correct density expression includes a number, a mass unit, and a volume unit.

Temperature scales. The three major temperature scales are the Celsius, Kelvin, and Fahrenheit scales. The size of the degree for the Celsius and Kelvin scale is the same. They differ only in the numerical values assigned to the reference points. The Fahrenheit scale has a smaller degree size than the other two temperature scales.

Heat energy. The most commonly used unit of measurement for heat energy is the calorie. A calorie is the amount of heat energy needed to raise the temperature of 1 gram of water by 1 degree Celsius.

KEY REACTIONS AND EQUATIONS

1. Density of a substance (Section 2.7)

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

2. Conversion of temperature readings from one scale to another (Section 2.8)

$$\text{K} = {}^\circ\text{C} + 273 \qquad {}^\circ\text{C} = \text{K} - 273$$

$$ {}^\circ\text{F} = \frac{9}{5}({}^\circ\text{C}) + 32 \qquad {}^\circ\text{C} = \frac{5}{9}({}^\circ\text{F} - 32)$$

KEY TERMS

Calorie (2.9)
Conversion factor (2.7)
Density (2.8)
Dimensional analysis (2.7)
Exact number (2.3)
Gram (2.2)
Inexact number (2.3)
Liter (2.2)
Mass (2.2)
Measurement (2.1)
Meter (2.2)
Rounding off (2.5)
Scientific notation (2.6)
Significant figures (2.4)
Weight (2.2)

ADDITIONAL PROBLEMS

2.36 Round off the number 4.7205059 to the indicated number of significant figures.
a. Six b. Five c. Four d. Two

2.37 Write each of the following numbers in scientific notation to the number of significant figures indicated in parentheses.
a. 0.00300300 (three) b. 936,000 (two)
c. 23.5003 (three) d. 450,000,001 (six)

2.38 For each of the pairs of units listed, indicate whether the first unit is larger or smaller than the second unit, and then indicate how many times larger or smaller it is.
a. Milliliter, liter b. Kiloliter, microliter
c. Nanoliter, deciliter d. Centiliter, megaliter

2.39 Indicate how each of the following conversion factors should be interpreted in terms of significant figures present.
a. $\frac{2.540\text{ cm}}{1.000\text{ in.}}$ b. $\frac{453.6\text{ g}}{1.000\text{ lb}}$ c. $\frac{2.113\text{ pt}}{1.00\text{ L}}$ d. $\frac{10^{-9}\text{ m}}{1\text{ nm}}$

2.40 Without actually performing the following multiplications, specify the number of significant figures that the answer should contain. The first listed number is an exact number, and the rest of the numbers are measured numbers.
a. $2 \times 2.00 \times 3.00$ b. $32 \times 4.31 \times 52{,}000$
c. $3 \times 3.0 \times 3.00$ d. $323 \times 320 \times 3200$

2.41 How many significant figures must the number Q possess, in each case, to make the following mathematical equations valid from the standpoint of significant figures?
a. $6.000 \times Q = 4.0$ b. $6.000 \times 4.0 = Q$
c. $5.250 + Q = 7.03$ d. $0.7777 - Q = 0.011$

2.42 A 1-gram sample of a powdery white solid is found to have a volume of 2 cubic centimeters. Calculate the solid's density using the following uncertainty specifications, and express your answers in scientific notation.
a. 1.0 g and 2.0 cm^3 b. 1.000 g and 2.00 cm^3
c. 1.0000 g and 2.0000 cm^3 d. 1.000 g and 2.0000 cm^3

2.43 Which is the higher temperature, -10°C or 10°F?

2.44 An individual weighs 83.2 kg and is 1.92 m tall. What are the person's equivalent measurements in pounds and feet?

2.45 What is wrong with the statement "The number of objects is exactly 12.00"?

PRACTICE TEST ▶ True/False

2.46 The density of an object is the ratio of its mass to its height.

2.47 Two conversion factors, which have a reciprocal relationship, can be derived from the equality 24 hours = 1 day.

2.48 A micrometer is a smaller metric system unit than a picometer.

2.49 All of the zeros in the measured number 0.0040400 are significant.

2.50 The number 344,700, when rounded to three significant figures, becomes 345.

2.51 The meter is the base unit of mass in the metric system.

2.52 1 mL, 1 cm^3, and 1 cc are three representations for the same volume.

2.53 All numbers have a degree of uncertainty associated with them.

2.54 The answer for the problem (3.11 + 9.2) should have an uncertainty of tenths.

2.55 The number 3.21×10^2 is larger than the number 314.

2.56 The number 0.0030, when it is expressed in scientific notation, becomes 3×10^{-3}.

2.57 The metric system prefix *micro* has the mathematical meaning 10^{-6}.

2.58 Addition of the value 273 to a Fahrenheit scale temperature reading will convert it into a Kelvin scale temperature reading.

2.59 Readings from a ruler scale with a smallest scale marking of 1 mm should be estimated to the closest 0.1 mm.

2.60 Solving the problem "How many feet are there in 3.72 yards?" using dimensional analysis requires the use of the conversion factor (1 yd/3 ft).

PRACTICE TEST ▶ Multiple Choice

2.61 The "mathematical meanings" associated with the metric system prefixes *centi, milli,* and *kilo* are respectively,
a. 10^{-3}, 10^{-4}, 10^{-6}
b. 10^{-2}, 10^{-3}, 10^{6}
c. 10^{-2}, 10^{-3}, 10^{3}
d. 10^{2}, 10^{-3}, 10^{3}

2.62 To what decimal position should a volume measurement be recorded if the smallest markings on the measurement scale are tenths of a milliliter?
a. To the closest milliliter
b. To tenths of a milliliter
c. To hundredths of a milliliter
d. To thousandths of a milliliter

2.63 Which of the following statements about the "significance" of zeros in recorded measurements is *incorrect*?
a. Leading zeros are never significant.
b. Confined zeros are always significant.
c. Trailing zeros are not always significant.
d. Trailing zeros and confined zeros are not always significant.

2.64 The number 43250, when rounded off to three significant figures, becomes
a. 432
b. 433
c. 43200
d. 43300

2.65 The number 105.00, when expressed in scientific notation, becomes
a. 1.05×10^{-2}
b. 1.0500×10^{-2}
c. 1.05×10^{2}
d. 1.0500×10^{2}

2.66 The calculator answer obtained by multiplying the measurements 62.32 and 7.00 is 436.24. This answer
a. is correct as written
b. should be rounded to 436.2
c. should be rounded to 436
d. should be rounded to 440

2.67 The correct answer obtained by adding the measurements 8.1, 2.16, and 3.123 contains
a. two significant figures
b. three significant figures
c. four significant figures
d. five significant figures

2.68 According to dimensional analysis, which of the following is the correct setup for the problem "How many milligrams are there in 67 kilograms?"

a. $67\text{ kg} \times \left(\frac{1\text{ g}}{10^3\text{ kg}}\right) \times \left(\frac{1\text{ mg}}{10^{-3}\text{ g}}\right)$

b. $67\text{ kg} \times \left(\frac{10^3\text{ g}}{1\text{ kg}}\right) \times \left(\frac{1\text{ mg}}{10^{-3}\text{ g}}\right)$

c. $67\text{ kg} \times \left(\frac{10^3\text{ mg}}{1\text{ kg}}\right) \times \left(\frac{10^{-3}\text{ g}}{1\text{ mg}}\right)$

d. $67\text{ kg} \times \left(\frac{1\text{ g}}{10^3\text{ kg}}\right) \times \left(\frac{10^{-3}\text{ kg}}{1\text{ mg}}\right)$

2.69 If object A weighs 8 g and has a volume of 4 mL and object B weighs 12 g and has a volume of 3 mL, then
a. B is less dense than A
b. A and B have equal densities
c. B is twice as dense as A
d. B is four times as dense as A

2.70 Which of the following comparisons of the size of a degree on the major temperature scales is *correct*?
a. A kelvin is larger than a Celsius degree.
b. A Fahrenheit degree and a Celsius degree are equal in size.
c. A Fahrenheit degree is larger than a kelvin.
d. A Celsius degree and a kelvin are equal in size.