1.4 Units of Measurement

Many properties of matter are *quantitative*; that is, they are associated with numbers. When a number represents a measured quantity, the units of that quantity must always be specified. To say that the length of a pencil is 17.5 is meaningless. To say that it is 17.5 centimeters (cm) properly specifies the length. The units used for scientific measurements are those of the **metric system**.

The metric system, which was first developed in France during the late eighteenth century, is used as the system of measurement in most countries throughout the world. The United States has traditionally used the English system, although use of the metric system has become more common in recent years. For example, the contents of most canned goods and soft drinks in grocery stores are now given in metric as well as in English units as shown in Figure 1.16.

Figure 1.16 Metric measurements are becoming increasingly common in the United States, as exemplified by the volume printed on this container.

SI Units

In 1960 an international agreement was reached specifying a particular choice of metric units for use in scientific measurements. These preferred units are called **SI units**, after the French *Système International d'Unités*. The SI system has seven *base units* from which all other units are derived. Table 1.4 lists these base units and their symbols. In this chapter we will consider the base units for length, mass, and temperature.

TABLE 1.4 SI Base Units

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

[a]The abbreviation sec is frequently used.

Prefixes are used to indicate decimal fractions or multiples of various units. For example, the prefix *milli-* represents a 10^{-3} fraction of a unit: A milligram (mg) is 10^{-3} gram (g), a millimeter (mm) is 10^{-3} meter (m), and so forth. Table 1.5 presents the prefixes commonly encountered in chemistry. In using the SI system and in working problems throughout this text, you must be comfortable using exponential notation. If you are unfamiliar with exponential notation or want to review it, refer to Appendix A.1.

TABLE 1.5 Selected Prefixes Used in the Metric System

Prefix	Abbreviation	Meaning	Example
Giga	G	10^9	1 gigameter (Gm) $= 1 \times 10^9$ m
Mega	M	10^6	1 megameter (Mm) $= 1 \times 10^6$ m
Kilo	k	10^3	1 kilometer (km) $= 1 \times 10^3$ m
Deci	d	10^{-1}	1 decimeter (dm) = 0.1 m
Centi	c	10^{-2}	1 centimeter (cm) = 0.01 m
Milli	m	10^{-3}	1 millimeter (mm) = 0.001 m
Micro	μ[a]	10^{-6}	1 micrometer (μm) $= 1 \times 10^{-6}$ m
Nano	n	10^{-9}	1 nanometer (nm) $= 1 \times 10^{-9}$ m
Pico	p	10^{-12}	1 picometer (pm) $= 1 \times 10^{-12}$ m
Femto	f	10^{-15}	1 femtometer (fm) $= 1 \times 10^{-15}$ m

[a]This is the Greek letter mu (pronounced "mew").

Although non-SI units are being phased out, there are still some that are commonly used by scientists. Whenever we first encounter a non-SI unit in the text, the proper SI unit will also be given.

Length and Mass

The SI base unit of *length* is the meter (m), a distance only slightly longer than a yard. The relations between the English and metric system units that we will use most frequently in this text appear on the back inside cover. We will discuss how to convert English units into metric units, and vice versa, in Section 1.6.

Mass* is a measure of the amount of material in an object. The SI base unit of mass is the kilogram (kg), which is equal to about 2.2 pounds (lb). This base unit is unusual because it uses a prefix, *kilo*-, instead of the word *gram* alone. We obtain other units for mass by adding prefixes to the word *gram*.

SAMPLE EXERCISE 1.2

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

Solution In each case we can refer to Table 1.5, finding the prefix related to each of the decimal fractions: **(a)** nanogram, ng; **(b)**microsecond, μs; **(c)** millimeter, mm.

PRACTICE EXERCISE

(a) What decimal fraction of a second is a picosecond, ps? **(b)** Express the measurement 6.0×10^3 m using a prefix to replace the power of ten. **(c)** Use standard exponential notation to express 3.76 mg in grams.

Answers: **(a)** 10^{-12} second; **(b)** 6.0 km; **(c)** 3.76×10^{-3} g

---

Temperature

We sense temperature as a measure of the hotness or coldness of an object. Indeed, temperature determines the direction of heat flow. Heat always flows spontaneously from a substance at higher temperature to one at lower temperature. Thus, we feel the influx of energy when we touch a hot object, and we know that the object is at a higher temperature than our hand.

The temperature scales commonly employed in scientific studies are the Celsius and Kelvin scales. The **Celsius scale** is also the everyday scale of temperature in most countries (Figure 1.17). It was originally based on the assignment of 0°C to the freezing point of water and 100°C to its boiling point at sea level (Figure 1.18).

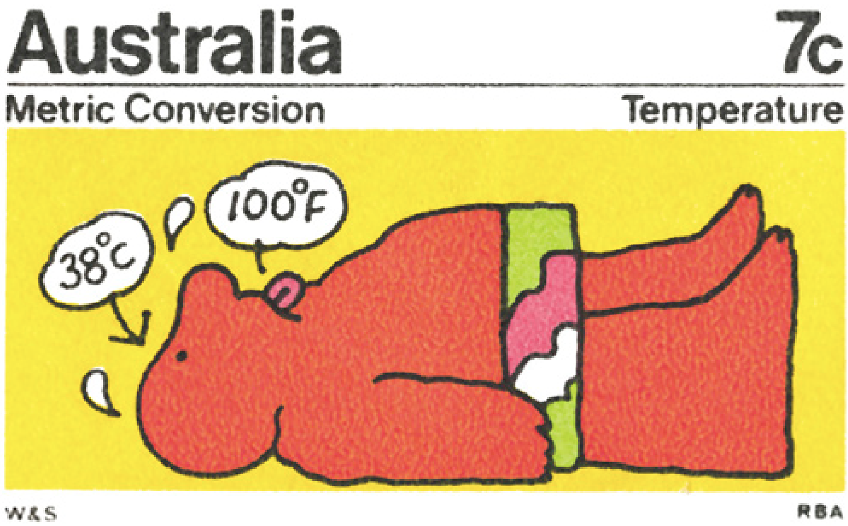

Figure 1.17 Many countries employ the Celsius temperature scale in everyday use, as illustrated by this Australian stamp.

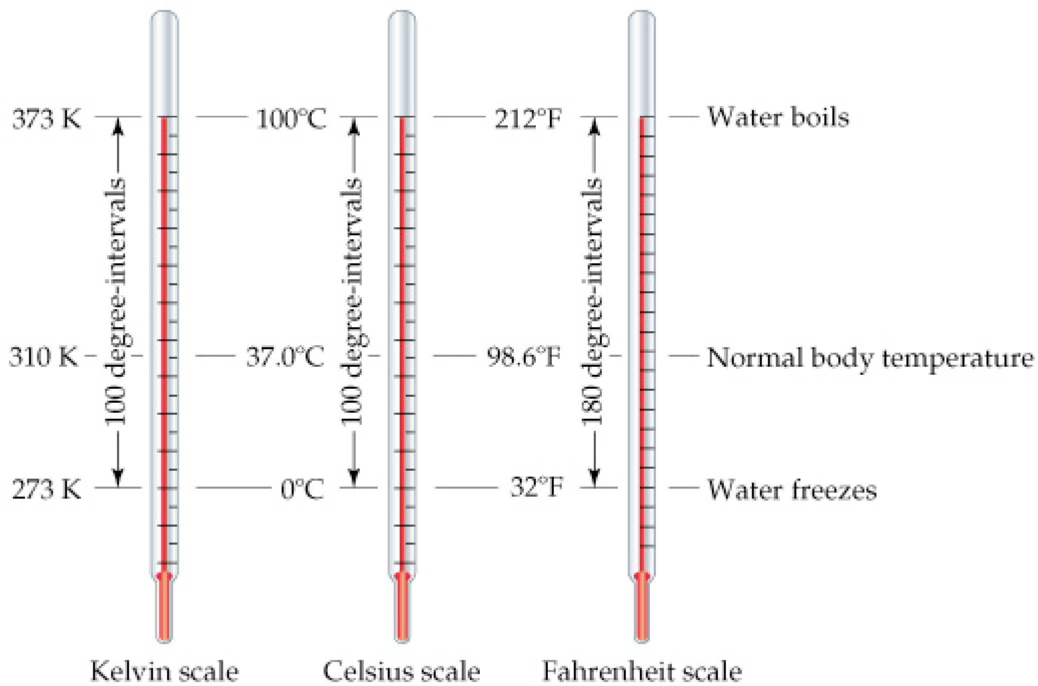

Figure 1.18 Comparison of the Kelvin, Celsius, and Fahrenheit temperature scales.

The **Kelvin scale** is the SI temperature scale, and the SI unit of temperature is the kelvin (K). Historically, the Kelvin scale was based on the properties of gases; its origins will be considered

in Chapter 10. Zero on this scale is the lowest attainable temperature, $-273.15°C$, a temperature referred to as *absolute zero*. Both the Celsius and Kelvin scales have equal-sized units—that is, a kelvin is the same size as a degree Celsius. Thus, the Kelvin and Celsius scales are related as follows:

$$K = °C + 273.15 \quad [1.1]$$

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

The common temperature scale in the United States is the *Fahrenheit scale*, which is not generally used in scientific studies. On the Fahrenheit scale water freezes at $32°F$ and boils at $212°F$.

The Fahrenheit and Celsius scales are related as follows:

$$°C = \frac{5}{9}(°F - 32) \quad \text{or} \quad °F = \frac{9}{5}(°C) + 32 \quad [1.2]$$

SAMPLE EXERCISE 1.3

If a weather forecaster predicts that the temperature for the day will reach $31°C$, what is the predicted temperature **(a)** in K; **(b)** in $°F$?

Solution

(a) Using Equation 1.1, we have $K = 31 + 273 = 304\ K$

(b) Using Equation 1.2, we have $°F = \frac{9}{5}(31) + 32 = 56 + 32 = 88°F$

PRACTICE EXERCISE

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

Answers: **(a)** 261.7 K; **(b)** $11.3°F$

Derived SI Units

The SI base units in Table 1.4 are used to derive the units of other quantities. To do so, we use the defining equation for the quantity, substituting the appropriate base units. For example, speed is defined as the ratio of distance to elapsed time. Thus, the SI unit for speed is the SI unit for distance (length) divided by the SI unit for time, m/s, which we read as "meters per second." We will encounter many derived units, such as those for force, pressure, and energy, later in this text. In this chapter we examine the derived units for volume and density.

Volume

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

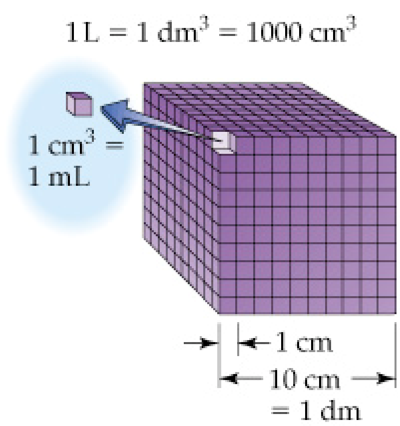

Figure 1.19 A liter is the same volume as a cubic decimeter, 1 L $=$ 1 dm^3. Each cubic decimeter contains 1000 cubic centimeters, 1 $dm^3 = 1000\ cm^3$. Each cubic centimeter equals a milliliter, 1 $cm^3 =$ 1 mL.

The terms *milliliter* and *cubic centimeter* are used interchangeably in expressing volume.

The devices used most frequently in chemistry to measure volume are illustrated in Figure 1.20. Syringes, burets, and pipets deliver liquids with more accuracy than graduated cylinders. Volumetric flasks are used to contain specific volumes of liquid.

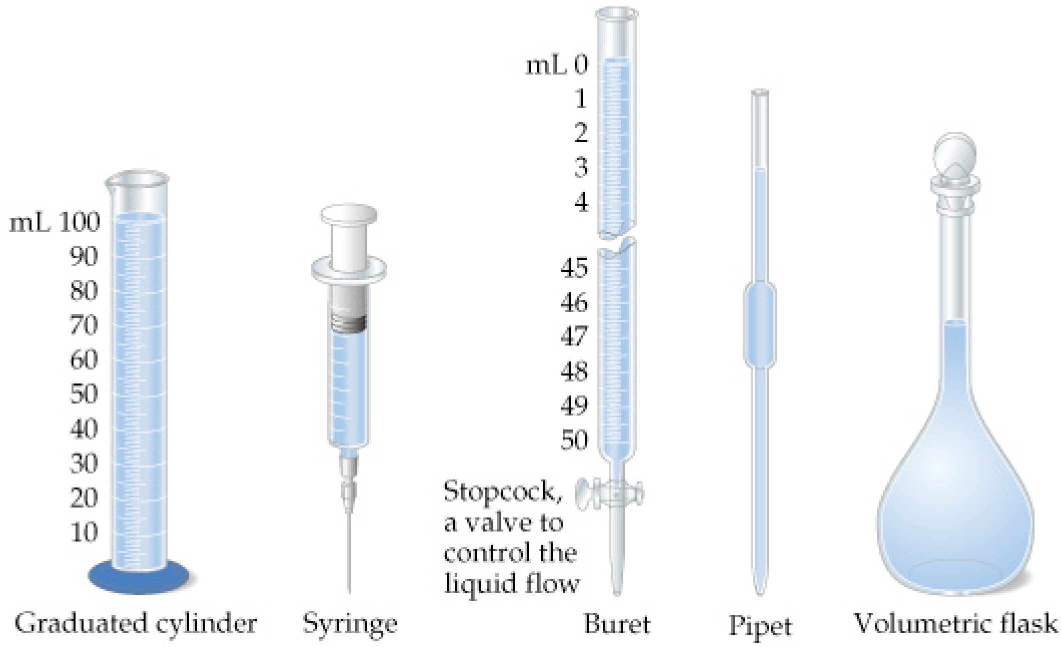

Figure 1.20 Common devices used in chemistry laboratories for the measurement and delivery of volumes of liquid. The graduated cylinder, syringe, and buret are used to deliver variable volumes of liquid; the pipet is used to deliver a specific volume of liquid. The volumetric flask contains a specific volume of liquid when filled to the mark.

Density

Density is widely used to characterize substances. It is defined as the amount of mass in a unit volume of the substance:

$$\text{Density} = \frac{\text{mass}}{\text{volume}} \quad [1.3]$$

The densities of solids and liquids are commonly expressed in units of grams per cubic centimeter (g/cm^3) or grams per milliliter (g/mL). The densities of some common substances are listed in Table 1.6. It is no coincidence that the density of water is $1.00\ g/mL$; the gram was originally defined as the mass of 1 mL of water at a specific temperature. Because most substances change volume when heated or cooled, densities are temperature dependent. When reporting densities, the temperature should be specified. We usually assume that the temperature is $25°C$, close to normal room temperature, if no temperature is reported.

TABLE 1.6 Densities of Some Selected Substances at 25°C

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

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

SAMPLE EXERCISE 1.4

(a) Calculate the density of mercury if 1.00×10^2 g occupies a volume of 7.36 cm^3.

(b) Calculate the volume of 65.0 g of the liquid methanol (wood alcohol) if its density is 0.791 g / mL.

(c) What is the mass in grams of a cube of gold (density $= 19.32\ g/cm^3$) if the length of the cube is 2.00 cm.

Solution

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

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

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

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

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

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

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

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

PRACTICE EXERCISE

(a) Calculate the density of a 374.5-g sample of copper if it has a volume of 41.8 cm^3. **(b)** A student needs 15.0 g of ethanol for an experiment. If the density of the alcohol is 0.789 g/mL, how many milliliters of alcohol are needed? **(c)** What is the mass, in grams, of 25.0 mL of mercury (density $= 13.6\ \text{g/mL}$)?

Answers: **(a)** 8.96 g/cm^3; **(b)** 19.0 mL; **(c)** 340 g

---

* Mass and weight are not interchangeable terms and are often incorrectly thought to be the same. The weight of an object is the force that the mass exerts due to gravity. In space, where gravitational forces are very weak, an astronaut can be weightless, but he or she cannot be massless. In fact, the astronaut's mass in space is the same as it is on Earth.