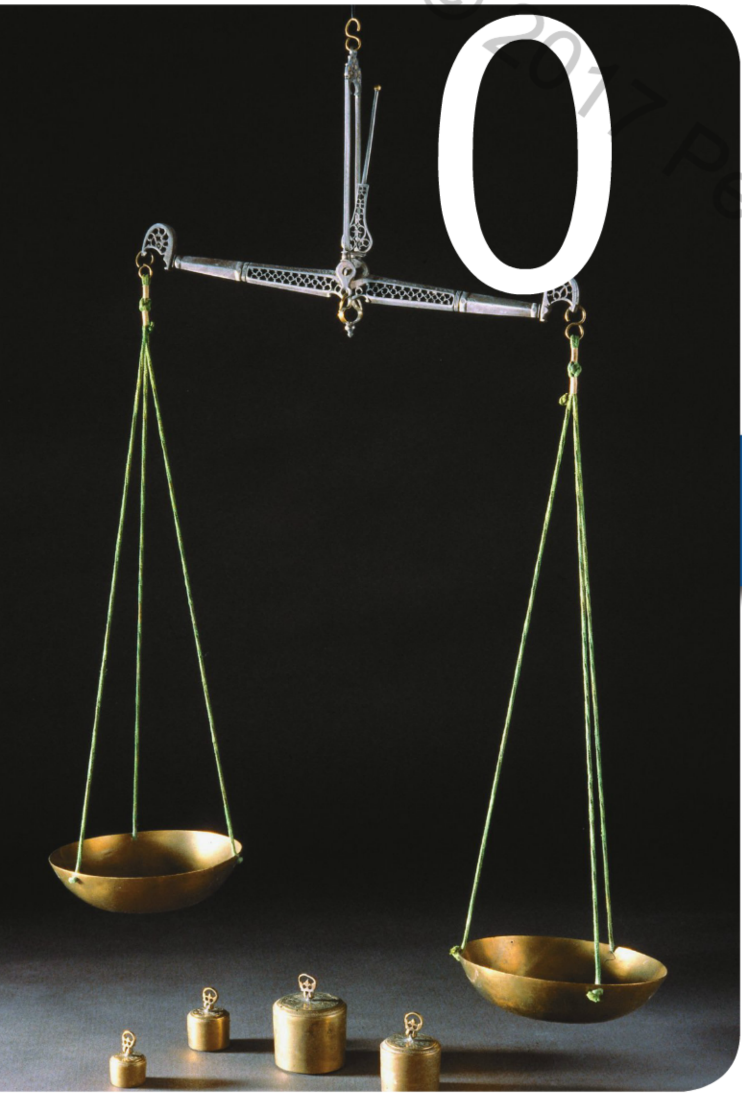
STUDENTS – PLEASE READ. REALLY.

Who ever heard of beginning a book with Chapter 0? It's not that the material in this chapter is unimportant. In fact, it's *very* important because it gives you some tools that you must know how to use before beginning your study of chemistry. The problem is that you and your classmates, who typically take Introductory Chemistry in the first year of college, are so diverse in your backgrounds and preparations that many of you will have

CHAPTER 0

Chemical Tools: Experimentation and Measurement



Accurate measurements, such as those obtained on this 300-year-old Persian balance, have always been a crucial part of scientific experiments.

Life has changed more in the past two centuries than in all the previously recorded span of human history. The Earth's population has increased sevenfold since 1800, from about 1 billion to 7 billion, and life expectancy has nearly doubled because of our ability to control diseases, increase crop yields, and synthesize medicines. Methods of transportation have changed from horses and buggies to automobiles and airplanes because of our ability to harness the energy in petroleum. Many goods are now made of polymers and ceramics instead of wood and metal because of our ability to manufacture materials with properties unlike any found in nature.

In one way or another, all these changes involve **chemistry**, the study of the composition, properties, and transformations of matter. Chemistry is deeply involved in both the changes that take place in nature and the profound social changes of the past two centuries. In addition, chemistry is central to the current revolution in molecular biology that is now exploring the details of how life is genetically controlled. No educated person today can understand the modern world without a basic knowledge of chemistry.

already taken a chemistry Advanced Placement course in high school while many others of you will have had either no high school background in chemistry or at best a one-semester introduction. Thus, some of you will have already learned the material in this chapter and will be comfortable using it, while others of you will have to carefully brush up on it or learn it for the first time.

The choice about whether you need to learn, brush up on, or move quickly past the material in Chapter 0 is yours (or your professor's). But be careful in making your choice. You *do* need to master and feel comfortable with this chapter before going on.

CONTENTS

0.1 EXPERIMENT → HYPOTHESIS → THEORY: APPROACHING CHEMISTRY

By opening this book, you have already decided that you need to know more about chemistry. Perhaps you want to learn how medicines are made, how genes can be sequenced and manipulated, how fertilizers and pesticides work, how living organisms function, how new high-temperature ceramics are used in space vehicles, or how microelectronic circuits are etched onto silicon chips. How do you approach chemistry?

One way to approach chemistry or any other science is to look around you and try to think of logical explanations for what you see. You would certainly observe, for instance, that different substances have different forms and appearances. Some substances are gases, some are liquids, and some are solids; some are hard and shiny, but others are soft and dull. You'd also observe that different substances behave differently: Iron rusts but gold does not; copper conducts electricity but sulfur doesn't. How can these and a vast number of other observations be explained?

► Gold, one of the most valuable of elements, has been prized since antiquity for its beauty and resistance to corrosion.

In fact, the natural world is far too complex to be understood by looking and thinking alone, so a more active approach is needed. Specific questions must be asked and experiments must be carried out to find their answers. Only when the results of many experiments are known can we devise a tentative interpretation, or *hypothesis*, that explains the results. The hypothesis, in turn, can be used to make more predictions and to suggest more experiments until a consistent explanation, or **theory**, is finally arrived at.

Experiments and observations come first in chemistry or any other science, and theories to explain those observations come later after much hard work and a lot of thought. Nevertheless, it's important to keep in mind that scientific theories are not laws of nature and can never be absolutely proven. There's always the chance that a new experiment might give results that can't be explained by present theory. All a theory can do is to represent the best explanation that we can come up with at any given time. If new experiments uncover results that present theories can't explain, the theories will have to be modified or perhaps even replaced.

0.2 EXPERIMENTATION AND MEASUREMENT IN CHEMISTRY

Chemistry is an experimental science. But if our experiments are to be reproducible, we must be able to describe fully the substances we're working with—their amounts, volumes, temperatures, and so forth. Thus, one of the most important requirements in chemistry is that we have a way to measure things.

Under an international agreement concluded in 1960, scientists throughout the world now use the International System of Units for measurement, abbreviated **SI** for the French *Système Internationale d'Unités*. Based on the metric system, which is used in all countries except the United States, Liberia, and Myanmar, the SI system has seven fundamental units (**Table 0.1**). These seven fundamental units, along with others derived from them, suffice for all scientific measurements. We'll look at three of the most commonly used units in this chapter—those for mass, length, and temperature—and will discuss others as the need arises in later chapters.

TABLE 0.1 The Seven Fundamental SI Units of Measure

Physical Quantity	Name of Unit	Abbreviation
Mass	kilogram	kg
Length	meter	m
Temperature	kelvin	K
Amount of substance	mole	mol
Time	second	s
Electric current	ampere	A
Luminous intensity	candela	cd

One problem with any system of measurement is that the sizes of the units often turn out to be inconveniently large or small. For example, a chemist describing the diameter of a sodium atom (0.000 000 000 372 m) would find the meter (m) to be inconveniently large, but an astronomer describing the average distance from the Earth to the Sun (150,000,000,000 m) would find the meter to be inconveniently small. For this reason, SI units are modified by using prefixes when they refer to either smaller or larger quantities. For example, the prefix *milli-* means one-thousandth, so a *milli*meter (mm) is 1/1000 of 1 meter. Similarly, the prefix *kilo-* means one thousand, so a *kilo*meter (km) is 1000 meters. [Note that the SI unit for mass (kilogram) already contains the *kilo-* prefix.] A list of prefixes is shown in **Table 0.2**, with the most commonly used ones in red.

TABLE 0.2 Some Prefixes for Multiples of SI Units

Factor	Prefix	Symbol	Example
$1,000,000,000,000 = 10^{12}$	tera	T	1 teragram (Tg) $= 10^{12}$ g
$1,000,000,000 = 10^{9}$	giga	G	1 gigameter (Gm) $= 10^{9}$ m
$1,000,000 = 10^{6}$	mega	M	1 megameter (Mm) $= 10^{6}$ m
$1,000 = 10^{3}$	kilo	k	1 kilogram (kg) $= 10^{3}$ g
$100 = 10^{2}$	hecto	h	1 hectogram (hg) = 100 g
$10 = 10^{1}$	deka	da	1 dekagram (dag) = 10 g
$0.1 = 10^{-1}$	deci	d	1 decimeter (dm) = 0.1 m
$0.01 = 10^{-2}$	centi	c	1 centimeter (cm) = 0.01 m
$0.001 = 10^{-3}$	milli	m	1 milligram (mg) = 0.001 g
$*0.000\,001 = 10^{-6}$	micro	μ	1 micrometer (μm) $= 10^{-6}$ m
$*0.000\,000\,001 = 10^{-9}$	nano	n	1 nanosecond (ns) $= 10^{-9}$ s
$*0.000\,000\,000\,001 = 10^{-12}$	pico	p	1 picosecond (ps) $= 10^{-12}$ s
$*0.000\,000\,000\,000\,001 = 10^{-15}$	femto	f	1 femtomole (fmol) $= 10^{-15}$ mol

*It is becoming common in scientific work to leave a thin space every three digits to the right of the decimal point in very small numbers, analogous to the comma placed every three digits to the left of the decimal point in large numbers.

Notice how numbers that are either very large or very small are indicated in Table 0.2 using an exponential format called **scientific notation**. For example, the number 55,000 is written in scientific notation as 5.5×10^4, and the number 0.003 20 as 3.20×10^{-3}. Review Appendix A if you are uncomfortable with scientific notation or if you need to brush up on how to do mathematical manipulations on numbers with exponents.

Notice also that all measurements contain both a number and a unit label. A number alone is not much good without a unit to define it. If you asked a friend how far it was to the nearest tennis court, the answer "3" alone wouldn't tell you much. 3 blocks? 3 kilometers? 3 miles?

PROBLEM 0.1 Express the following quantities in scientific notation:
(a) The diameter of a sodium atom, 0.000 000 000 372 m
(b) The distance from the Earth to the Sun, 150,000,000,000 m

PROBLEM 0.2 What units do the following abbreviations represent?
(a) μg **(b)** dm **(c)** ps **(d)** kA **(e)** mmol

0.3 FUNDAMENTAL UNITS: MEASURING MASS

Let's look in more detail at the measurement of some common quantities in chemistry, beginning with mass. **Mass** is defined as the amount of *matter* in an object. **Matter**, in turn, is a catchall term used to describe anything with a physical presence—anything you can touch, taste, or smell. (Stated more scientifically, matter is anything that has mass.) Mass is measured in SI units by the **kilogram (kg**; 1 kg = 2.205 U.S. lb). Because the kilogram is too large for many purposes in chemistry, the metric **gram** (**g**; 1 g = 0.001 kg), the **milligram** (**mg**; 1 mg = 0.001 g $= 10^{-6}$ kg), and the **microgram** (**μg**; 1 μg = 0.001 mg $= 10^{-6}$ g $= 10^{-9}$ kg) are more commonly used. One gram is a bit less than half the mass of a new U.S. dime.

1 kg = 1000 g = 1,000,000 mg = 1,000,000,000 μg (2.205 lb)
1 g = 1000 mg = 1,000,000 μg (0.035 27 oz)
1 mg = 1000 μg

▲ The mass of a U.S. dime is approximately 2.27 g. How many milligrams is this?

The standard kilogram is defined as the mass of the *international prototype kilogram*, a cylindrical bar of platinum–iridium alloy stored in a vault in a suburb of Paris, France. There are 40 original copies of this bar distributed throughout the world, with two (Numbers K4 and K20) stored at the U.S. National Institute of Standards and Technology near Washington, D.C.

The terms "mass" and "weight," although often used interchangeably, have quite different meanings. *Mass* is a measure of the amount of matter in an object, whereas *weight* is a measure of the force that gravity exerts on an object. Mass is independent of an object's location: Your body has the same amount of matter whether you're on Earth or on the Moon. Weight, however, *does* depend on an object's location. If you weigh 140 lb on Earth, you would weigh only about 23 lb on the Moon, which has a lower gravity than Earth.

At the same location on Earth, two objects with identical masses experience an identical pull of Earth's gravity and have identical weights. Thus, the mass of an object can be measured by comparing its weight to the weight of a reference standard of known mass. Much of the confusion between mass and weight is simply due to a language problem. We speak of "weighing" when we really mean that we are measuring mass by comparing two weights. **Figure 0.1** shows two types of balances normally used for measuring mass in the laboratory.

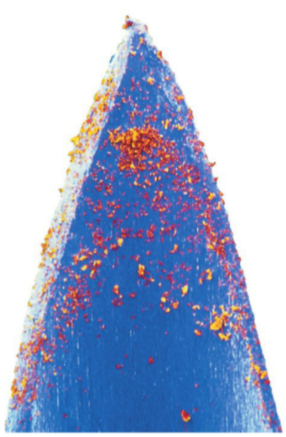
FIGURE 0.1

Some balances used for measuring mass in the laboratory.

0.4 FUNDAMENTAL UNITS: MEASURING LENGTH

The **meter (m)** is the standard unit of length in the SI system. Although originally defined in 1790 as being 1 ten-millionth of the distance from the equator to the North Pole, the meter was redefined in 1889 as the distance between two thin lines on a bar of platinum–iridium alloy stored near Paris, France. To accommodate an increasing need for precision, the meter was redefined again in 1983 as equal to the distance traveled by light through a vacuum in 1/299,792,458 second. Although this new definition isn't as easy to grasp as the distance between two scratches on a bar, it has the great advantage that it can't be lost or damaged.

One meter is 39.37 inches, about 10% longer than an English yard and much too large for most measurements in chemistry. Other, more commonly used measures of length are the **centimeter (cm**; 1 cm = 0.01 m, a bit less than half an inch), the **millimeter (mm**; 1 mm = 0.001 m, about the thickness of a U.S. dime), the **micrometer** (**μm**; 1 μm = 10^{-6} m), the **nanometer** (**nm**; 1 nm = 10^{-9} m), and the **picometer** (**pm**; 1 pm = 10^{-12} m). Thus, a chemist might refer to the diameter of a sodium atom as 372 pm (3.72×10^{-10} m).



▲ The bacteria on the tip of this pin have a length of about 5×10^{-7} m. How many nanometers is this?

1 m = 100 cm = 1000 mm = 1,000,000 μm = 1,000,000,000 nm (1.0936 yd)

1 cm = 10 mm = 10,000 μm = 10,000,000 nm (0.3937 in.)

1 mm = 1000 μm = 1,000,000 nm

0.5 FUNDAMENTAL UNITS: MEASURING TEMPERATURE

Just as the kilogram and the meter are slowly replacing the pound and the yard as common units for mass and length measurement in the United States, the **degree Celsius** (°C) is slowly replacing the degree Fahrenheit (°F) as the common unit for temperature measurement. In scientific work, however, the **kelvin (K)** has replaced both. (Note that we say only "kelvin," not "degree kelvin.")

For all practical purposes, the kelvin and the degree Celsius are the same size—both are one-hundredth of the interval between the freezing point of water and the boiling point of water at standard atmospheric pressure. The only real difference between the two units is that the numbers assigned to various points on the scales differ. Whereas the Celsius scale assigns a value of 0 °C to the freezing point of water and 100 °C to the boiling point of water, the Kelvin scale assigns a value of 0 K to the coldest possible temperature, −273.15 °C, sometimes called *absolute zero*. Thus, 0 K = −273.15 °C and 273.15 K = 0 °C. For example, a warm spring day with a Celsius temperature of 25 °C has a Kelvin temperature of 25° + 273.15° = 298 K.

$$\text{Temperature in K} = \text{Temperature in °C} + 273.15°$$
$$\text{Temperature in °C} = \text{Temperature in K} - 273.15°$$

In contrast to the Kelvin and Celsius scales, the common Fahrenheit scale specifies an interval of 180° between the freezing point (32 °F) and the boiling point (212 °F) of water. Thus, it takes 180 degrees Fahrenheit to cover the same range as 100 degrees Celsius (or kelvins) so a degree Fahrenheit is only 100/180 = 5/9 as large as a degree Celsius. **Figure 0.2** compares the Fahrenheit, Celsius, and Kelvin scales.

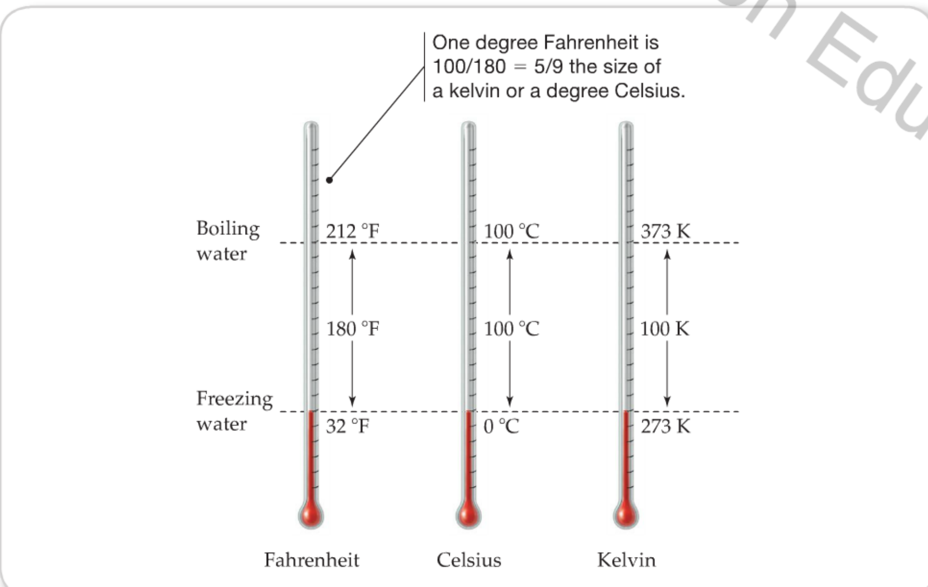


FIGURE 0.2
A comparison of the Fahrenheit, Celsius, and Kelvin temperature scales.

Two adjustments are needed to convert between Fahrenheit and Celsius scales—one to adjust for the difference in degree size and one to adjust for the difference in zero points. The size adjustment is made using the relationships 1 °C = 9/5 °F and 1 °F = 5/9 °C. The zero-point adjustment is made by remembering that the freezing point of water is higher by 32° on the Fahrenheit scale than on the Celsius scale. Thus, if you want to convert from Celsius to Fahrenheit, you first change the size of the Celsius value to that of a Fahrenheit value (multiply °C by 9/5) and then adjust the zero point of that Fahrenheit value (add 32°). If you want to convert from Fahrenheit to Celsius, you first adjust the zero-point of the Fahrenheit value (by subtracting 32°) and then change the size of that Fahrenheit value to a Celsius value (multiply by 5/9). The following formulas describe the conversions, and Worked Example 0.1 shows how to do a calculation.

Celsius to Fahrenheit

$$°F = \left(\frac{9\,°F}{5\,\cancel{°C}} \times \cancel{°C}\right) + 32\,°F$$

Fahrenheit to Celsius

$$°C = \frac{5\,°C}{9\,\cancel{°F}} \times (\cancel{°F} - 32\,\cancel{°F})$$

Worked Example 0.1

Converting from Fahrenheit to Celsius

The melting point of table salt is 1474 °F. What temperature is this on the Celsius and Kelvin scales?

SOLUTION

There are two ways to do this and every other problem in chemistry. One is to plug numbers into a formula and hope for the best; the other is to think things through to be sure you understand what's going on. The formula approach works only if you use the right equation; the thinking approach always works. Let's try both ways.

The formula approach: Set up an equation using the temperature conversion formula for changing from Fahrenheit to Celsius:

$$°C = \left(\frac{5\,°C}{9\,\cancel{°F}}\right)(1474\,\cancel{°F} - 32\,\cancel{°F}) = 801\,°C$$

Converting to kelvin gives a temperature of 801° + 273.15° = 1074 K.

The thinking approach: We're given a temperature in degrees Fahrenheit, and we need to convert to degrees Celsius. A temperature of 1474 °F corresponds to 1474 °F − 32 °F = 1442 °F above the freezing point of water. Because a degree Fahrenheit is only 5/9 as large as a degree Celsius, 1442 degrees Fahrenheit above freezing equals 1442 × 5/9 = 801 degrees Celsius above freezing (0 °C), or 801 °C. The same number of degrees above freezing on the Kelvin scale (273.15 K) corresponds to a temperature of 273.15 + 801 = 1074 K.

Because the answers obtained by the two approaches agree, we can feel fairly confident that our thinking is following the right lines and that we understand the subject. (If the answers did *not* agree, we'd be alerted to a misunderstanding somewhere.)

▲ The melting point of sodium chloride is 1474 °F, or 801 °C.

PROBLEM 0.3 The normal body temperature of a healthy adult human is 98.6 °F. What is this value on both Celsius and Kelvin scales?

PROBLEM 0.4 Carry out the indicated temperature conversions.
(a) −78 °C = ? K (b) 158 °C = ? °F (c) 375 K = ? °F

0.6 DERIVED UNITS: MEASURING VOLUME

Look back at the seven fundamental SI units given in Table 0.1 and you'll find that measures for such familiar quantities as area, volume, density, speed, and pressure are missing. All are examples of *derived* quantities rather than fundamental quantities because they can be expressed by using a combination of one or more of the seven fundamental units (**Table 0.3**).

TABLE 0.3 Some Derived Units and the Quantities They Measure

Quantity	Definition	Derived Unit (Name)
Area	Length times length	m^2
Volume	Area times length	m^3
Density	Mass per unit volume	kg/m^3
Speed	Distance per unit time	m/s
Acceleration	Change in speed per unit time	m/s^2
Force	Mass times acceleration	$(kg \cdot m)/s^2$ (newton, N)
Pressure	Force per unit area	$kg/(m \cdot s^2)$ (pascal, Pa)
Energy	Force times distance	$(kg \cdot m^2)/s^2$ (joule, J)

Volume, the amount of space occupied by an object, is measured in SI units by the **cubic meter (m^3)**, defined as the amount of space occupied by a cube 1 meter on edge (**Figure 0.3**).

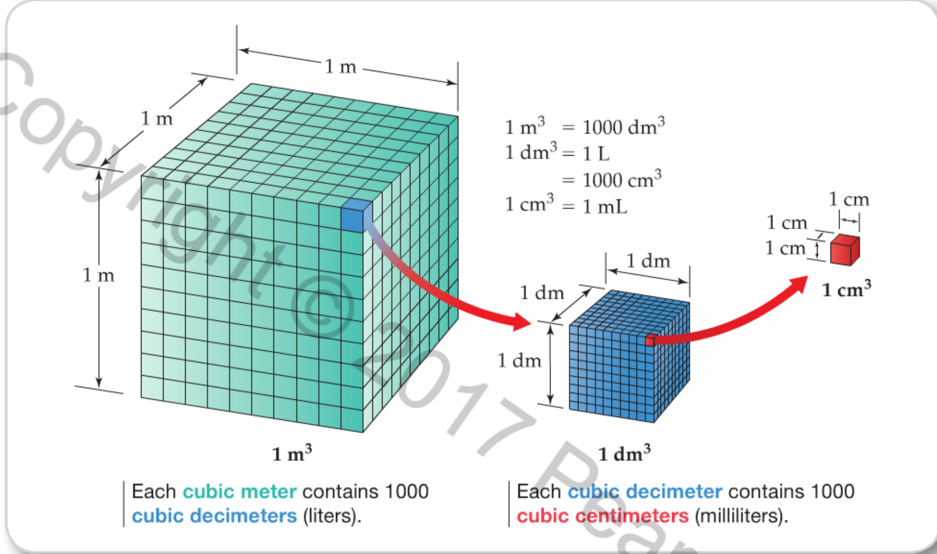


FIGURE 0.3

Units for measuring volume. A cubic meter is the volume of a cube 1 meter along each edge.

A cubic meter equals 264.2 U.S. gallons, much too large a quantity for normal use in chemistry. As a result, smaller, more convenient measures are commonly employed. Both the **cubic decimeter** ($1\ \mathbf{dm^3} = 0.001\ m^3$), equal in size to the more familiar metric **liter (L)**, and the **cubic centimeter** ($1\ \mathbf{cm^3} = 0.001\ dm^3 = 10^{-6}\ m^3$), equal in size to the metric **milliliter (mL)**, are particularly convenient. Slightly larger than 1 U.S. quart, a liter has the volume of a cube 1 dm on edge. Similarly, a milliliter has the volume of a cube 1 cm on edge (Figure 0.3).

$$1\ m^3 = 1000\ dm^3 = 1{,}000{,}000\ cm^3 \qquad (264.2\ \text{gal})$$
$$1\ dm^3 = 1\ L = 1000\ mL \qquad (1.057\ \text{qt})$$

Figure 0.4 shows some of the equipment frequently used in the laboratory for measuring liquid volume.

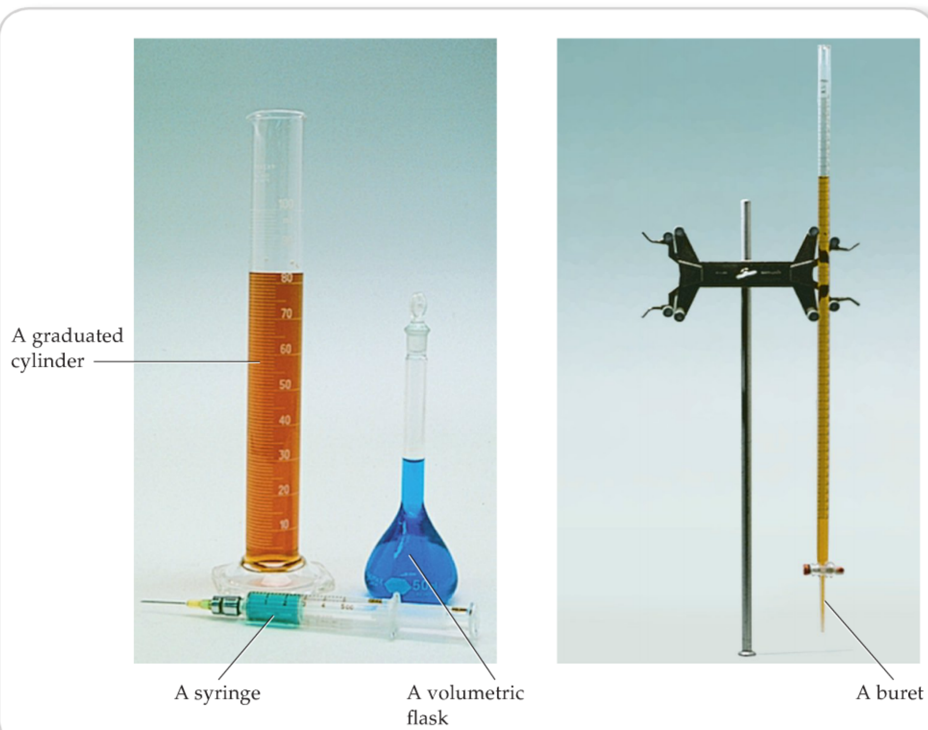


FIGURE 0.4

Common items of laboratory equipment used for measuring liquid volume.

0.7 DERIVED UNITS: MEASURING DENSITY

▲ Which weighs more, the brass weight or the pillow? Actually, both have identical masses and weights, but the brass weight has a higher density because its volume is much smaller.

The property that relates the mass of an object to its volume is called *density*. **Density**, which is simply the mass of an object divided by its volume, is expressed in the SI derived unit g/mL for a liquid or g/cm^3 for a solid. The densities of some common materials are given in **Table 0.4**.

$$\text{Density} = \frac{\text{Mass (g)}}{\text{Volume (mL or cm}^3\text{)}}$$

TABLE 0.4 Densities of Some Common Materials

Substance	Density (g/cm^3)	Substance	Density (g/cm^3)
Ice (0 °C)	0.917	Human fat	0.94
Water (3.98 °C)	1.0000	Human muscle	1.06
Gold	19.31	Cork	0.22–0.26
Helium (25 °C)	0.000 164	Balsa wood	0.12
Air (25 °C)	0.001 185	Earth	5.54

Because most substances change in volume when heated or cooled, densities are temperature-dependent. At 3.98 °C, for instance, a 1.0000 mL container holds exactly 1.0000 g of water (density = 1.0000 g/mL). As the temperature is raised, however, the volume occupied by the water expands so that only 0.9584 g fits in the 1.0000 mL container at 100 °C (density = 0.9584 g/mL). When reporting a density, the temperature must also be specified.

Although most substances expand when heated and contract when cooled, water behaves differently. Water contracts when cooled from 100 °C to 3.98 °C, but below this temperature it begins to expand again. Thus, the density of liquid water is at its maximum of 1.0000 g/mL at 3.98 °C but decreases to 0.999 87 g/mL at 0 °C (**Figure 0.5**). When freezing occurs, the density drops still further to a value of 0.917 g/cm^3 for ice at 0 °C. Ice and any other substance with a density less than that of water will float, but any substance with a density greater than that of water will sink.

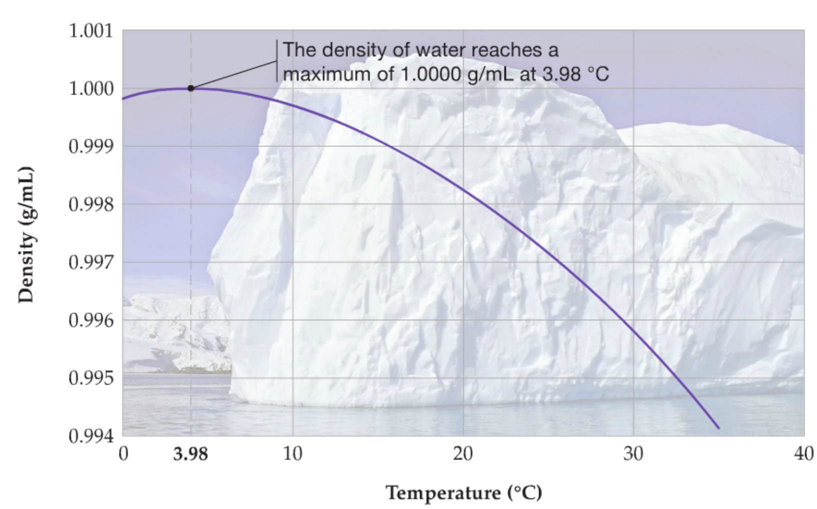


FIGURE 0.5

The density of water at different temperatures.

Knowing the density of a substance, particularly a liquid, can be useful in the chemistry laboratory because it's often easier to measure a liquid amount by volume rather than by mass. Suppose, for example, that you needed 1.55 g of ethyl alcohol. Rather than trying to weigh exactly the right amount of the liquid, it would be much easier to look up the density of ethyl alcohol in a standard reference book such as *The Handbook of Chemistry and Physics* (0.7893 g/mL at 20 °C) and then measure the correct volume with a syringe.

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} \qquad \text{so} \qquad \text{Volume} = \frac{\text{Mass}}{\text{Density}}$$

$$\text{Volume} = \frac{1.55\ \cancel{g}\ \text{ethyl alcohol}}{0.7893\ \dfrac{\cancel{g}}{\text{mL}}} = 1.96\ \text{mL ethyl alcohol}$$

▲ The mass of a liquid is easily measured with a syringe if the density of the liquid is known.

Worked Example 0.2

Calculating a Density

What is the density of the element copper in g/cm^3 if a sample weighing 324.5 g has a volume of 36.2 cm^3?

SOLUTION

Density is mass divided by volume:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} = \frac{324.5\ \text{g}}{36.2\ \text{cm}^3} = 8.96\ \text{g/cm}^3$$

Worked Example 0.3

Using Density to Calculate a Volume

What is the volume in cm^3 of 454 g of gold? (See Table 0.4.)

STRATEGY AND SOLUTION

As described in the text, density is mass divided by volume, so volume is mass divided by density:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}} \qquad \text{so} \qquad \text{Volume} = \frac{\text{Mass}}{\text{Density}}$$

$$\text{Volume} = \frac{454\ \cancel{g}\ \text{gold}}{19.31\ \cancel{g}/\text{cm}^3} = 23.5\ \text{cm}^3\ \text{gold}$$

PROBLEM 0.5 What is the density of glass in g/cm^3 if a sample weighing 27.43 g has a volume of 12.40 cm^3?

PROBLEM 0.6 Chloroform, a substance once used as an anesthetic, has a density of 1.483 g/mL at 20 °C. How many milliliters would you use if you needed 9.37 g?

0.8 DERIVED UNITS: MEASURING ENERGY

The last derived unit we'll consider at this point is **energy**, a word familiar to everyone though a bit hard to define without a lengthy explanation. We'll wait until Chapter 8 to give that explanation and discuss the topic more fully but will introduce it now because energy is so important in chemistry and shows up in so many contexts.

Energy has the unit of mass (kg) times the square of velocity (m/s)^2, or (kg · m^2)/s^2, an SI derived unit that is given the name **joule (J)** after the English physicist James Prescott Joule (1818–1889). The joule is a fairly small amount of energy—it takes roughly 100,000 J to heat a coffee cup full of water from room temperature to boiling—so kilojoules (kJ) are more frequently used in chemistry.

In addition to the SI energy unit joule, some chemists and biochemists still use the unit calorie (cal, with a lowercase c). Originally defined as the amount of energy necessary to raise

LOOKING AHEAD ...

The simplest definition of **energy** is to say that it is the capacity to supply heat or do work. All of Chapter 8 is devoted to explaining this topic and to seeing the importance of energy relationships in chemistry.

▲ The 75 watt incandescent bulb in this lamp uses energy at the rate of 75 J/s. Only about 5% of that energy appears as light, however; the remaining 95% is given off as heat.

the temperature of 1 g of water by 1 °C (specifically, from 14.5 °C to 15.5 °C), 1 calorie is now defined as exactly 4.184 J.

$$1 \text{ cal} = 4.184 \text{ J} \qquad \text{and} \qquad 1 \text{ J} = 0.2390 \text{ cal}$$

Nutritionists also use the somewhat confusing unit Calorie (Cal, with a capital C), which is equal to 1000 calories, or 1 kilocalorie (kcal), or 4.184 kilojoules (kJ).

$$1 \text{ Cal} = 1000 \text{ cal} = 1 \text{ kcal} = 4.184 \text{ kJ}$$

The energy value, or caloric content, of food is measured in Calories. Thus, the statement that a banana contains 70 Calories means that 70 Cal (70 kcal, or 290 kJ) of energy is released when the banana is used by the body for fuel.

PROBLEM 0.7 A Big Mac hamburger from McDonald's contains 540 Calories.
(a) How many kilojoules does a Big Mac contain?
(b) For how many hours could the amount of energy in a Big Mac light a 100 watt lightbulb? (1 watt = 1 J/s)

0.9 ACCURACY, PRECISION, AND SIGNIFICANT FIGURES IN MEASUREMENT

Making measurements like those discussed in the previous sections is something that most of us do every day, whether in cooking, carpentry, or chemistry. But how good are those measurements? Any measurement is only as good as the skill of the person doing the work and the reliability of the equipment being used. You've probably noticed, for instance, that you often get slightly different readings when you weigh yourself on a common bathroom scale and on a scale at the doctor's office, so there's always some uncertainty about your real weight. The same is true in chemistry—there is always some uncertainty in the value of a measurement.

In talking about the degree of uncertainty in a measurement, we use the words *accuracy* and *precision*. Although most of us use the words interchangeably in daily life, there's actually an important distinction between them. **Accuracy** refers to how close to the true value a given measurement is, whereas **precision** refers to how well a number of independent measurements agree with one another. To see the difference, imagine that you weigh a tennis ball whose true mass is 54.441 778 g. Assume that you take three independent measurements on each of three different types of balance to obtain the data shown in the following table.

▲ This tennis ball has a mass of about 54 g.

Measurement #	Bathroom Scale	Lab Balance	Analytical Balance
1	0.1 kg	54.4 g	54.4418 g
2	0.0 kg	54.5 g	54.4417 g
3	0.1 kg	54.3 g	54.4418 g
(average)	(0.07 kg)	(54.4 g)	(54.4418 g)

If you use a bathroom scale, your measurement (average = 0.07 kg) is neither accurate nor precise. Its accuracy is poor because it measures to only one digit that is far from the true value, and its precision is poor because any two measurements may differ substantially. If you now weigh the ball on an inexpensive laboratory balance, the value you get (average = 54.4 g) has three digits and is fairly accurate, but it is still not very precise because the three readings vary from 54.3 g to 54.5 g, perhaps due to air movements in the room or to a sticky mechanism. Finally, if you weigh the ball on an expensive analytical balance like those found in research laboratories, your measurement (average = 54.4418 g) is both precise and accurate. It's accurate because the measurement is very close to the true value, and it's precise because it has six digits that vary little from one reading to another.

To indicate the uncertainty in a measurement, the value you record should use all the digits you are sure of plus one additional digit that you estimate. In reading a thermometer that has a mark for each degree, for example, you could be certain about the digits of the nearest mark—say 25 °C—but you would have to estimate between two marks—say between 25 °C and 26 °C—to obtain a value of 25.3 °C.

The total number of digits recorded for a measurement is called the measurement's number of **significant figures**. For instance, the mass of the tennis ball as determined on the single-pan balance (54.4 g) has three significant figures, whereas the mass determined on the analytical balance (54.4418 g) has six significant figures. All digits but the last are certain; the final digit is an estimate, which we generally assume to have an error of plus or minus one (±1).

Finding the number of significant figures in a measurement is usually easy but can be troublesome if zeros are present. Look at the following four quantities:

4.803 cm	Four significant figures: 4, 8, 0, 3
0.006 **61** g	Three significant figures: 6, 6, 1
55.220 K	Five significant figures: 5, 5, 2, 2, 0
34,200 m	Anywhere from three (3, 4, 2) to five (3, 4, 2, 0, 0) significant figures

The following rules cover the different situations that arise:

1. **Zeros in the middle of a number are like any other digit; they are always significant.** Thus, 4.803 cm has four significant figures.
2. **Zeros at the beginning of a number are not significant;** they act only to locate the decimal point. Thus, 0.006 61 g has three significant figures. (Note that 0.006 61 g can be rewritten as 6.61×10^{-3} g or as 6.61 mg.)
3. **Zeros at the end of a number and *after* the decimal point are always significant.** The assumption is that these zeros would not be shown unless they were significant. Thus, 55.220 K has five significant figures. (If the value were known to only four significant figures, we would write 55.22 K.)
4. **Zeros at the end of a number and *before* the decimal point may or may not be significant.** We can't tell whether they are part of the measurement or whether they just locate the decimal point. Thus, 34,200 m may have three, four, or five significant figures. Often, however, a little common sense is helpful. A temperature reading of 20 °C probably has two significant figures rather than one, since one significant figure would imply a temperature anywhere from 10 °C to 30 °C and would be of little use. Similarly, a volume given as 300 mL probably has three significant figures. On the other hand, a figure of 93,000,000 mi for the distance between the Earth and the Sun probably has only two or three significant figures.

The fourth rule shows why it's helpful to write numbers in scientific notation rather than ordinary notation: Doing so makes it possible to indicate the number of significant figures. Thus, writing the number 34,200 as 3.42×10^4 indicates three significant figures but writing it as 3.4200×10^4 indicates five significant figures.

One further point about significant figures: Certain numbers, such as those obtained when counting objects, are exact and have an effectively infinite number of significant figures. A week has exactly 7 days, for instance, not 6.9 or 7.0 or 7.1, and a foot has exactly 12 inches, not 11.9 or 12.0 or 12.1. In addition, the power of 10 used in scientific notation is an exact number. That is, the number 10^3 is exactly 1000, but the number 1×10^3 has one significant figure.

Worked Example 0.4

Significant Figures

How many significant figures does each of the following measurements have?

(a) 0.036 653 m **(b)** 7.2100×10^{-3} g **(c)** 72,100 km **(d)** $25.03

SOLUTION

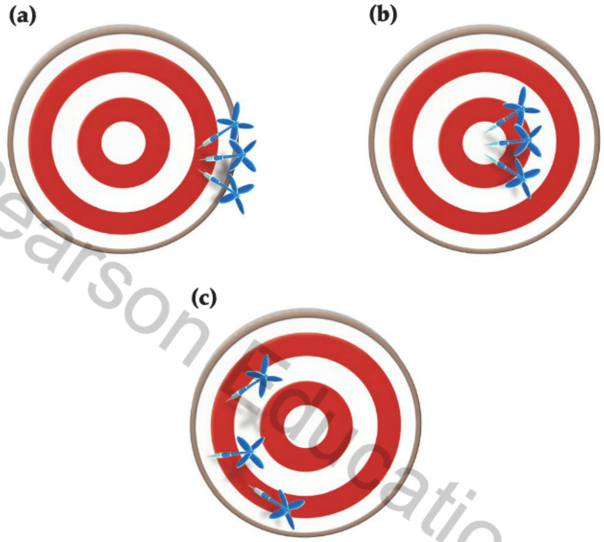
(a) 5 (by rule 2) **(b)** 5 (by rule 3) **(c)** 3, 4, or 5 (by rule 4)
(d) $25.03 is an exact number

PROBLEM 0.8 A 1.000 mL sample of acetone, a common solvent sometimes used as a paint remover, was placed in a small bottle whose mass was known to be 38.0015 g. The following values were obtained when the acetone-filled bottle was weighed: 38.7798 g, 38.7795 g, and 38.7801 g. How would you characterize the precision and accuracy of these measurements if the true mass of the acetone was 0.7791 g?

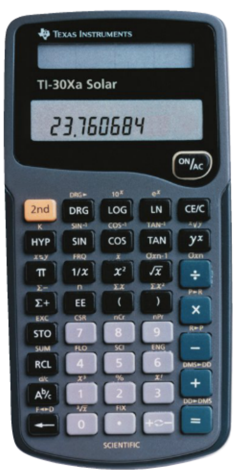
PROBLEM 0.9 How many significant figures does each of the following quantities have? Explain your answers.

(a) 76.600 kJ **(b)** 4.50200×10^{3} g **(c)** 3000 nm **(d)** 0.00300 mL
(e) 18 students **(f)** 3×10^{-5} g **(g)** 47.60 mL **(h)** 2070 mi

CONCEPTUAL PROBLEM 0.10 Characterize each of the following dartboards according to the accuracy and precision of the results.


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

0.10 ROUNDING NUMBERS



▲ Calculators often display more figures than are justified by the precision of the data.

It often happens when dealing with measurements, particularly when doing arithmetic on a calculator, that a quantity appears to have more significant figures than are really justified. You might calculate the gas mileage of your car, for instance, by finding that it takes 11.70 gallons of gasoline to drive 278 miles:

$$\text{Mileage} = \frac{\text{Miles}}{\text{Gallons}} = \frac{278\text{ mi}}{11.70\text{ gal}} = 23.760684\text{ mi/gal (mpg)}$$

Although the answer on the calculator has eight digits, your measurement is really not as precise as it appears. In fact, your answer is precise to only three significant figures and should be **rounded off** to 23.8 mi/gal by removing all nonsignificant digits.

How do you decide how many figures to keep and how many to remove? For most purposes, a simple procedure using just two rules is sufficient.

1. **In carrying out a multiplication or division, the answer can't have more significant figures than either of the original numbers.** If you think about it, this rule is just common sense. After all, if you don't know the number of miles you drove to better than three significant figures (278 could mean 277, 278, or 279), you can't calculate your mileage to more than the same number of significant figures.

Three significant figures (278) — Three significant figures (23.8)

$$\frac{278\text{ mi}}{11.70\text{ gal}} = 23.8\text{ mi/gal}$$

Four significant figures (11.70)

2. **In carrying out an addition or subtraction, the answer can't have more digits to the right of the decimal point than either of the original numbers.** For example, if you have 3.18 L of water and you add 0.013 15 L more, you now have 3.19 L. Again, this rule is just common sense. If you don't know the volume you started with past the second decimal place (it could be 3.17, 3.18, or 3.19), you can't know the total of the combined volumes past the same decimal place.

$$\begin{array}{r} 3.18?\ ?? \\ +\ 0.013\ 15 \\ \hline 3.19?\ ?? \end{array}$$

3.18? ?? ← Ends **two** places past decimal point
+ 0.013 15 ← Ends **five** places past decimal point
3.19? ?? ← Ends **two** places past decimal point

Once you decide how many digits to retain for your answer, the rules for rounding off numbers are straightforward:

1. **If the first digit you remove is less than 5, round down by dropping it and all following digits.** Thus, 3.864 becomes 3.86 when rounded to three significant figures because the first (and only) of the dropped digits is less than 5.
2. **If the first digit you remove is 5 or greater, round up by adding 1 to the digit on the left.** Thus, 3.865 becomes 3.87 when rounded to three significant figures because the first (and only) of the dropped digits is 5.

These two rules arise because of the nature of the decimal system, in which the first digit dropped in rounding can have any of ten values, 0 through 9. By rounding down in the five cases when the first dropped digit is 0, 1, 2, 3, or 4 and rounding up in the five cases when the first dropped digit is 5, 6, 7, 8, or 9, rounding up and rounding down have the same 50% chance of occurring,

Worked Example 0.5

A Calculation Using Significant Figures

It takes 9.25 hours to fly from London, England, to Chicago, Illinois, a distance of 3952 miles. What is the average speed of the airplane in miles per hour?

SOLUTION

First, set up an equation dividing the number of miles flown by the number of hours:

$$\text{Average speed} = \frac{3952 \text{ mi}}{9.25 \text{ h}} = 427.243\,24 \text{ mi/h}$$

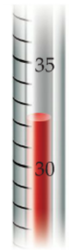
Next, decide how many significant figures should be in your answer. Because the problem involves a division, and because one of the quantities you started with (9.25 h) has only three significant figures, the answer must also have three significant figures. Finally, round off your answer. The first digit to be dropped (2) is less than 5, so the answer 427.243 24 must be rounded off to 427 mi/h.

In doing this or any other problem, use all figures for the calculation, whether significant or not, and then round off the final answer. Don't round off at any intermediate step.

PROBLEM 0.11 Round off each of the following quantities to the number of significant figures indicated in parentheses:
(a) 3.774 499 L (4) **(b)** 255.0974 K (3)
(c) 55.265 kg (4) **(d)** 906.401 kJ (5)

PROBLEM 0.12 Carry out the following calculations, expressing each result with the correct number of significant figures:
(a) 24.567 g + 0.044 78 g = ? g
(b) 4.6742 g ÷ 0.003 71 L = ? g/L
(c) 0.378 mL + 42.3 mL − 1.5833 mL = ? mL

CONCEPTUAL PROBLEM 0.13 What is the temperature reading on the following Celsius thermometer? How many significant figures do you have in your answer?



0.11 CONVERTING MEASUREMENTS FROM ONE UNIT TO ANOTHER

▲ Speed skaters have to convert from laps to meters to find out how far they have gone.

Because so many scientific activities involve numerical calculations—measuring, weighing, preparing solutions, and so forth—it's often necessary to convert a quantity from one unit to another. Converting between units isn't difficult; we all do it every day. If you run 7.5 laps around a 200 meter track, for instance, you have to convert between the distance unit *lap* and the distance unit *meter* to find that you have run 1500 m (7.5 laps times 200 meters per lap). Converting from one scientific unit to another is just as easy.

$$7.5\ \cancel{\text{laps}} \times \frac{200\ \text{meters}}{1\ \cancel{\text{lap}}} = 1500\ \text{meters}$$

The simplest way to carry out calculations that involve different units is to use the **dimensional-analysis method**. In this method, a quantity described in one unit is converted into an equivalent quantity with a different unit by using a **conversion factor** to express the relationship between units.

$$\text{Original quantity} \times \text{Conversion factor} = \text{Equivalent quantity}$$

As an example, we know from Section 0.4 that 1 meter equals 39.37 inches. Writing this relationship as a fraction restates it in the form of a conversion factor, either meters per inch or inches per meter.

$$\textbf{Conversion factors between meters and inches} \quad \frac{1\ \text{m}}{39.37\ \text{in.}} \quad \text{equals} \quad \frac{39.37\ \text{in.}}{1\ \text{m}} \quad \text{equals} \quad 1$$

Note that this and all other conversion factors are effectively equal to 1 because the quantity above the division line (the numerator) is equal in value to the quantity below the division line (the denominator). Thus, multiplying by a conversion factor is equivalent to multiplying by 1 and so does not change the value of the quantity.

The key to the dimensional-analysis method of problem solving is that units are treated like numbers and can thus be multiplied and divided just as numbers can. The idea when solving a problem is to set up an equation so that unwanted units cancel, leaving only the desired units. Usually it's best to start by writing what you know and then manipulating that known quantity. For example, if you know your height is 69.5 inches and want to find it in meters, you can write down the height in inches and set up an equation multiplying the height by the conversion factor in meters per inch. The unit "in." cancels from the equation because it appears both above and below the division line, and the only unit that remains is “m.”

$$\underset{\text{Starting quantity}}{69.5\ \cancel{\text{in.}}} \times \underset{\text{Conversion factor}}{\frac{1\ \text{m}}{39.37\ \cancel{\text{in.}}}} = \underset{\text{Equivalent quantity}}{1.77\ \text{m}}$$

The dimensional-analysis method gives the right answer only if the equation is set up so that the unwanted units cancel. If the equation is set up in any other way, the units won’t cancel properly and you won’t get the right answer. Thus, if you were to multiply your height in inches by the incorrect conversion factor inches per meter, you would end up with a meaningless answer using meaningless units.

Wrong! $69.5\ \text{in.} \times \frac{39.37\ \text{in.}}{1\ \text{m}} = 2740\ \text{in.}^2/\text{m}$ **??**

The main drawback to using the dimensional-analysis method is that it’s easy to get the right answer without really understanding what you’re doing. It’s therefore best after solving a problem to think through a rough estimate, or “ballpark” solution, as a check on your work. If your ballpark check isn’t close to the answer you get from the detailed solution, there’s a misunderstanding somewhere and you should think through the problem again.

Even if you don’t make an estimate, it’s important to be sure that your calculated answer makes sense. If, for example, you were trying to calculate the volume of a human cell and you came up with the answer 5.3 cm^3, you should realize that such an answer couldn’t possibly be right. Cells are too tiny to be distinguished with the naked eye, but a volume of 5.3 cm^3 is about the size of a walnut.

The dimensional-analysis method and the use of ballpark checks are techniques that will help you solve problems of many kinds, not just unit conversions. Problems sometimes seem complicated, but you can usually sort out the complications by analyzing the problem properly.

- Identify the information you’re given, including units.
- Identify the information you need in the answer, including units.
- Find a relationship between the known information and needed answer, and plan a strategy for getting from one to the other.
- Solve the problem.
- Make a rough estimate to be sure your calculated answer is reasonable.

▲ What is the volume of a red blood cell?

Examples 0.6–0.9 show how to devise strategies and estimate answers. To conserve space, we’ll use this approach routinely in only the next few chapters and more sparingly thereafter, but you should make it a standard part of all your problem solving.

Worked Example 0.6

Unit Conversions and Significant Figures

The Bugatti Veyron Super Sport is currently the fastest production car in the world, with a measured top speed of 267 miles per hour (and a price of $2.4 million). What is this speed in kilometers per hour?

STRATEGY

The known information is the speed in mi/h; the unknown is the speed in km/h. Find the appropriate conversion factor inside the back cover of this book, and use the dimensional-analysis method to set up an equation so the “mi” units cancel. The answer should be rounded to three significant figures.

SOLUTION

$$\frac{267\ \cancel{\text{mi}}}{1\ \text{h}} \times \frac{1.609\ \text{km}}{1\ \cancel{\text{mi}}} = 430\ \frac{\text{km}}{\text{h}}$$

A very fast car!

▲ The Bugatti Veyron Super Sport has a top speed of 267 mph.

continued on next page

✔ BALLPARK CHECK

The answer is certainly large, perhaps several hundred km/h. A better estimate is to realize that, because 1 mi = 1.609 km, it takes a bit more than 1 1/2 times as many kilometers as miles to measure the same distance. Thus, 267 mi is a bit more than 400 km, and 267 mi/h is over 400 km/h. The estimate agrees with the detailed solution.

Worked Example 0.7

Complex Unit Conversions

A large sport utility vehicle moving at a speed of 125 km/h might use gasoline at a rate of 16 L per 100 km. What does this correspond to in mi/gal?

STRATEGY

We are given a gasoline mileage in the units L/km (or km/L), and we need to find the mileage in the units mi/gal. Thus, two conversions are necessary, one from kilometers to miles and one from liters to gallons. It's best to do multiple conversions one step at a time until you get used to them. First, convert the distance from kilometers to miles and the amount of fuel from liters to gallons, and then divide the distance by the amount of fuel to find the mileage.

SOLUTION

$$100\ \text{km} \times \frac{0.6214\ \text{mi}}{1\ \text{km}} = 62.14\ \text{mi} \qquad 16\ \text{L} \times \frac{1\ \text{gal}}{3.79\ \text{L}} = 4.22\ \text{gal}$$

$$\frac{62.14\ \text{mi}}{4.22\ \text{gal}} = 14.73\ \frac{\text{mi}}{\text{gal}} \qquad \text{Round off to 15 mi/gal}$$

Notice that extra digits are carried through the intermediate calculations and only the final answer is rounded off.

When you become more confident in working multiple conversion problems, you can set up one large equation in which all unwanted units cancel.

$$\frac{100\ \text{km}}{16\ \text{L}} \times \frac{3.79\ \text{L}}{1\ \text{gal}} \times \frac{0.6214\ \text{mi}}{1\ \text{km}} = 14.73\ \frac{\text{mi}}{\text{gal}} \qquad \text{Round off to 15 mi/gal}$$

✔ BALLPARK CHECK

The mileage is probably low, perhaps around 15 mi/gal. This is a difficult problem to estimate, however, because it requires several different conversions. It's therefore best to think the problem through one step at a time, writing down the intermediate estimates:

- A distance of 100 km per 16 L is approximately 6 km/L.
- Because 1 km is about 0.6 mi, 6 km/L is about 4 mi/L.
- Because 1 L is approximately 1 qt, or 1/4 gal, 4 mi/L is about 16 mi/gal.

This estimate agrees with the detailed solution.

Worked Example 0.8

Complex Unit Conversions

The volcanic explosion that destroyed the Indonesian island of Krakatau on August 27, 1883, released an estimated 4.3 cubic miles (mi^3) of debris into the atmosphere and affected global weather for years. In SI units, how many cubic meters (m^3) of debris were released?

STRATEGY

We are given a volume in cubic miles and need to convert to cubic meters. It's probably simplest to convert first from mi^3 to km^3 and then from km^3 to m^3.

SOLUTION

$$4.3\ \text{mi}^3 \times \left(\frac{1\ \text{km}}{0.6214\ \text{mi}}\right)^3 = 17.92\ \text{km}^3$$

$$17.92\ \text{km}^3 \times \left(\frac{1000\ \text{m}}{1\ \text{km}}\right)^3 = 1.792 \times 10^{10}\ \text{m}^3$$

$$= 1.8 \times 10^{10}\ \text{m}^3 \quad \text{Rounded off}$$

✔ BALLPARK CHECK

As in the previous Worked Example, this is a difficult problem to estimate because it requires several different conversions. Take it one step at a time: One meter is much less than 1 mile, so it takes a large number of cubic meters to equal 1 mi^3, and the answer is going to be very large. Because 1 km is about 0.6 mi, 1 km^3 is about $(0.6)^3 = 0.2$ times as large as 1 mi^3. Thus, each mi^3 contains about 5 km^3, and 4.3 mi^3 contains about 20 km^3. Each km^3, in turn, contains $(1000\text{ m})^3 = 10^9\text{ m}^3$. The volume of debris from the Krakatau explosion was therefore about $20 \times 10^9\text{ m}^3$, or $2 \times 10^{10}\text{ m}^3$. The estimate agrees with the detailed solution.

Worked Example 0.9

Complex Unit Conversions

Precious metals such as gold, silver, and platinum are measured in *troy ounces* (ozt) rather than common avoirdupois ounces (oz), where 1 ozt = 31.103 g and 1 oz = 28.35 g. Furthermore, a troy pound (lbt) contains only 12 ozt while an avoirdupois pound (lb) contains 16 oz. How many ozt are in 1 lb, and how many oz are in 1 lbt? Which would you rather receive as a gift from a rich relative, 1 avoirdupois ounce of gold or 1 troy ounce? One avoirdupois pound of gold or 1 troy pound?

▲ Gold and other precious metals are measured in troy ounces rather than common avoirdupois ounces.

STRATEGY

This problem contains a lot of information that makes it sound more difficult than it is. You can cut through the difficulties, however, by setting up equations in which unwanted units cancel and only desired units remain.

SOLUTION

$$\frac{1\text{ ozt}}{31.103\text{ g}} \times \frac{28.35\text{ g}}{1\text{ oz}} \times \frac{16\text{ oz}}{1\text{ lb}} = 14.58\ \frac{\text{ozt}}{\text{lb}}$$

$$\frac{12\text{ ozt}}{1\text{ lbt}} \times \frac{31.103\text{ g}}{1\text{ ozt}} \times \frac{1\text{ oz}}{28.35\text{ g}} = 13.17\ \frac{\text{oz}}{\text{lbt}}$$

One troy ounce of gold (31.103 g) is larger than one avoirdupois ounce (28.35 g) and makes a nicer gift, but one avoirdupois pound (16 oz) is much larger than one troy pound (13.17 oz) and makes a great gift.

✔ BALLPARK CHECK

One troy ounce (31.103 g) is slightly larger than one avoirdupois ounce (28.35 g) so there will be slightly fewer of them in an avoirdupois pound, perhaps about 15 ozt/lb rather than 16 oz/lb. Furthermore, because an avoirdupois ounce is slightly smaller than a troy ounce, there will be slightly more of them in a troy pound, perhaps about 13 oz/lbt rather than 12 ozt/lbt. These estimates agree with the detailed solution.

▲ At 44.4 carats, the Hope Diamond is the largest blue diamond in the world.

PROBLEM 0.14 Calculate answers to the following problems, and check your solutions by making ballpark estimates.

(a) The melting point of gold is 1064 °C. What is this temperature in degrees Fahrenheit?

(b) How large, in cubic centimeters, is the volume of a red blood cell if the cell has a cylindrical shape with a diameter of 6×10^{-6} m and a height of 2×10^{-6} m?

PROBLEM 0.15 Gemstones are weighed in *carats*, with 1 carat = 200 mg (exactly). What is the mass in grams of the Hope Diamond, the world's largest blue diamond at 44.4 carats? What is this mass in ounces?

PROBLEM 0.16 A pure diamond with a mass of 0.1000 g contains 5.014×10^{21} carbon atoms and has a density of 3.52 g/cm^3. What is the volume of the Hope Diamond (Problem 0.15), and how many carbon atoms does it contain?

▲ Is this a poison or a treatment for leukemia?

FYI THE RISKS AND BENEFITS OF CHEMICALS

Life is not risk-free—we all take many risks each day, often without even thinking about it. We may decide to ride a bike rather than drive, even though the likelihood per mile of being killed on a bicycle is 10 times greater than in a car. We may decide to smoke cigarettes, even though smoking kills more than 170,000 people each year in the United States.

What about risks from "chemicals"? News reports sometimes make it seem that our food is covered with pesticides and filled with dangerous additives, that our land is polluted by toxic waste dumps, and that our medicines are unsafe. How bad are the risks from chemicals, and how are the risks evaluated?

First, it's important to realize that *everything*, including your own body, is made of chemicals—that's what matter is. There is no such thing as a "chemical-free" food, cosmetic, cleanser, or anything else. Second, there is no meaningful distinction between a "natural" substance and a "synthetic" one; a chemical is a chemical. Many naturally occurring substances—snake venom and botulism toxin, for example—are extraordinarily poisonous, and many synthetic substances—polyethylene and nylon, for example—are harmless.

Risk evaluation of chemicals is carried out by exposing test animals, usually mice or rats, to a chemical and then monitoring for signs of harm. To limit the expense and time needed for testing, the amounts administered are often hundreds or thousands of times larger than those a person might normally encounter. The *acute chemical toxicity* (as opposed to chronic toxicity) observed in animal tests is reported as an *LD_{50} value*, the amount of a substance per kilogram of body weight that is a lethal dose for 50% of the test animals. Some LD_{50} values of different substances are shown in **Table 0.5**. The lower the value, the more toxic the substance.

TABLE 0.5 Some LD_{50} Values in Rats

Substance	LD_{50} (g/kg)	Substance	LD_{50} (g/kg)
Strychnine	0.005	Chloroform	1.2
Arsenic trioxide	0.015	Iron(II) sulfate	1.5
DDT	0.115	Ethyl alcohol	7.1
Aspirin	1.1	Sodium cyclamate	12.8

Even with an LD_{50} value established in test animals, the risk of human exposure to a given substance is still hard to assess. If a substance is harmful to rats, is it necessarily harmful to humans? How can a large dose for a small animal be translated into a small dose for a large human? All substances are toxic to some organisms to some extent, and the difference between help and harm is often a matter of degree. Vitamin A, for example, is necessary for vision, yet it can promote cancer at high doses. Arsenic trioxide is the most classic of poisons, yet it induces remissions in some types of leukemia and was approved in 2000 by the U.S. Food and Drug Administration for drug use under the name Trisenox.

Even water can be toxic if drunk in large amounts because it dilutes the salt in body fluids and causes a potentially life-threatening condition called *hyponatremia* that has resulted in the death of several marathon runners. Furthermore, how we evaluate risk is strongly influenced by familiarity. Many foods contain natural ingredients far more toxic than synthetic additives or pesticide residues, but the ingredients are ignored because the foods are familiar.

All decisions involve tradeoffs. Does the benefit of a pesticide that increases food production outweigh a possible health risk to 1 person in 1 million? Do the beneficial effects of a new drug outweigh a potentially dangerous side effect in a small number of users? Different people will have different opinions, but an honest evaluation of the facts is surely better than a purely emotional response.

PROBLEM 0.17 Oxalic acid, found in dark-green leafy vegetables such as spinach and rhubarb, has a reported LD_{50} of 600 mg/kg in humans. If dry rhubarb leaves (not stalks) contain 0.5% oxalic acid by weight, how many pounds of dry leaves would a 175-pound person have to consume to have a 50% chance of dying? How many pounds of undried leaves, whose mass is about 95% water?

SUMMARY

Chemistry is the study of the composition, properties, and transformations of **matter**. It is best approached by posing questions, conducting experiments, and devising **theories** to interpret the experimental results.

Accurate measurement is crucial to scientific experimentation. The units used are those of the Système Internationale (**SI units**). There are seven fundamental SI units, together with other derived units. **Mass**, the amount of matter in an object, is measured in **kilograms (kg)**; **length** is measured in **meters (m)**; **temperature** is measured in **kelvin (K)**; and **volume** is measured in **cubic meters (m^3)**. The more familiar metric **liter (L)** and **milliliter (mL)** are also still used for measuring volume, and the **Celsius degree (°C)** is still used for measuring temperature. **Density** relates mass to volume and is measured in the derived SI unit g/cm^3 or g/mL. **Energy** is a quantity used throughout chemistry and is measured in the derived SI unit $(kg \cdot m^2)/s^2$, or **joule (J)**.

Because many experiments involve numerical calculations, it's often necessary to manipulate and convert different units of measure. The simplest way to carry out such conversions is to use the **dimensional-analysis method**, in which an equation is set up so that unwanted units cancel and only the desired units remain. It's also important when measuring physical quantities or carrying out calculations to indicate the precision of the measurement by **rounding off** the result to the correct number of **significant figures**.

KEY WORDS

accuracy *12*
Celsius degree (°C) *7*
centimeter (cm) *6*
chemistry *2*
conversion factor *16*
cubic centimeter (cm^3) *9*
cubic decimeter (dm^3) *9*
cubic meter (m^3) *9*
density *10*
dimensional-analysis method *16*
energy *11*
gram (g) *5*
joule (J) *11*
kelvin (K) *7*
kilogram (kg) *5*
liter (L) *9*
mass *5*
matter *5*
meter (m) *6*
microgram (μg) *5*
micrometer (μm) *6*
milligram (mg) *5*
milliliter (mL) *9*
millimeter (mm) *6*
nanometer (nm) *6*
picometer (pm) *6*
precision *12*
rounding off *14*
scientific notation *5*
SI unit *4*
significant figure *13*
theory *4*
volume *9*

KEY EQUATIONS

- Converting between Celsius and Fahrenheit temperatures **(Section 0.5)**

CELSIUS TO FAHRENHEIT

$$°F = \left(\frac{9\,°F}{5\,\cancel{°C}} \times \cancel{°C}\right) + 32\,°F$$

FAHRENHEIT TO CELSIUS

$$°C = \frac{5\,°C}{9\,\cancel{°F}} \times (\cancel{°F} - 32\,\cancel{°F})$$

LEARNING OUTCOMES

- Express and do calculations with numbers using scientific notation. **(Section 0.2)** ***Problems 1, 34, 35, 40–45, 48–51, 67***
- Learn the fundamental SI units for physical quantities and the prefixes that modify them by size. **(Section 0.2)** ***Problems 2, 22–29, 32, 33, 39***
- Use and interconvert SI and non-SI units for mass and length. **(Sections 0.3, 0.4)** ***Problems 35, 36, 66–73, 81, 93***
- Interconvert temperatures on the Fahrenheit, Celsius, and Kelvin scales. **(Section 0.5)** ***Problems 3, 4, 26, 27, 52–57, 76, 84, 88, 92***
- Know and use the derived SI units for volume **(Section 0.6)** and density. **(Section 0.7)** ***Problems 5, 6, 20, 21, 58–63, 68, 69, 74, 75, 78–80, 83, 85, 89, 90***
- Know and use the derived SI unit for energy. **(Section 0.8)** ***Problems 7, 64, 65, 80, 82***
- Determine the number of significant figures in a measurement **(Section 0.9)**, and round the answer from a calculation to the correct number of significant figures. **(Section 0.10)** ***Problems 8, 9, 11–13, 30, 31, 40, 41, 46–51, 67***
- Use conversion factors and the dimensional-analysis method to convert measurements from one unit to another. **(Section 0.11)** ***Problems 14–16, 38, 52–57, 66–73, 76, 77, 80, 91***

CONCEPTUAL PROBLEMS

Problems at the end of each chapter begin with a section called Conceptual Problems. These problems are visual rather than numerical and are intended to probe your understanding rather than your facility with numbers and formulas. Answers to even-numbered problems can be found at the end of the book following the appendixes.

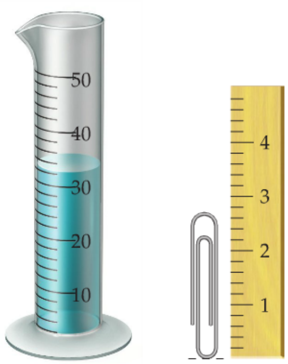
Problems 0.1–0.17 appear within the chapter.

0.18 Which block in each of the following drawings of a balance is more dense, red or green? Explain.

(a) **(b)**

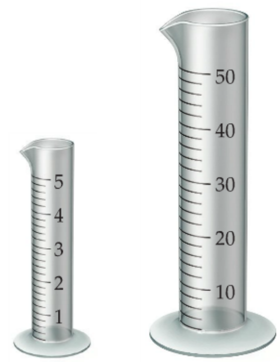
0.19 How many milliliters of water does the graduated cylinder in **(a)** contain, and how tall in centimeters is the paper clip in **(b)**? How many significant figures do you have in each answer?

(a) **(b)**

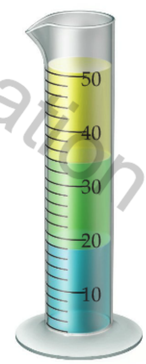


0.20 Assume that you have two graduated cylinders, one with a capacity of 5 mL **(a)** and the other with a capacity of 50 mL **(b)**. Draw a line in each, showing how much liquid you would add if you needed to measure 2.64 mL of water. Which cylinder will give the more accurate measurement? Explain.

(a) **(b)**



0.21 The following cylinder contains three liquids that don't mix with one another: water (density = 1.0 g/mL), vegetable oil (density = 0.93 g/mL), and mercury (density = 13.5 g/mL). Which liquid is which?



SECTION PROBLEMS

The Section Problems at the end of each chapter cover specific topics from the various sections of the chapter. These problems are presented in pairs, with each even-numbered problem followed by an odd-numbered one requiring similar skills.

Units, Significant Figures, and Rounding (Sections 0.2, 0.9, 0.10)

0.22 What is the difference between mass and weight?

0.23 What is the difference between a derived SI unit and a fundamental SI unit? Give an example of each.

0.24 What SI units are used for measuring the following quantities? For derived units, express your answers in terms of the six fundamental units.

(a) Mass **(b)** Length
(c) Temperature **(d)** Volume
(e) Energy **(f)** Density

0.25 What SI prefix corresponds to each of the following multipliers?

(a) 10^3 **(b)** 10^{-6}
(c) 10^9 **(d)** 10^{-12}
(e) 10^{-2} **(f)** 10^{-9}

0.26 Which is larger, a Fahrenheit degree or a Celsius degree? By how much?

0.27 What is the difference between a kelvin and a Celsius degree?

0.28 What is the difference between a cubic decimeter (SI) and a liter (metric)?

0.29 What is the difference between a cubic centimeter (SI) and a milliliter (metric)?

0.30 Which of the following statements use exact numbers?

(a) 1 ft = 12 in. **(b)** 1 cal = 4.184 J
(c) The height of Mt. Everest is 29,035 ft.
(d) The world record for the 1 mile run, set by Morocco's Hicham El Guerrouj in July, 1999, is 3 minutes, 43.13 seconds.

0.31 What is the difference in mass between a nickel that weighs 4.8 g and a nickel that weighs 4.8673 g?

0.32 Bottles of wine sometimes carry the notation "Volume = 75 cL." What does the unit cL mean?

0.33 What do the following abbreviations stand for?

(a) dL (b) dm
(c) µm (d) nL
(e) MJ (f) pg

0.34 Which quantity in each of the following pairs is larger?

(a) 5.63×10^6 cm or 6.02×10^1 km
(b) 46 µs or 3.2×10^{-2} ms
(c) 200,098 g or 17×10^1 kg

0.35 Which quantity in each of the following pairs is smaller?

(a) 154 pm or 7.7×10^{-9} cm
(b) 1.86×10^{11} µm or 2.02×10^2 km
(c) 2.9 GA or 3.1×10^{15} µA

0.36 How many picograms are in 1 mg? In 35 ng?

0.37 How many microliters are in 1 L? In 20 mL?

0.38 Carry out the following conversions:

(a) 5 pm = ____ cm = ____ nm
(b) 8.5 cm^3 = ____ m^3 ____ mm^3
(c) 65.2 mg = ____ g = ____ pg

0.39 Which is larger, and by approximately how much?

(a) A liter or a quart (b) A mile or a kilometer
(c) A gram or an ounce (d) A centimeter or an inch

0.40 How many significant figures are in each of the following measurements?

(a) 35.0445 g (b) 59.0001 cm
(c) 0.030 03 kg (d) 0.004 50 m
(e) 67,000 m^2 (f) 3.8200×10^3 L

0.41 How many significant figures are in each of the following measurements?

(a) $130.95 (b) 2000.003 g
(c) 5 ft 3 in. (d) 510 J
(e) 5.10×10^2 J (f) 10 students

0.42 The Vehicle Assembly Building at the John F. Kennedy Space Center in Cape Canaveral, Florida, is the largest building in the world, with a volume of 3,666,500 m^3. Express this volume in scientific notation.

0.43 The diameter of the Earth at the equator is 7926.381 mi. Round this quantity to four significant figures; to two significant figures. Express the answers in scientific notation.

0.44 Express the following measurements in scientific notation:

(a) 453.32 mg (b) 0.000 042 1 mL (c) 667,000 g

0.45 Convert the following measurements from scientific notation to standard notation:

(a) 3.221×10^{-3} mm (b) 8.940×10^5 m
(c) $1.350\,82 \times 10^{-12}$ m^3 (d) 6.4100×10^2 km

0.46 Round the following quantities to the number of significant figures indicated in parentheses:

(a) 35,670.06 m (4, 6) (b) 68.507 g (2, 3)
(c) 4.995×10^3 cm (3) (d) $2.309\,85 \times 10^{-4}$ kg (5)

0.47 Round the following quantities to the number of significant figures indicated in parentheses:

(a) 7.0001 kg (4) (b) 1.605 km (3)
(c) 13.2151 g/cm^3 (3) (d) 2,300,000.1 (7)

0.48 Express the results of the following calculations with the correct number of significant figures:

(a) 4.884×2.05 (b) $94.61 \div 3.7$
(c) $3.7 \div 94.61$ (d) $5502.3 + 24 + 0.01$
(e) $86.3 + 1.42 - 0.09$ (f) 5.7×2.31

0.49 Express the results of the following calculations with the correct number of significant figures:

(a) $\dfrac{3.41 - 0.23}{5.233} \times 0.205$ (b) $\dfrac{5.556 \times 2.3}{4.223 - 0.08}$

0.50 The world record for the women's outdoor 20,000 meter run, set in 2000 by Tegla Loroupe, is 1:05:26.6 (seconds are given to the nearest tenth). What was her average speed, expressed in miles per hour with the correct number of significant figures? Assume that the race distance is accurate to 5 significant figures.

0.51 In the U.S., the emissions limit for carbon monoxide in motorcycle engine exhaust is 12.0 g of carbon monoxide per kilometer driven. What is this limit expressed in mg per mile with the correct number of significant figures?

Temperature (Section 0.5)

0.52 The normal body temperature of a goat is 39.9 °C, and that of an Australian spiny anteater is 22.2 °C. Express these temperatures in degrees Fahrenheit.

0.53 Of the 90 or so naturally occurring elements, only four are liquid near room temperature: mercury (melting point = −38.87 °C), bromine (melting point = −7.2 °C), cesium (melting point = 28.40 °C), and gallium (melting point = 29.78 °C). Convert these melting points to degrees Fahrenheit.

0.54 Tungsten, the element used to make filaments in lightbulbs, has a melting point of 6192 °F. Convert this temperature to degrees Celsius and to kelvin.

0.55 Suppose that your oven is calibrated in degrees Fahrenheit but a recipe calls for you to bake at 175 °C. What oven setting should you use?

0.56 Suppose that you were dissatisfied with both Celsius and Fahrenheit units and wanted to design your own temperature scale based on ethyl alcohol (ethanol). On the Celsius scale, ethanol has a melting point of −117.3 °C and a boiling point of 78.5 °C, but on your new scale calibrated in units of degrees ethanol, °E, you define ethanol to melt at 0 °E and boil at 200 °E.

(a) How does your ethanol degree compare in size with a Celsius degree?

(b) How does an ethanol degree compare in size with a Fahrenheit degree?

(c) What are the melting and boiling points of water on the ethanol scale?

(d) What is normal human body temperature (98.6 °F) on the ethanol scale?

(e) If the outside thermometer reads 130 °E, how would you dress to go out?

0.57 Answer parts **(a)–(d)** of Problem 0.56 assuming that your new temperature scale is based on ammonia, NH_3. On the Celsius scale, ammonia has a melting point of −77.7 °C and a boiling point of −33.4 °C, but on your new scale calibrated in units of degrees ammonia, °A, you define ammonia to melt at 0 °A and boil at 100 °A.

Density (Section 0.7)

0.58 The density of silver is 10.5 g/cm^3. What is the mass (in kilograms) of a cube of silver that measures 0.62 m on each side?

0.59 A vessel contains 4.67 L of bromine, whose density is 3.10 g/cm^3. What is the mass of the bromine in the vessel (in kilograms)?

0.60 Aspirin has a density of 1.40 g/cm^3. What is the volume in cubic centimeters of an aspirin tablet weighing 250 mg? Of a tablet weighing 500 lb?

0.61 Gaseous hydrogen has a density of 0.0899 g/L at 0 °C, and gaseous chlorine has a density of 3.214 g/L at the same temperature. How many liters of each would you need if you wanted 1.0078 g of hydrogen and 35.45 g of chlorine?

0.62 What is the density of lead in g/cm^3 if a rectangular bar measuring 0.50 cm in height, 1.55 cm in width, and 25.00 cm in length has a mass of 220.9 g?

0.63 What is the density of lithium metal in g/cm^3 if a cylindrical wire with a diameter of 2.40 mm and a length of 15.0 cm has a mass of 0.3624 g?

Energy (Section 0.8)

0.64 The combustion of 45.0 g of methane (natural gas) releases 2498 kJ of heat energy. How much energy in kilocalories (kcal) would combustion of 0.450 ounces of methane release?

0.65 Sodium (Na) metal undergoes a chemical reaction with chlorine (Cl) gas to yield sodium chloride, or common table salt. If 1.00 g of sodium reacts with 1.54 g of chlorine, 2.54 g of sodium chloride is formed and 17.9 kJ of heat is released. How much sodium and how much chlorine in grams would have to react to release 171 kcal of heat?

Unit Conversions (Section 0.11)

0.66 Carry out the following conversions:

(a) How many grams of meat are in a quarter-pound hamburger (0.25 lb)?

(b) How tall in meters is the Willis Tower (formerly called the Sears Tower) in Chicago (1454 ft)?

(c) How large in square meters is the land area of Australia (2,941,526 mi^2)?

0.67 Convert the following quantities into SI units with the correct number of significant figures:

(a) 5.4 in. **(b)** 66.31 lb **(c)** 0.5521 gal
(d) 65 mi/h **(e)** 978.3 yd^3 **(f)** 2.380 mi^2

0.68 The volume of water used for crop irrigation is measured in acre-feet, where 1 acre-foot is the amount of water needed to cover 1 acre of land to a depth of 1 ft.

(a) If there are 640 acres per square mile, how many cubic feet of water are in 1 acre-foot?

(b) How many acre-feet are in Lake Erie (total volume = 116 mi^3)?

0.69 The height of a horse is usually measured in *hands* instead of in feet, where 1 hand equals 1/3 ft (exactly).

(a) How tall in centimeters is a horse of 18.6 hands?

(b) What is the volume in cubic meters of a box measuring 6 × 2.5 × 15 hands?

0.70 Concentrations of substances dissolved in solution are often expressed as mass per unit volume. For example, normal human blood has a cholesterol concentration of about 200 mg/100 mL. Express this concentration in the following units:

(a) mg/L **(b)** µg/mL **(c)** g/L **(d)** ng/µL

(e) How much total blood cholesterol in grams does a person have if the normal blood volume in the body is 5 L?

0.71 Weights in England are commonly measured in *stones*, where 1 stone = 14 lb. What is the weight in pounds of a person who weighs 8.65 stones?

0.72 Among the many alternative units that might be considered as a measure of time is the *shake* rather than the second. Based on the expression "faster than a shake of a lamb's tail," we'll define 1 shake as equal to 2.5×10^{-4} s. If a car is traveling at 55 mi/h, what is its speed in cm/shake?

0.73 Administration of digitalis, a drug used to control atrial fibrillation in heart patients, must be carefully controlled because even a modest overdosage can be fatal. To take differences between patients into account, drug dosages are prescribed in terms of mg/kg body weight. Thus, a child and an adult differ greatly in weight, but both receive the same dosage per kilogram of body weight. At a dosage of 20 µg/kg body weight, how many milligrams of digitalis should a 160 lb patient receive?

CHAPTER PROBLEMS

Chapter Problems are unpaired, are not identified by subject, and can cover topics from any of the various sections of the chapter. In addition, they are generally a bit more challenging than earlier Section Problems.

0.74 When an irregularly shaped chunk of silicon weighing 8.763 g was placed in a graduated cylinder containing 25.00 mL of water, the water level in the cylinder rose to 28.76 mL. What is the density of silicon in g/cm^3?

0.75 Lignum vitae is a hard, durable, and extremely dense wood used to make ship bearings. A sphere of this wood with a diameter of 7.60 cm has a mass of 313 g.

(a) What is the density of the lignum vitae sphere?

(b) Will the sphere float or sink in water?

(c) Will the sphere float or sink in chloroform? (The density of chloroform is 1.48 g/mL.)

0.76 Sodium chloride has a melting point of 1074 K and a boiling point of 1686 K. Convert these temperatures to degrees Celsius and to degrees Fahrenheit.

0.77 A large tanker truck for carrying gasoline has a capacity of 3.4×10^4 L.

(a) What is the tanker's capacity in gallons?

(b) If the retail price of gasoline is $3.00 per gallon, what is the value of the truck's full load of gasoline?

0.78 The density of chloroform, a widely used solvent, is 1.4832 g/mL at 20 °C. How many milliliters would you use if you wanted 112.5 g of chloroform?

0.79 Sulfuric acid (density = 1.8302 g/cm^3) is produced in a larger amount than any other chemical, approximately 3.6×10^{11} lb/yr worldwide. What is the volume of this amount in liters?

0.80 Answer the following questions:

(a) An old rule of thumb in cooking says: "A pint's a pound the world around." What is the density in g/mL of a substance for which 1 pt = 1 lb exactly?

(b) There are exactly 640 acres in 1 square mile. How many square meters are in 1 acre?

(c) A certain type of wood has a density of 0.40 g/cm^3. What is the mass of 1.0 cord of this wood in kg, where 1 cord is 128 cubic feet of wood?

(d) A particular sample of crude oil has a density of 0.85 g/mL. What is the mass of 1.00 barrel of this crude oil in kg, where a barrel of oil is exactly 42 gallons?

(e) A gallon of ice cream contains exactly 32 servings, and each serving has 165 Calories, of which 30.0% are derived from fat. How many Calories derived from fat would you consume if you ate one half-gallon of ice cream?

0.81 A 1.0 ounce piece of chocolate contains 15 mg of caffeine, and a 6.0 ounce cup of regular coffee contains 105 mg of caffeine. How much chocolate would you have to consume to get as much caffeine as you would from 2.0 cups of coffee?

0.82 A bag of Hershey's Kisses contains the following information:

Serving size: 9 pieces = 41 grams

Calories per serving: 230

Total fat per serving: 13 g

(a) The bag contains 2.0 lbs of Hershey's Kisses. How many Kisses are in the bag?

(b) The density of a Hershey's Kiss is 1.4 g/mL. What is the volume of a single Hershey's Kiss?

(c) How many Calories are in one Hershey's Kiss?

(d) Each gram of fat yields 9 Calories when metabolized. What percent of the Calories in Hershey's Kisses are derived from fat?

0.83 Vinaigrette salad dressing consists mainly of olive oil and vinegar. The density of olive oil is 0.918 g/cm^3, the density of vinegar is 1.006 g/cm^3, and the two do not mix. If a certain mixture of olive oil and vinegar has a total mass of 397.8 g and a total volume of 422.8 cm^3, what is the volume of oil and what is the volume of vinegar in the mixture?

0.84 At a certain point, the Celsius and Fahrenheit scales "cross," giving the same numerical value on both. At what temperature does this crossover occur?

0.85 Imagine that you place a cork measuring 1.30 cm × 5.50 cm × 3.00 cm in a pan of water and that on top of the cork you place a small cube of lead measuring 1.15 cm on each edge. The density of cork is 0.235 g/cm^3, and the density of lead is 11.35 g/cm^3. Will the combination of cork plus lead float or sink?

0.86 The LD_{50} of aspirin in rats is given in the *FYI* at the end of this chapter. If a baby aspirin tablet contains 81 mg of aspirin, how many whole tablets would a 0.75 lb rat have to consume to have at least a 50% chance of dying from the dose?

0.87 An Eastern diamondback rattlesnake was milked until 0.134 g of venom was obtained. The venom was then administered subcutaneously in equal portions to 550 mice with an average weight of 0.70 oz, and exactly half the mice died. What is the LD_{50} (in g/kg) for the snake venom in mice? See the *FYI* at the end of this chapter.

0.88 A 125 mL sample of water at 293.2 K was heated for 8 min, 25 s so as to give a constant temperature increase of 3.0 °F/min. What is the final temperature of the water in degrees Celsius?

0.89 A calibrated flask was filled to the 25.00 mL mark with ethyl alcohol. By weighing the flask before and after adding the alcohol, it was determined that the flask contained 19.7325 g of alcohol. In a second experiment, 25.0920 g of metal beads were added to the flask, and the flask was again filled to the 25.00 mL mark with ethyl alcohol. The total mass of the metal plus alcohol in the flask was determined to be 38.4704 g. What is the density of the metal in g/mL?

0.90 Brass is an alloy, or solid mixture, of copper and zinc. What is the mass in grams of a brass cylinder having a length of 1.62 in. and a diameter of 0.514 in. if the composition of the brass is 67.0% copper and 33.0% zinc by mass? The density of copper is 8.92 g/cm^3, and the density of zinc is 7.14 g/cm^3. Assume that the density of the brass varies linearly with composition.

0.91 Ocean currents are measured in *sverdrups* (Sv) where 1 Sv = 10^9 m^3/s. The Gulf Stream off the tip of Florida, for instance, has a flow of 35 Sv.

(a) What is the flow of the Gulf Stream in milliliters per minute?

(b) What mass of water in the Gulf Stream flows past a given point in 24 hours? The density of seawater is 1.025 g/mL.

(c) How much time is required for 1 petaliter (PL; 1 PL = 10^{15} L) of seawater to flow past a given point?

0.92 The element gallium (Ga) has the second largest liquid range of any element, melting at 29.78 °C and boiling at 2204 °C at atmospheric pressure.

(a) What is the density of gallium in g/cm^3 at 25 °C if a 1 in. cube has a mass of 0.2133 lb?

(b) Assume that you construct a thermometer using gallium as the fluid instead of mercury, and that you define the melting point of gallium as 0 °G and the boiling point of gallium as 1000 °G. What is the melting point of sodium chloride (801 °C) on the gallium scale?

0.93 Distances over land are measured in *statute miles* (5280 ft), but distances over water are measured in *nautical miles*, where 1 nautical mile was originally defined as 1 minute of arc along an Earth meridian, or 1/21,600 of the Earth's circumference through the poles. A ship's speed through the water is measured in *knots*, where 1 knot = 1 nautical mile per hour. Historically, the unit *knot* derived from the practice of measuring a ship's speed by throwing a log tied to a knotted line over the side. The line had a knot tied in it at intervals of 47 ft 3 in., and the number of knots run out in 28 seconds was counted to determine speed.

(a) How many feet are in a nautical mile? How many meters?

(b) The northern bluefin tuna can weigh up to 1500 pounds and can swim at speeds up to 48 miles per hour. How fast is this in knots?

(c) A *league* is defined as 3 nautical miles. The Mariana Trench, with a depth of 35,798 feet, is the deepest point in the ocean. How many leagues deep is this?

(d) By international agreement, the nautical mile is now defined as exactly 1852 meters. By what percentage does this current definition differ from the original definition, and by what percentage does it differ from a statute mile?