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Chemistry

An Atoms First Approach

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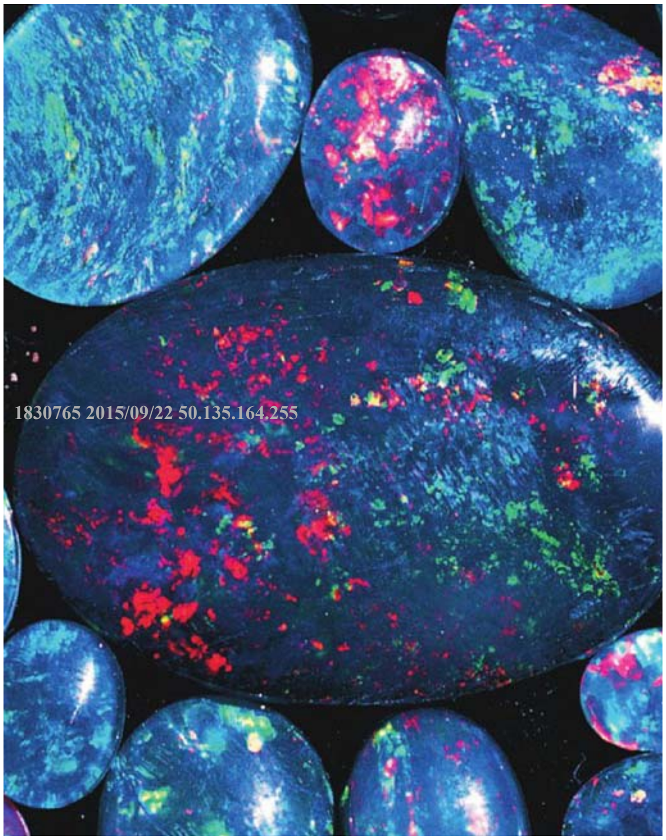

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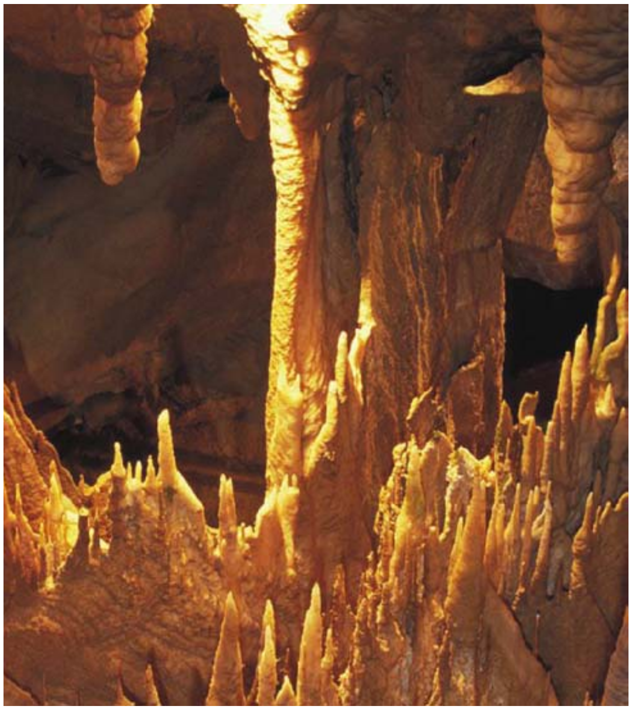

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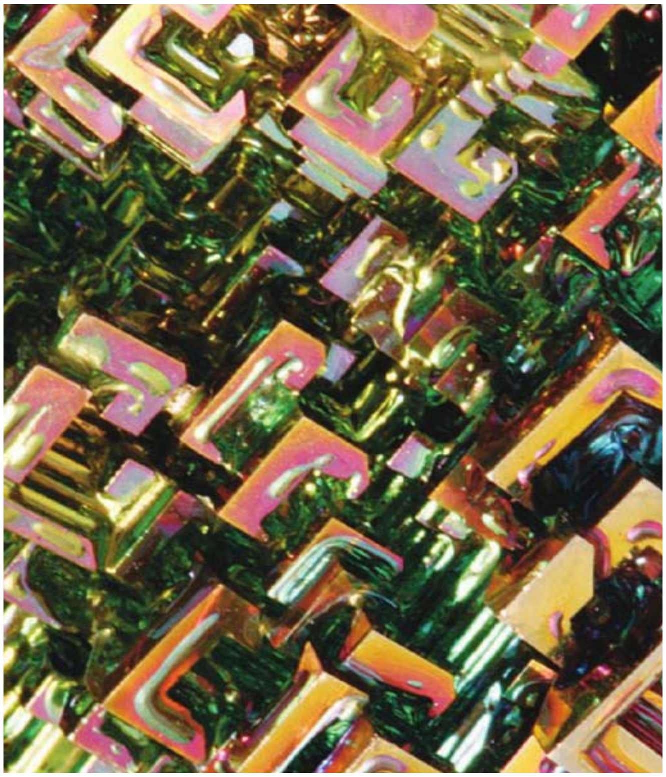
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to the student

As you jump into the study of chemistry we hope that you will find our text helpful and interesting. Our job is to present the concepts and ideas of chemistry in a way you can understand. We hope to encourage you in your studies and to help you learn to solve problems in ways you can apply in all areas of your professional and personal lives.

Our main goal is to help you learn to become a truly creative problem solver. Our world badly needs people who can "think outside the box." Our focus is to help you learn to think like a chemist. Why would you want to do that? Chemists are great problem solvers. They use logic, trial and error, and intuition—along with lots of patience—to work through complex problems. Chemists make mistakes, as we all do in our lives. The important thing that a chemist does is to learn from the mistakes and to try again. This "can do" attitude is useful in all careers.

In this book we develop the concepts in a natural way: The observations come first and then we develop models to explain the observed behavior. Models help us to understand and explain our world. They are central to scientific thinking. Models are very useful, but they also have limitations, which we will point out. By understanding the basic concepts in chemistry we lay the foundation for solving problems.

Our main goal is to help you learn a thoughtful method of problem solving. True learning is more than memorizing facts. Truly educated people use their factual knowledge as a starting point—a basis for creative problem solving. Our strategy for solving problems is explained in Section 6.2. To solve a problem we ask ourselves questions, which help us think through the problem. We let the problem guide us to the solution. This process can be applied to all types of problems in all areas of life.

As you study the text, use the *Examples* and the problem-solving strategies to help you. The strategies are boxed to highlight them for you and the *Examples* show how these strategies are applied.

After you have read and studied each chapter of the text you'll need to practice your problem-solving skills. To do this we have provided plenty of review questions and end-of-chapter exercises. Your instructor may assign these on paper or online; in either case, you'll want to work with your fellow students. One of the most effective ways to learn chemistry is through the exchange of ideas that comes from helping one another. The online homework assignments will give you instant feedback and, in print, we have provided answers to some of the exercises in the back of the text. In all cases, your main goal is not just to get the correct answer, but to understand the process for getting the answer. Memorizing solutions for specific problems is not a very good way to prepare for an exam (or to solve problems in the real world!).

To become a great problem solver you'll need these skills:

1. Look within the problem for the solution. (Let the problem guide you.)
2. Use the concepts you have learned along with a systematic, logical approach to find the solution.
3. Solve the problem by asking questions and learn to trust yourself to think it out.

You will make mistakes, but the important thing is to learn from these errors. The only way to gain confidence is to practice, practice, practice and to use your mistakes to find your weaknesses. Be patient with yourself and work hard to understand rather than simply memorize.

We hope you'll have an interesting and successful year learning to think like a chemist!

Steve and Susan Zumdahl

review

Measurement and Calculations in Chemistry

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Measuring quantities of liquids precisely is important in chemistry.

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Making observations is fundamental to all science. These observations can be qualitative or quantitative. A quantitative observation is called a **measurement**, which always has two parts: a **number** and a scale (called a **unit**). Both parts must be present for a measurement to be meaningful. A qualitative observation does not involve a number. Examples of qualitative observations are "the substance is blue" and "the sun is very hot."

In this chapter we will discuss measurements in detail and explain the various properties of the numbers and units associated with measurements. This material should be familiar to you from previous science courses, but we include it here to provide a review of these topics that are critical to the operations of chemistry.

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Soda is commonly sold in 2-liter bottles—an example of the use of SI units in everyday life.

R.1 › Units of Measurement

In our study of chemistry we will use measurements of mass, length, time, temperature, electric current, and the amount of a substance, among others. Scientists recognized long ago that standard systems of units had to be adopted if measurements were to be useful. If every scientist had a different set of units, complete chaos would result. Unfortunately, different standards were adopted in different parts of the world. The two major systems are the *English system* used in the United States and the *metric system* used by most of the rest of the industrialized world. This duality causes a good deal of trouble; for example, parts as simple as bolts are not interchangeable between machines built according to the two systems. As a result, the United States has begun to adopt the metric system.

Most scientists in all countries have for many years used the metric system. In 1960, an international agreement set up a system of units called the *International System* (*le Système International* in French), or the **SI system**. This system is based on the metric system and units derived from the metric system. The fundamental SI units are listed in Table R.1. We will discuss how to manipulate these units later in this chapter.

Because the fundamental units are not always convenient (expressing the mass of a pin in kilograms is awkward), prefixes are used to change the size of the unit. These are listed in Table R.2. Some common objects and their measurements in SI units are listed in Table R.3.

One physical quantity that is very important in chemistry is *volume*, which is not a fundamental SI unit but is derived from length. A cube that measures 1 meter (m) on each edge is represented in Fig. R.1. This cube has a volume of $(1\ \text{m})^3 = 1\ \text{m}^3$. Because there are 10 decimeters (dm) in a meter, the volume of this cube is $(1\ \text{m})^3 = (10\ \text{dm})^3 = 1000\ \text{dm}^3$. A cubic decimeter, that is $(1\ \text{dm})^3$, is commonly called a *liter (L)*, which is a

TABLE R.1 › The Fundamental SI Units

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

TABLE R.2 › **The Prefixes Used in the SI System (Those most commonly encountered are shown in blue.)**

Prefix	Symbol	Meaning	Exponential Notation*
exa	E	1,000,000,000,000,000,000	10^{18}
peta	P	1,000,000,000,000,000	10^{15}
tera	T	1,000,000,000,000	10^{12}
giga	G	1,000,000,000	10^{9}
mega	M	1,000,000	10^{6}
kilo	k	1,000	10^{3}
hecto	h	100	10^{2}
deka	da	10	10^{1}
—	—	1	10^{0}
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}
pico	p	0.000000000001	10^{-12}
femto	f	0.000000000000001	10^{-15}
atto	a	0.000000000000000001	10^{-18}

*See Appendix 1.1 if you need a review of exponential notation.

TABLE R.3 › **Some Examples of Commonly Used Units**

Length	A dime is 1 mm thick. A quarter is 2.5 cm in diameter. The average height of an adult man is 1.8 m.
Mass	A nickel has a mass of about 5 g. A 120-lb person has a mass of about 55 kg.
Volume	A 12-oz can of soda has a volume of about 360 mL.

unit of volume slightly larger than a quart. As shown in Fig. R.1, 1000 liters are contained in a cube with a volume of 1 cubic meter. Similarly, since 1 decimeter equals 10 centimeters (cm), the liter can be divided into 1000 cubes each with a volume of 1 cubic centimeter:

$$1 \text{ liter} = (1 \text{ dm})^3 = (10 \text{ cm})^3 = 1000 \text{ cm}^3$$

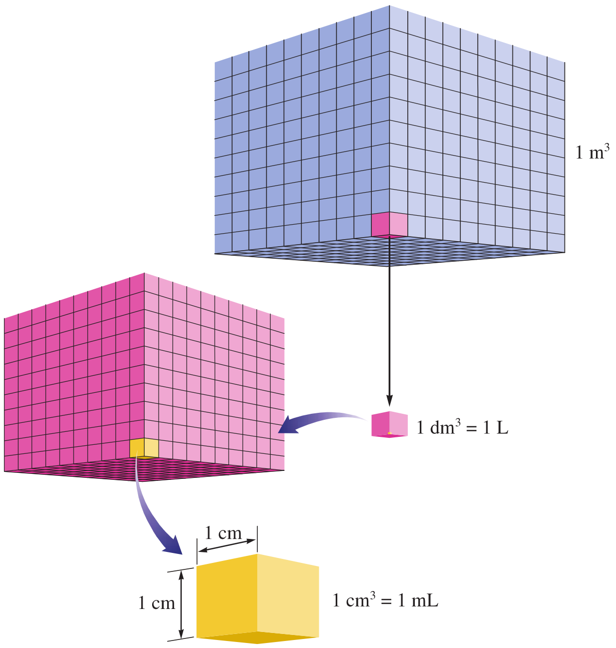

FIGURE R.1
The largest cube has sides 1 m in length and a volume of 1 m^3. The middle-sized cube has sides 1 dm in length and a volume of 1 dm^3, or 1 L. The smallest cube has sides 1 cm in length and a volume of 1 cm^3, or 1 mL.

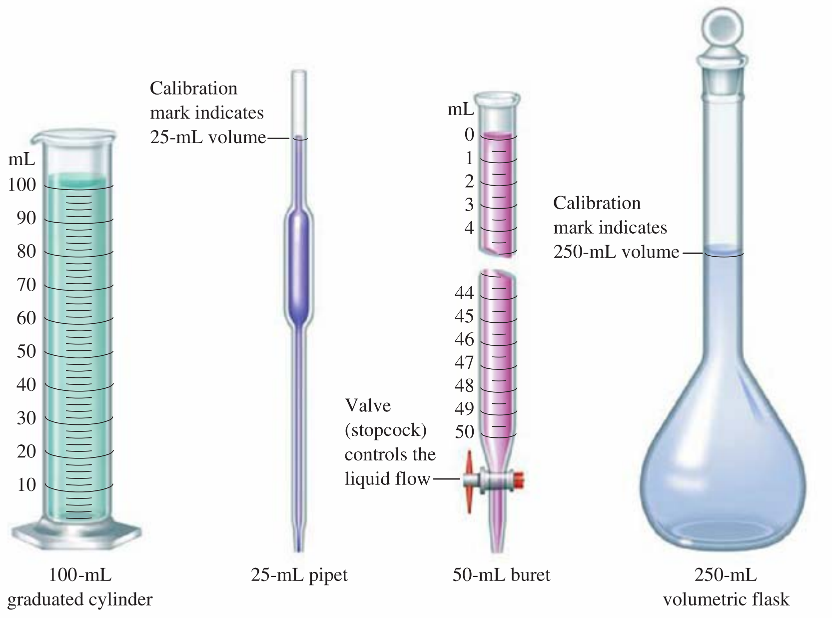

FIGURE R.2
Common types of laboratory equipment used to measure liquid volume.

Also, since 1 cm^3 = 1 milliliter (mL),

$$1 \text{ liter} = 1000 \text{ cm}^3 = 1000 \text{ mL}$$

Thus 1 liter contains 1000 cubic centimeters, or 1000 milliliters.

Chemical laboratory work frequently requires measurement of the volumes of liquids. Several devices for the accurate determination of liquid volume are shown in Fig. R.2.

An important point concerning measurements is the relationship between mass and weight. Although these terms are sometimes used interchangeably, they are *not* the same. ***Mass** is a measure of the resistance of an object to a change in its state of motion*. Mass is measured by the force necessary to give an object a certain acceleration. On earth we use the force that gravity exerts on an object to measure its mass. We call this force the object's **weight**. Since weight is the response of mass to gravity, it varies with the strength of the gravitational field. Therefore, your body mass is the same on the earth or on the moon, but your weight would be much less on the moon than on earth because of the moon's smaller gravitational field.

Because weighing something on a chemical balance involves comparing the mass of that object to a standard mass, the terms *weight* and *mass* are sometimes used interchangeably, although this is incorrect.

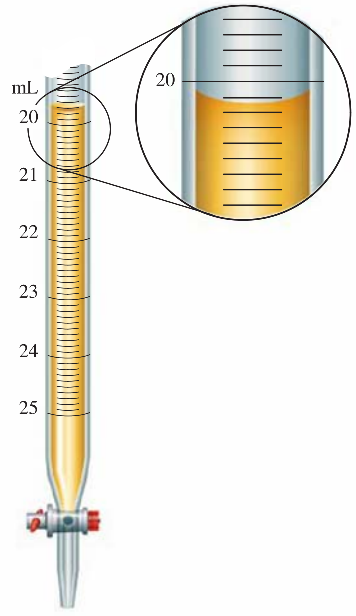

FIGURE R.3
Measurement of volume using a buret. The volume is read at the bottom of the liquid curve (called the meniscus).

R.2 › Uncertainty in Measurement

The number associated with a measurement is obtained using some measuring device. For example, consider the measurement of the volume of a liquid using a buret (shown in Fig. R.3 with the scale greatly magnified). Notice that the meniscus of the liquid occurs at about 20.15 milliliters. This means that about 20.15 mL of liquid has been delivered from the buret (if the initial position of the liquid meniscus was 0.00 mL). Note that we must estimate the last number of the volume reading by interpolating between the 0.1-mL marks. Since the last number is estimated, its value may be different if another person makes the same measurement. If five different people read the same volume, the results might be as follows:

Person	Results of Measurement
1	20.15 mL
2	20.14 mL
3	20.16 mL
4	20.17 mL
5	20.16 mL

These results show that the first three numbers (20.1) remain the same regardless of who makes the measurement; these are called *certain* digits. However, the digit to the right of the 1 must be estimated and therefore varies; it is called an *uncertain* digit. We customarily report a measurement by recording all the certain digits plus the *first* uncertain digit. In our example it would not make any sense to try to record the volume of thousandths of a milliliter, because the value for hundredths of a milliliter must be estimated when using the buret.

It is very important to realize that a *measurement always has some degree of **uncertainty***. The uncertainty of a measurement depends on the precision of the measuring device. For example, using a bathroom scale, you might estimate the mass of a grapefruit to be approximately 1.5 pounds. Weighing the same grapefruit on a highly precise balance might produce a result of 1.476 pounds. In the first case, the uncertainty occurs in the tenths of a pound place; in the second case, the uncertainty occurs in the thousandths of a pound place. Suppose we weigh two similar grapefruits on the two devices and obtain the following results:

A measurement always has some degree of uncertainty.

	Bathroom Scale	Balance
Grapefruit 1	1.5 lb	1.476 lb
Grapefruit 2	1.5 lb	1.518 lb

Do the two grapefruits have the same mass? The answer depends on which set of results you consider. Thus a conclusion based on a series of measurements depends on the certainty of those measurements. For this reason, it is important to indicate the uncertainty in any measurement. This is done by always recording the certain digits and the first uncertain digit (the estimated number). These numbers are called the **significant figures** of a measurement.

Uncertainty in measurement is discussed in more detail in Appendix 1.5.

The convention of significant figures automatically indicates something about the uncertainty in a measurement. The uncertainty in the last number (the estimated number) is usually assumed to be ±1 unless otherwise indicated. For example, the measurement 1.86 kilograms can be taken to mean 1.86 ± 0.01 kilograms.

EXAMPLE R.1 Uncertainty in Measurement

In analyzing a sample of polluted water, a chemist measured out a 25.00-mL water sample with a pipet (see Fig. R.2). At another point in the analysis, the chemist used a graduated cylinder (see Fig. R.2) to measure 25 mL of a solution. What is the difference between the measurements 25.00 mL and 25 mL?

Solution

Even though the two volume measurements appear to be equal, they really convey different information. The quantity 25 mL means that the volume is between 24 mL and 26 mL, whereas the quantity 25.00 mL means that the volume is between 24.99 mL and 25.01 mL. The pipet measures volume with much greater precision than does the graduated cylinder.

See Question R.19

When making a measurement, it is important to record the results to the appropriate number of significant figures. For example, if a certain buret can be read to ±0.01 mL, you should record a reading of twenty-five milliliters as 25.00 mL, not 25 mL. This way, at some later time when you are using your results to do calculations, the uncertainty in the measurement will be known to you.

Precision and Accuracy

Two terms often used to describe the reliability of measurements are *precision* and *accuracy*. Although these words are frequently used interchangeably in everyday life, they have different meanings in the scientific context. **Accuracy** refers to the agreement of a particular value with the true value. **Precision** refers to the degree of agreement among several measurements of the same quantity. Precision reflects the *reproducibility* of a given type of measurement. The difference between these terms is illustrated by the results of three different dart throws shown in Fig. R.4.

Two different types of errors are illustrated in Fig. R.4. A **random error** (also called an *indeterminate error*) means that a measurement has an equal probability of being high or low. This type of error occurs in estimating the value of the last digit of a measurement. The second type of error is called **systematic error** (or *determinate error*). This type of error occurs in the same direction each time; it is either always high or always low. Fig. R.4(a) indicates large random errors (poor technique). Fig. R.4(b) indicates small random errors but a large systematic error, and Fig. R.4(c) indicates small random errors and no systematic error.

In quantitative work, precision is often used as an indication of accuracy; we assume that the *average* of a series of precise measurements (which should "average out" the random errors because of their equal probability of being high or low) is accurate, or close to the "true" value. However, this assumption is valid only if systematic errors are absent. Suppose we weigh a piece of brass five times on a very precise balance and obtain the following results:

Weighing	Result
1	2.486 g
2	2.487 g
3	2.485 g
4	2.484 g
5	2.488 g

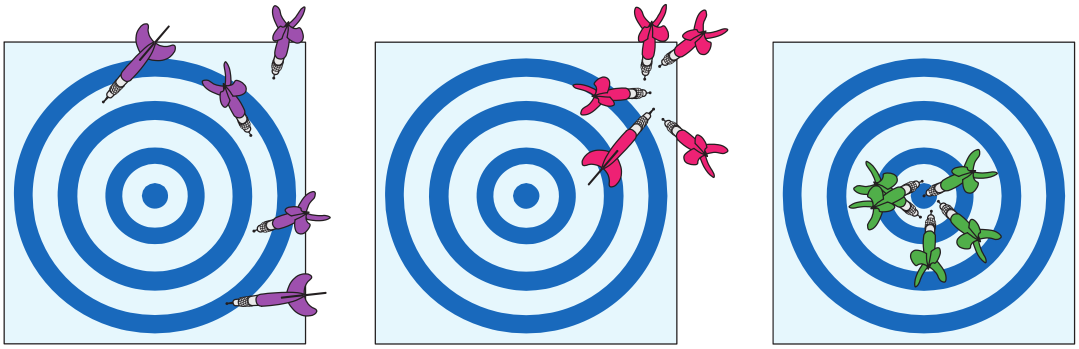
(a) (b) (c)

FIGURE R.4
The results of several dart throws show the difference between precise and accurate. (a) Neither accurate nor precise (large random errors). (b) Precise but not accurate (small random errors, large systematic error). (c) Bull's-eye! Both precise and accurate (small random errors, no systematic error).

Normally, we would assume that the true mass of the piece of brass is very close to 2.486 grams, which is the average of the five results:

$$\frac{2.486\text{ g} + 2.487\text{ g} + 2.485\text{ g} + 2.484\text{ g} + 2.488\text{ g}}{5} = 2.486\text{ g}$$

However, if the balance has a defect causing it to give a result that is consistently 1.000 gram too high (a systematic error of +1.000 gram), then the measured value of 2.486 grams would be seriously in error. The point here is that high precision among several measurements is an indication of accuracy *only* if systematic errors are absent.

EXAMPLE R.2

Precision and Accuracy

To check the accuracy of a graduated cylinder, a student filled the cylinder to the 25-mL mark using water delivered from a buret (see Fig. R.2) and then read the volume delivered. Following are the results of five trials:

Trial	Volume Shown by Graduated Cylinder	Volume Shown by the Buret
1	25 mL	26.54 mL
2	25 mL	26.51 mL
3	25 mL	26.60 mL
4	25 mL	26.49 mL
5	25 mL	26.57 mL
Average	25 mL	26.54 mL

Is the graduated cylinder accurate?

Precision is an indication of accuracy only if there are no systematic errors.

Solution

The results of the trials show very good precision (for a graduated cylinder). The student has good technique. However, note that the average value measured using the buret is significantly different from 25 mL. Thus this graduated cylinder is not very accurate. It produces a systematic error (in this case, the indicated result is low for each measurement)

See Question R.2

R.3 › Significant Figures and Calculations

Calculating the final result for an experiment usually involves adding, subtracting, multiplying, or dividing the results of various types of measurements. Since it is very important that the uncertainty in the final result is known correctly, we have developed rules for counting the significant figures in each number and for determining the correct number of significant figures in the final result.

Rules for Counting Significant Figures

1. *Nonzero integers*. Nonzero integers always count as significant figures.
2. *Zeros*. There are three classes of zeros:
    a. *Leading zeros* are zeros that *precede* all the nonzero digits. These do not count as significant figures. In the number 0.0025, the three zeros simply indicate the position of the decimal point. This number has only two significant figures.

Leading zeros are never significant figures.

(continued)

Captive zeros are always significant figures.

Trailing zeros are sometimes significant figures.

Exact numbers never limit the number of significant figures in a calculation.

Rules for Counting Significant Figures (continued)

- **b.** *Captive zeros* are zeros *between* nonzero digits. These always count as significant figures. The number 1.008 has four significant figures.
- **c.** *Trailing zeros* are zeros at the *right end* of the number. They are significant only if the number contains a decimal point. The number 100 has only one significant figure, whereas the number 1.00×10^2 has three significant figures. The number one hundred written as 100. also has three significant figures.

3. *Exact numbers*. Many times calculations involve numbers that were not obtained using measuring devices but were determined by counting: 10 experiments, 3 apples, 8 molecules. Such numbers are called *exact numbers*. They can be assumed to have an infinite number of significant figures. Other examples of exact numbers are the 2 in $2\pi r$ (the circumference of a circle) and the 4 and the 3 in $\frac{4}{3}\pi r^3$ (the volume of a sphere). Exact numbers also can arise from definitions. For example, one inch is defined as *exactly* 2.54 centimeters. Thus, in the statement 1 in = 2.54 cm, neither the 2.54 nor the 1 limits the number of significant figures when used in a calculation.

Exponential notation is reviewed in Appendix 1.1.

Note that the number 1.00×10^2 above is written in **exponential notation**. This type of notation has at least two advantages: the number of significant figures can be easily indicated, and fewer zeros are needed to write a very large or very small number. For example, the number 0.000060 is much more conveniently represented as 6.0×10^{-5}. (The number has two significant figures.)

INTERACTIVE EXAMPLE R.3

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Significant Figures

Give the number of significant figures for each of the following results.

a. A student's extraction procedure on tea yields 0.0105 g of caffeine.
b. A chemist records a mass of 0.050080 g in an analysis.
c. In an experiment a span of time is determined to be 8.050×10^{-3} s.

Solution

a. The number contains three significant figures. The zeros to the left of the 1 are leading zeros and are not significant, but the remaining zero (a captive zero) is significant.
b. The number contains five significant figures. The leading zeros (to the left of the 5) are not significant. The captive zeros between the 5 and the 8 are significant, and the trailing zero to the right of the 8 is significant because the number contains a decimal point.
c. This number has four significant figures. Both zeros are significant.

See Exercises R.13 through R.16

To this point we have learned to count the significant figures in a given number. Next, we must consider how uncertainty accumulates as calculations are carried out. The detailed analysis of the accumulation of uncertainties depends on the type of calculation involved and can be complex. However, in this textbook we will employ the following simple rules that have been developed for determining the appropriate number of significant figures in the result of a calculation.

Rules for Significant Figures in Mathematical Operations

1. *For multiplication or division*, the number of significant figures in the result is the same as the number in the least precise measurement used in the calculation. For example, consider the calculation

$$4.56 \times \underset{\substack{\uparrow \\ \text{Limiting term has} \\ \text{two significant} \\ \text{figures}}}{1.4} = 6.38 \xrightarrow{\text{Corrected}} \underset{\substack{\uparrow \\ \text{Two significant} \\ \text{figures}}}{6.4}$$

The product should have only two significant figures, since 1.4 has two significant figures.

2. *For addition or subtraction*, the result has the same number of decimal places as the least precise measurement used in the calculation. For example, consider the sum

$$\begin{array}{l} 12.11 \\ 18.0 \quad \leftarrow \text{Limiting term has one decimal place} \\ \underline{1.013} \\ 31.123 \xrightarrow{\text{Corrected}} \underset{\substack{\uparrow \\ \text{One decimal place}}}{31.1} \end{array}$$

The correct result is 31.1, since 18.0 has only one decimal place.

Note that for multiplication and division, significant figures are counted. For addition and subtraction, the decimal places are counted.

For multiplication and division: significant figures are counted. For addition and subtraction: decimal places are counted.

In most calculations you will need to round numbers to obtain the correct number of significant figures. The following rules should be applied when rounding.

Rules for Rounding

1. In a series of calculations, carry the extra digits through to the final result, *then* round.
2. If the digit to be removed
   - **a.** is less than 5, the preceding digit stays the same. For example, 1.33 rounds to 1.3.
   - **b.** is equal to or greater than 5, the preceding digit is increased by 1. For example, 1.36 rounds to 1.4.

Rule 2 is consistent with the operation of electronic calculators.

Although rounding is generally straightforward, one point requires special emphasis. As an illustration, suppose that the number 4.348 needs to be rounded to two significant figures. In doing this, we look *only* at the *first number* to the right of the 3:

$$4.3\underset{\substack{\uparrow \\ \text{Look at this number to} \\ \text{round to two significant figures.}}}{4}8$$

The number is rounded to 4.3 because 4 is less than 5. It is incorrect to round sequentially. For example, do *not* round the 4 to 5 to give 4.35 and then round the 3 to 4 to give 4.4.

Do not round sequentially. The number 6.8347 rounded to three significant figures is 6.83, not 6.84.

When rounding, *use only the first number to the right of the last significant figure*.

It is important to note that Rule 1 above usually will not be followed in the Examples in this text because we want to show the correct number of significant figures in *each step* of a problem. This same practice is followed for the detailed solutions given in the *Solutions Guide*. However, when you are doing problems, you should carry extra digits throughout a series of calculations and round to the correct number of significant figures only at the end. This is the practice you should follow. The fact that your rounding procedures are different from those used in this text must be taken into account when you check your answer with the one given at the end of the book or in the *Solutions Guide*. Your answer (based on rounding only at the end of a calculation) may differ in the last place from that given here as the "correct" answer because we have rounded after each step. To help you understand the difference between these rounding procedures, we will consider them further in Example R.4.

INTERACTIVE EXAMPLE R.4

Significant Figures in Mathematical Operations

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Carry out the following mathematical operations, and give each result with the correct number of significant figures.

a. $1.05 \times 10^{-3} \div 6.135$
b. $21 - 13.8$
c. As part of a lab assignment to determine the value of the gas constant (R), a student measured the pressure (P), volume (V), and temperature (T) for a sample of gas, where

$$R = \frac{PV}{T}$$

The following values were obtained: $P = 2.560$, $T = 275.15$, and $V = 8.8$. (Gases will be discussed in detail in Chapter 7; we will not be concerned at this time about the units for these quantities.) Calculate R to the correct number of significant figures.

Solution

a. The result is 1.71×10^{-4}, which has three significant figures because the term with the least precision (1.05×10^{-3}) has three significant figures.
b. The result is 7 with no decimal point because the number with the least number of decimal places (21) has none.
c.

$$R = \frac{PV}{T} = \frac{(2.560)(8.8)}{275.15}$$

The correct procedure for obtaining the final result can be represented as follows:

$$\frac{(2.560)(8.8)}{275.15} = \frac{22.528}{275.15} = 0.0818753$$
$$= 0.082 = 8.2 \times 10^{-2} = R$$

The final result must be rounded to two significant figures because 8.8 (the least precise measurement) has two significant figures. To show the effects of rounding at intermediate steps, we will carry out the calculation as follows:

$$\frac{(2.560)(8.8)}{275.15} = \frac{22.528}{275.15} = \frac{\overset{\substack{\text{Rounded to two}\\ \text{significant figures}\\ \downarrow}}{23}}{275.15}$$

Now we proceed with the next calculation:

$$\frac{23}{275.15} = 0.0835908$$

Rounded to two significant figures, this result is

$$0.084 = 8.4 \times 10^{-2}$$

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This number must be rounded to two significant figures.

Note that intermediate rounding gives a significantly different result than was obtained by rounding only at the end. Again, we must reemphasize that in *your* calculations you should round *only at the end*. However, because rounding is carried out at intermediate steps in this text (to always show the correct number of significant figures), the final answer given in the text may differ slightly from the one you obtain (rounding only at the end).

See Exercises R.21 through R.24

There is a useful lesson to be learned from part c of Example R.4. The student measured the pressure and temperature to greater precision than the volume. A more precise value of R (one with more significant figures) could have been obtained if a more precise measurement of V had been made. As it is, the efforts expended to measure P and T very precisely were wasted. Remember that a series of measurements to obtain some final result should all be done to about the same precision.

R.4 › Dimensional Analysis

It is often necessary to convert a given result from one system of units to another. The best way to do this is by a method called the **unit factor method** or, more commonly, **dimensional analysis**. To illustrate the use of this method, we will consider several unit conversions. Some equivalents in the English and metric systems are listed in Table R.4. A more complete list of conversion factors given to more significant figures appears in Appendix 6.

TABLE R.4 › English–Metric Equivalents

Length	1 m = 1.094 yd 2.54 cm = 1 in
Mass	1 kg = 2.205 lb 453.6 g = 1 lb
Volume	1 L = 1.06 qt 1 ft^3 = 28.32 L

Consider a pin measuring 2.85 centimeters in length. What is its length in inches? To accomplish this conversion, we must use the equivalence statement

$$2.54\ \text{cm} = 1\ \text{in}$$

If we divide both sides of this equation by 2.54 centimeters, we get

$$1 = \frac{1\ \text{in}}{2.54\ \text{cm}}$$

This expression is called a *unit factor*. Since 1 inch and 2.54 centimeters are exactly equivalent, multiplying any expression by this unit factor will not change its *value*.

The pin has a length of 2.85 centimeters. Multiplying this length by the appropriate unit factor gives

$$2.85\ \cancel{\text{cm}} \times \frac{1\ \text{in}}{2.54\ \cancel{\text{cm}}} = \frac{2.85}{2.54}\ \text{in} = 1.12\ \text{in}$$

Note that the centimeter units cancel to give inches for the result. This is exactly what we wanted to accomplish. Note also that the result has three significant figures, as required by the number 2.85. Recall that the 1 and 2.54 in the conversion factor are exact numbers by definition.

INTERACTIVE EXAMPLE R.5

OWL Sign in to OWL at **www.cengage.com/owl** to view an interactive version of this problem.

Unit Conversions I

A pencil is 7.00 in long. What is its length in centimeters?

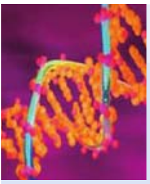
Solution

In this case we want to convert from inches to centimeters. Therefore, we must use the reciprocal of the unit factor used above to do the opposite conversion:

$$7.00\ \cancel{\text{in}} \times \frac{2.54\ \text{cm}}{1\ \cancel{\text{in}}} = (7.00)(2.54)\ \text{cm} = 17.8\ \text{cm}$$

Here the inch units cancel, leaving centimeters, as requested.

See Exercises R.27 and R.28



CHEMICAL CONNECTIONS

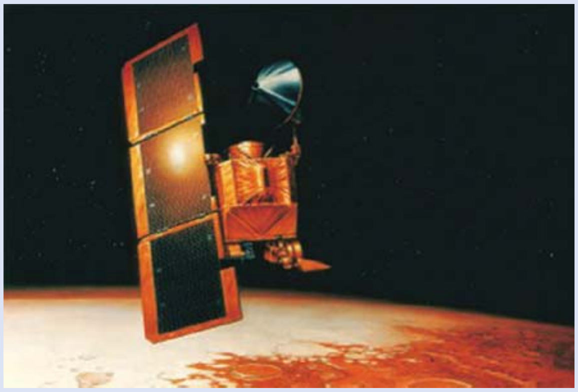
Critical Units!

How important are conversions from one unit to another? If you ask the National Aeronautics and Space Administration (NASA), very important! In 1999, NASA lost a $125 million Mars Climate Orbiter because of a failure to convert from English to metric units.

The problem arose because two teams working on the Mars mission were using different sets of units. NASA's scientists at the Jet Propulsion Laboratory in Pasadena, California, assumed that the thrust data for the rockets on the Orbiter they received from Lockheed Martin Astronautics in Denver, which built the spacecraft, were in metric units. In reality, the units were English. As a result the Orbiter dipped 100 kilometers lower into the Mars atmosphere than planned, and the friction from the atmosphere caused the craft to burn up.

NASA's mistake refueled the controversy over whether Congress should require the United States to switch to the metric system. About 95% of the world now uses the metric system, and the United States is slowly switching from English to metric. For example, the automobile industry has adopted metric fasteners, and we buy our soda in two-liter bottles.

Units can be very important. In fact, they can mean the difference between life and death on some occasions. In 1983, for example, a Canadian jetliner almost ran out of fuel when someone pumped 22,300 pounds of fuel into the aircraft instead of 22,300 kilograms. Remember to watch your units!



NASA

Artist's conception of the lost Mars Climate Orbiter.

Note that two unit factors can be derived from each equivalence statement. For example, from the equivalence statement 2.54 cm = 1 in, the two unit factors are

$$\frac{2.54\text{ cm}}{1\text{ in}} \quad \text{and} \quad \frac{1\text{ in}}{2.54\text{ cm}}$$

Consider the direction of the required change to select the correct unit factor.

How do you choose which one to use in a given situation? Simply look at the *direction* of the required change. To change from inches to centimeters, the inches must cancel. Thus the factor 2.54 cm/1 in is used. To change from centimeters to inches, centimeters must cancel, and the factor 1 in/2.54 cm is appropriate.

PROBLEM-SOLVING STRATEGY

Converting from One Unit to Another

- To convert from one unit to another, use the equivalence statement that relates the two units.
- Derive the appropriate unit factor by looking at the direction of the required change (to cancel the unwanted units).
- Multiply the quantity to be converted by the unit factor to give the quantity with the desired units.

EXAMPLE R.6 Unit Conversions II

You want to order a bicycle with a 25.5-in frame, but the sizes in the catalog are given only in centimeters. What size should you order?

Solution

You need to go from inches to centimeters, so 2.54 cm = 1 in is appropriate:

$$25.5\ \cancel{\text{in}} \times \frac{2.54\ \text{cm}}{1\ \cancel{\text{in}}} = 64.8\ \text{cm}$$

See Exercises R.27 and R.28

To ensure that the conversion procedure is clear, a multistep problem is considered in Example R.7.

INTERACTIVE EXAMPLE R.7

OWL Sign in to OWL at **www.cengage.com/owl** to view an interactive version of this problem.

Unit Conversions III

A student has entered a 10.0-km run. How long is the run in miles?

Solution

This conversion can be accomplished in several different ways. Since we have the equivalence statement 1 m = 1.094 yd, we will proceed by a path that uses this fact. Before we start any calculations, let us consider our strategy. We have kilometers, which we want to change to miles. We can do this by the following route:

$$\text{kilometers} \longrightarrow \text{meters} \longrightarrow \text{yards} \longrightarrow \text{miles}$$

To proceed in this way, we need the following equivalence statements:

$$\begin{aligned} 1\ \text{km} &= 1000\ \text{m} \\ 1\ \text{m} &= 1.094\ \text{yd} \\ 1760\ \text{yd} &= 1\ \text{mi} \end{aligned}$$

To make sure the process is clear, we will proceed step by step:

Kilometers to Meters

$$10.0\ \cancel{\text{km}} \times \frac{1000\ \text{m}}{1\ \cancel{\text{km}}} = 1.00 \times 10^4\ \text{m}$$

Meters to Yards

$$1.00 \times 10^4\ \cancel{\text{m}} \times \frac{1.094\ \text{yd}}{1\ \cancel{\text{m}}} = 1.094 \times 10^4\ \text{yd}$$

Note that we should have only three significant figures in the result. Since this is an intermediate result, however, we will carry the extra digit. Remember, round off only the final result.

Yards to Miles

In the text we round to the correct number of significant figures after each step to show the correct significant figures for each calculation. However, since you use a calculator and combine steps on it, you should round only at the end.

$$1.094 \times 10^4\ \cancel{\text{yd}} \times \frac{1\ \text{mi}}{1760\ \cancel{\text{yd}}} = 6.216\ \text{mi}$$

Note in this case that 1 mi equals exactly 1760 yd *by designation*. Thus 1760 is an exact number.

Since the distance was originally given as 10.0 km, the result can have only three significant figures and should be rounded to 6.22 mi. Thus

$$10.0\ \text{km} = 6.22\ \text{mi}$$

Alternatively, we can combine the steps:

$$10.0\ \cancel{\text{km}} \times \frac{1000\ \cancel{\text{m}}}{1\ \cancel{\text{km}}} \times \frac{1.094\ \cancel{\text{yd}}}{1\ \cancel{\text{m}}} \times \frac{1\ \text{mi}}{1760\ \cancel{\text{yd}}} = 6.22\ \text{mi}$$

See Exercises R.27 and R.28

In using dimensional analysis, your verification that everything has been done correctly is that you end up with the correct units. In doing chemistry problems, you should always include the units for the quantities used. Always check to see that the units cancel to give the correct units for the final result. This provides a very valuable check, especially for complicated problems.

Study the procedures for unit conversions in the following Examples.

INTERACTIVE EXAMPLE R.8

OWL Sign in to OWL at **www.cengage.com/owl** to view an interactive version of this problem.

Unit Conversions IV

The speed limit on many highways in the United States is 55 mi/h. What number would be posted in kilometers per hour?

Solution

$$\frac{55\ \cancel{\text{mi}}}{\text{h}} \times \frac{1760\ \cancel{\text{yd}}}{1\ \cancel{\text{mi}}} \times \frac{1\ \cancel{\text{m}}}{1.094\ \cancel{\text{yd}}} \times \frac{1\ \text{km}}{1000\ \cancel{\text{m}}} = \overset{\substack{\text{Result obtained by}\\ \text{rounding only at the}\\ \text{end of the calculation}\\ \downarrow}}{88\ \text{km/h}}$$

Note that all units cancel except the desired kilometers per hour.

See Exercises R.33 through R.35

EXAMPLE R.9

Unit Conversions V

A Japanese car is advertised as having a gas mileage of 15 km/L. Convert this rating to miles per gallon.

Solution

$$\frac{15\ \cancel{\text{km}}}{\cancel{\text{L}}} \times \frac{1000\ \cancel{\text{m}}}{1\ \cancel{\text{km}}} \times \frac{1.094\ \cancel{\text{yd}}}{1\ \cancel{\text{m}}} \times \frac{1\ \text{mi}}{1760\ \cancel{\text{yd}}} \times \frac{1\ \cancel{\text{L}}}{1.06\ \cancel{\text{qt}}} \times \frac{4\ \cancel{\text{qt}}}{1\ \text{gal}} = \overset{\substack{\text{Result obtained by}\\ \text{rounding only at the}\\ \text{end of the calculation}\\ \downarrow}}{35\ \text{mi/gal}}$$

See Exercise R.36

EXAMPLE R.10

Unit Conversions VI

The latest model Corvette has an engine with a displacement of 6.20 L. What is the displacement in units of cubic inches?

Solution

$$6.20\ \text{L} \times \frac{1\ \text{ft}^3}{28.32\ \text{L}} \times \frac{(12\ \text{in})^3}{(1\ \text{ft})^3} = 378\ \text{in}^3$$

Note that the unit factor for conversion of feet to inches must be cubed to accommodate the conversion of ft^3 to in^3.

See Exercise R.38

R.5 › Temperature

Three systems for measuring temperature are widely used: the Celsius scale, the Kelvin scale, and the Fahrenheit scale. The first two temperature systems are used in the physical sciences, and the third is used in many of the engineering sciences. Our purpose here is to define the three temperature scales and show how conversions from one scale to another can be performed. Although these conversions can be carried out routinely on most calculators, we will consider the process in some detail here to illustrate methods of problem solving.

The three temperature scales are defined and compared in Fig. R.5. Note that the size of the temperature unit (the *degree*) is the same for the Kelvin and Celsius scales. The fundamental difference between these two temperature scales is in their zero points. Conversion between these two scales simply requires an adjustment for the different zero points.

$T_K = T_C + 273.15$

$$\text{Temperature (Kelvin)} = \text{temperature (Celsius)} + 273.15$$

or

$T_C = T_K - 273.15$

$$\text{Temperature (Celsius)} = \text{temperature (Kelvin)} - 273.15$$

For example, to convert 300.00 K to the Celsius scale, we do the following calculation:

$$300.00 - 273.15 = 26.85°\text{C}$$

Note that in expressing temperature in Celsius units, the designation °C is used. The degree symbol is not used when writing temperature in terms of the Kelvin scale. The unit of temperature on this scale is called a *kelvin* and is symbolized by the letter K.

Converting between the Fahrenheit and Celsius scales is somewhat more complicated because both the degree sizes and the zero points are different. Thus we need to consider two adjustments: one for degree size and one for the zero point. First, we must account for the difference in degree size. This can be done by reconsidering Fig. R.5. Notice that since 212°F = 100°C and 32°F = 0°C,

$$212 - 32 = 180 \text{ Fahrenheit degrees} = 100 - 0 = 100 \text{ Celsius degrees}$$

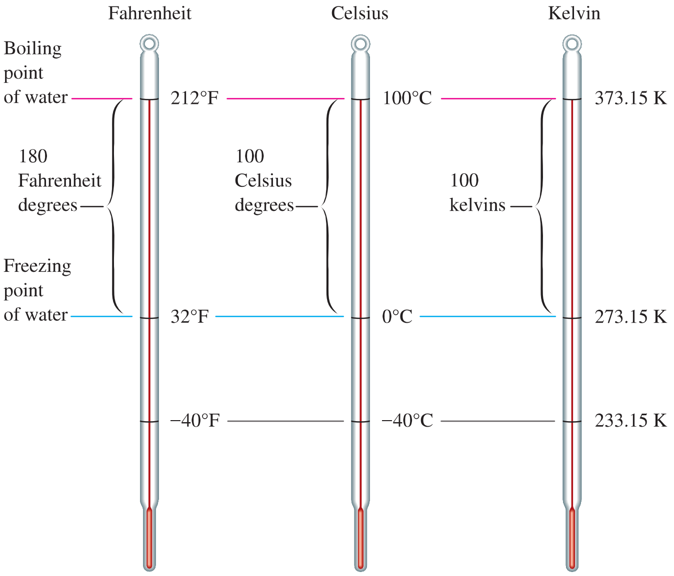


FIGURE R.5
The three major temperature scales.

Thus 180° on the Fahrenheit scale is equivalent to 100° on the Celsius scale, and the unit factor is

$$\frac{180°\text{F}}{100°\text{C}} \quad \text{or} \quad \frac{9°\text{F}}{5°\text{C}}$$

or the reciprocal, depending on the direction in which we need to go.

Next, we must consider the different zero points. Since 32°F = 0°C, we obtain the corresponding Celsius temperature by first subtracting 32 from the Fahrenheit temperature to account for the different zero points. Then the unit factor is applied to adjust for the difference in the degree size. This process is summarized by the equation

$$(T_{\text{F}} - 32°\text{F})\frac{5°\text{C}}{9°\text{F}} = T_{\text{C}} \qquad \textbf{(R.1)}$$

where T_{F} and T_{C} represent a given temperature on the Fahrenheit and Celsius scales, respectively. In the opposite conversion, we first correct for degree size and then correct for the different zero point. This process can be summarized in the following general equation:

$$T_{\text{F}} = T_{\text{C}} \times \frac{9°\text{F}}{5°\text{C}} + 32°\text{F} \qquad \textbf{(R.2)}$$

Understand the process of converting from one temperature scale to another; do not simply memorize the equations.

Equations (R.1) and (R.2) are really the same equation in different forms. See if you can obtain Equation (R.2) by starting with Equation (R.1) and rearranging.

At this point it is worthwhile to weigh the two alternatives for learning to do temperature conversions: You can simply memorize the equations, or you can take the time to learn the differences between the temperature scales and to understand the processes involved in converting from one scale to another. The latter approach may take a little more effort, but the understanding you gain will stick with you much longer than the memorized formulas. This choice also will apply to many of the other chemical concepts. Try to think things through!

INTERACTIVE EXAMPLE R.11

Temperature Conversions I

OWL Sign in to OWL at **www.cengage.com/owl** to view an interactive version of this problem.

Normal body temperature is 98.6°F. Convert this temperature to the Celsius and Kelvin scales.

Solution

Rather than simply using the formulas to solve this problem, we will proceed by thinking it through. The situation is diagramed in Fig. R.6. First, we want to convert 98.6°F to the Celsius scale. The number of Fahrenheit degrees between 32.0°F and 98.6°F is 66.6°F. We must convert this difference to Celsius degrees:

$$66.6°\text{F} \times \frac{5°\text{C}}{9°\text{F}} = 37.0°\text{C}$$

Thus 98.6°F corresponds to 37.0°C.

Now we can convert to the Kelvin scale:

$$T_{\text{K}} = T_{\text{C}} + 273.15 = 37.0 + 273.15 = 310.2 \text{ K}$$

Note that the final answer has only one decimal place (37.0 is limiting).

See Exercises R.39, R.41, and R.42


Peter Steiner

A physician taking the temperature of a patient.

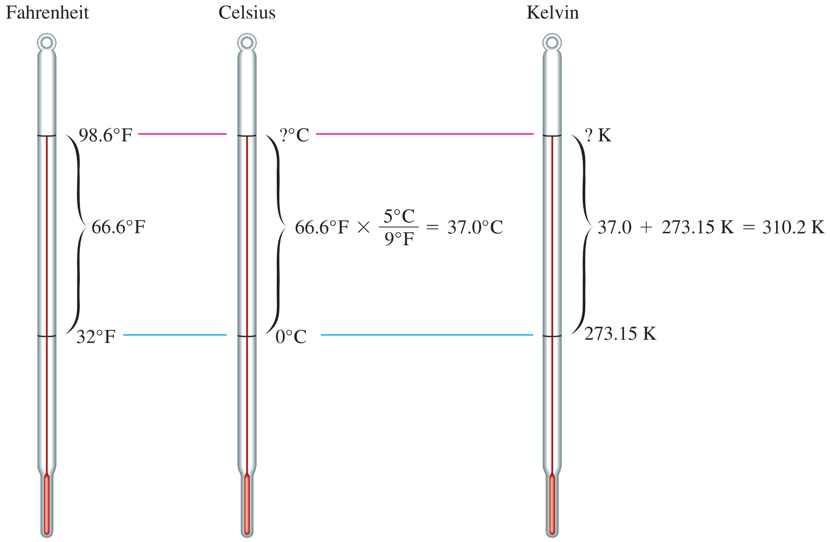


FIGURE R.6
Normal body temperature on the Fahrenheit, Celsius, and Kelvin scales.

EXAMPLE R.12 Temperature Conversions II

One interesting feature of the Celsius and Fahrenheit scales is that −40°C and −40°F represent the same temperature, as shown in Fig. R.5. Verify that this is true.

Solution

The difference between 32°F and −40°F is 72°F. The difference between 0°C and −40°C is 40°C. The ratio of these is

$$\frac{72°\text{F}}{40°\text{C}} = \frac{8 \times 9°\text{F}}{8 \times 5°\text{C}} = \frac{9°\text{F}}{5°\text{C}}$$

as required. Thus −40°C is equivalent to −40°F.

See Challenge Problem R.43

Since, as shown in Example R.12, −40° on both the Fahrenheit and Celsius scales represents the same temperature, this point can be used as a reference point (like 0°C and 32°F) for a relationship between the two scales:

$$\frac{\text{Number of Fahrenheit degrees}}{\text{Number of Celsius degrees}} = \frac{T_F - (-40)}{T_C - (-40)} = \frac{9°\text{F}}{5°\text{C}}$$

$$\frac{T_F + 40}{T_C + 40} = \frac{9°\text{F}}{5°\text{C}} \qquad \text{(R.3)}$$

where T_F and T_C represent the same temperature (but not the same number). This equation can be used to convert Fahrenheit temperatures to Celsius, and vice versa, and may be easier to remember than Equations (R.1) and (R.2).

EXAMPLE R.13 Temperature Conversions III

Liquid nitrogen, which is often used as a coolant for low-temperature experiments, has a boiling point of 77 K. What is this temperature on the Fahrenheit scale?


Richard Megna/Fundamental Photographs

Liquid nitrogen is so cold that water condenses out of the surrounding air, forming a cloud as the nitrogen is poured.

Solution

We will first convert 77 K to the Celsius scale:

$$T_C = T_K - 273.15 = 77 - 273.15 = -196°C$$

To convert to the Fahrenheit scale, we will use Equation (R.3):

$$\frac{T_F + 40}{T_C + 40} = \frac{9°F}{5°C}$$

$$\frac{T_F + 40}{-196°C + 40} = \frac{T_F + 40}{-156°C} = \frac{9°F}{5°C}$$

$$T_F + 40 = \frac{9°F}{5°C}(-156°C) = -281°F$$

$$T_F = -281°F - 40 = -321°F$$

See Exercises R.39, R.41, and R.42

R.6 › Density

A property of matter that is often used by chemists as an "identification tag" for a substance is **density**, the mass of substance per unit volume of the substance:

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

The density of a liquid can be determined easily by weighing an accurately known volume of liquid. This procedure is illustrated in Example R.14.

INTERACTIVE EXAMPLE R.14

OWL Sign in to OWL at **www.cengage.com/owl** to view an interactive version of this problem.

Determining Density

A chemist, trying to identify the main component of a cleaning fluid, finds that 25.00 cm^3 of the substance has a mass of 19.625 g at 20°C. The following are the names and densities of the compounds that might be the main component:

Compound	Density in g/cm^3 at 20°C
Chloroform	1.492
Diethyl ether	0.714
Ethanol	0.789
Isopropyl alcohol	0.785
Toluene	0.867

Which of these compounds is the most likely to be the main component of the cleaner?

Solution

There are two ways of indicating units that occur in the denominator. For example, we can write g/cm^3 or $g\ cm^{-3}$. Although we will use the former system here, the other system is widely used.

To identify the unknown substance, we must determine its density. This can be done by using the definition of density:

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{19.625\text{ g}}{25.00\text{ cm}^3} = 0.7850\text{ g/cm}^3$$

This density corresponds exactly to that of isopropyl alcohol, which is therefore the most likely main component of the cleaner. However, note that the density of ethanol is also very close. To be sure that the compound is isopropyl alcohol, we should run several more density experiments. (In the modern laboratory, many other types of tests could be done to distinguish between these two liquids.)

See Exercises R.49 and R.50

Besides being a tool for the identification of substances, density has many other uses. For example, the liquid in your car's lead storage battery (a solution of sulfuric acid) changes density because the sulfuric acid is consumed as the battery discharges. In a fully charged battery, the density of the solution is about 1.30 g/cm^3. If the density falls below 1.20 g/cm^3, the battery will have to be recharged. Density measurement is also used to determine the amount of antifreeze, and thus the level of protection against freezing, in the cooling system of a car.

The densities of various common substances are given in Table R.5.

R.7 › Classification of Matter

Before we can hope to understand the changes we see going on around us—the growth of plants, the rusting of steel, the aging of people, rain becoming more acidic—we must find out how matter is organized. **Matter**, best defined as anything occupying space and having mass, is the material of the universe. Matter is complex and has many levels of organization. In this section we introduce basic ideas about the structure of matter and its behavior.

We will start by considering the definitions of the fundamental properties of matter. Matter exists in three **states**: solid, liquid, and gas. A *solid* is rigid; it has a fixed volume and shape. A *liquid* has a definite volume but no specific shape; it assumes the shape of its container. A *gas* has no fixed volume or shape; it takes on the shape and volume of its container. In contrast to liquids and solids, which are only slightly compressible, gases are highly compressible; it is relatively easy to decrease the volume of a gas. Molecular-level pictures of the three states of water are given in Fig. R.7. The different properties of ice, liquid water, and steam are determined by the different arrangements of the molecules in these substances. Table R.5 gives the states of some common substances at 20°C and 1 atmosphere of pressure.

Most of the matter around us consists of **mixtures** of pure substances. Wood, gasoline, wine, soil, and air are all mixtures. The main characteristic of a mixture is that it has *variable composition*. For example, wood is a mixture of many substances, the proportions

TABLE R.5 › Densities of Various Common Substances* at 20°C

Substance	Physical State	Density (g/cm^3)
Oxygen	Gas	0.00133
Hydrogen	Gas	0.000084
Ethanol	Liquid	0.789
Benzene	Liquid	0.880
Water	Liquid	0.9982
Magnesium	Solid	1.74
Salt (sodium chloride)	Solid	2.16
Aluminum	Solid	2.70
Iron	Solid	7.87
Copper	Solid	8.96
Silver	Solid	10.5
Lead	Solid	11.34
Mercury	Liquid	13.6
Gold	Solid	19.32

*At 1 atmosphere pressure.

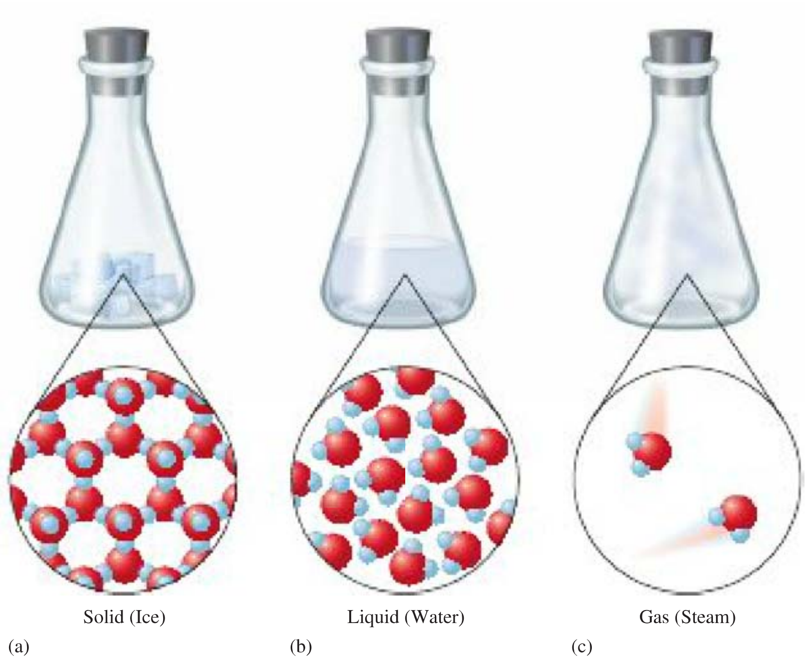


FIGURE R.7
The three states of water (where red spheres represent oxygen atoms and blue spheres represent hydrogen atoms). (a) Solid: the water molecules are locked into rigid positions and are close together. (b) Liquid: the water molecules are still close together but can move around to some extent. (c) Gas: the water molecules are far apart and move randomly.

of which vary depending on the type of wood and where it grows. Mixtures can be classified as **homogeneous** (having visibly indistinguishable parts) or **heterogeneous** (having visibly distinguishable parts).

A homogeneous mixture is called a **solution**. Air is a solution consisting of a mixture of gases. Wine is a complex liquid solution. Brass is a solid solution of copper and zinc. Sand in water and iced tea with ice cubes are examples of heterogeneous mixtures. Heterogeneous mixtures usually can be separated into two or more homogeneous mixtures or pure substances (for example, the ice cubes can be separated from the tea).

Mixtures can be separated into pure substances by physical methods. A **pure substance** is one with constant composition. Water is a good illustration of these ideas. As we will discuss in detail later, pure water is composed solely of H_2O molecules, but the water found in nature (groundwater or the water in a lake or ocean) is really a mixture. Seawater, for example, contains large amounts of dissolved minerals. Boiling seawater produces steam, which can be condensed to pure water, leaving the minerals behind as solids. The dissolved minerals in seawater also can be separated out by freezing the mixture, since pure water freezes out. The processes of boiling and freezing are **physical changes**: When water freezes or boils, it changes its state but remains water; it is still composed of H_2O molecules. A physical change is a change in the form of a substance, not in its chemical composition. A physical change can be used to separate a mixture into pure compounds, but it will not break compounds into elements.

One of the most important methods for separating the components of a mixture is **distillation**, a process that depends on differences in the volatility (how readily substances become gases) of the components. In simple distillation, a mixture is heated in a device such as that shown in Fig. R.8. The most volatile component vaporizes at the lowest temperature, and the vapor passes through a cooled tube (a condenser), where it condenses back into its liquid state.

The term *volatile* refers to the ease with which a substance can be changed to its vapor.

The simple, one-stage distillation apparatus shown in Fig. R.8 works very well when only one component of the mixture is volatile. For example, a mixture of water and sand is easily separated by boiling off the water. Water containing dissolved minerals behaves in much the same way. As the water is boiled off, the minerals remain behind as non-

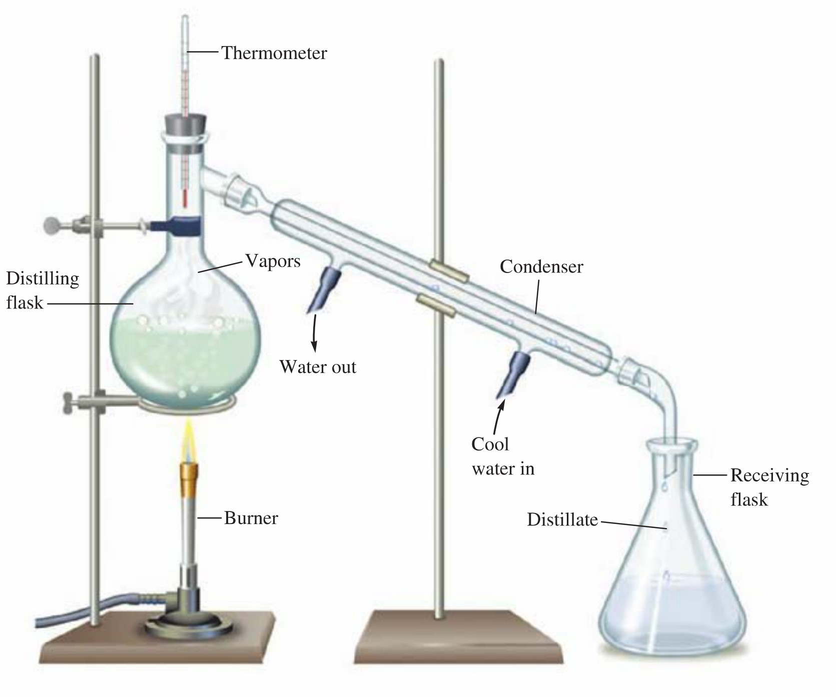


FIGURE R.8
Simple laboratory distillation apparatus. Cool water circulates through the outer portion of the condenser, causing vapors from the distilling flask to condense into a liquid. The nonvolatile component of the mixture remains in the distilling flask.

volatile solids. Simple distillation of seawater using the sun as the heat source is an excellent way to desalinate (remove the minerals from) seawater.

However, when a mixture contains several volatile components, the one-step distillation does not give a pure substance in the receiving flask, and more elaborate methods are required.

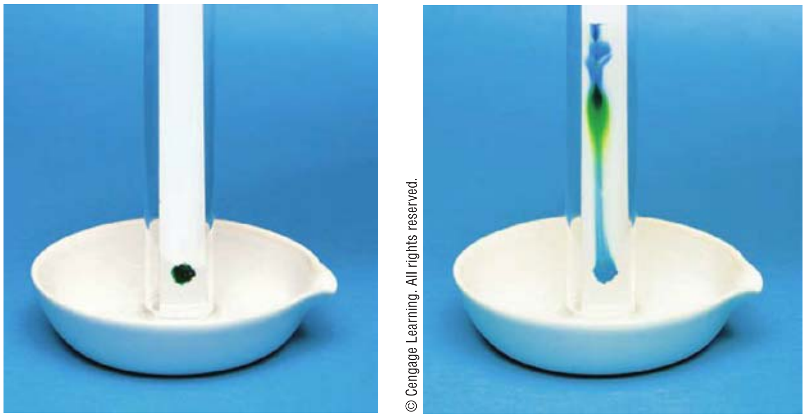
Another method of separation is simple **filtration**, which is used when a mixture consists of a solid and a liquid. The mixture is poured onto a mesh, such as filter paper, which passes the liquid and leaves the solid behind.

A third method of separation is called **chromatography**. Chromatography is the general name applied to a series of methods that employ a system with two *phases* (states) of matter: a mobile phase and a stationary phase. The *stationary phase* is a solid, and the *mobile phase* is either a liquid or a gas. The separation process occurs because the components of the mixture have different affinities for the two phases and thus move through the system at different rates. A component with a high affinity for the mobile phase moves relatively quickly through the chromatographic system, whereas one with a high affinity for the solid phase moves more slowly.

One simple type of chromatography, **paper chromatography**, employs a strip of porous paper, such as filter paper, for the stationary phase. A drop of the mixture to be separated is placed on the paper, which is then dipped into a liquid (the mobile phase) that travels up the paper as though it were a wick (see Fig. R.9). This method of separating a mixture is often used by biochemists, who study the chemistry of living systems.

It should be noted that when a mixture is separated, the absolute purity of the separated components is an ideal. Because water, for example, inevitably comes into contact with other materials when it is synthesized or separated from a mixture, it is never absolutely pure. With great care, however, substances can be obtained in very nearly pure form.

Pure substances are either compounds (combinations of elements) or free elements. A **compound** is a substance with *constant composition* that can be broken down into elements by chemical processes. An example of a chemical process is the electrolysis of water, in which an electric current is passed through water to break it down into the free elements hydrogen and oxygen. This process produces a chemical change because the water molecules have been broken down. The water is gone, and in its place we have the free elements hydrogen and oxygen. A **chemical change** is one in which a given substance becomes a new substance or substances with different properties and different composition. **Elements** are substances that cannot be decomposed into simpler substances by chemical or physical means. We have seen that the matter around us has various levels of organization. Fig. R.10 summarizes our discussion of the organization of matter.



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(a)

(b)

FIGURE R.9
Paper chromatography of ink. (a) A dot of the mixture to be separated is placed at one end of a sheet of porous paper. (b) The paper acts as a wick to draw up the liquid.

Kristen Brochmann/Fundamental Photographs

The element mercury (top left) combines with the element iodine (top right) to form the compound mercuric iodide (bottom). This is an example of a chemical change.

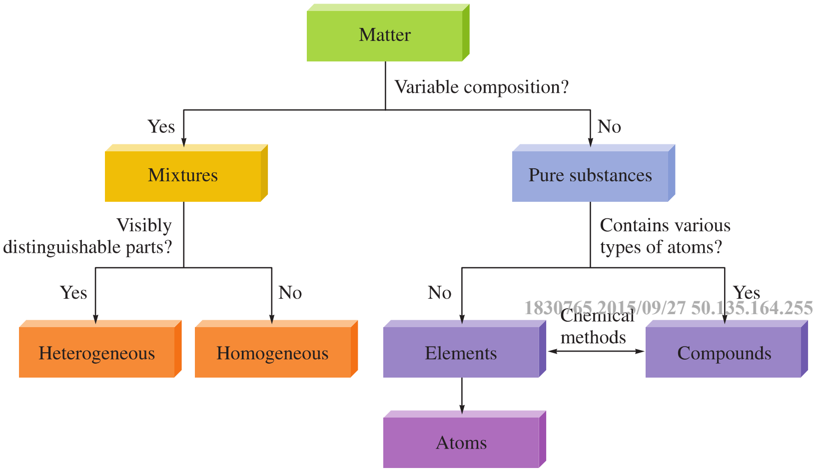

FIGURE R.10
The organization of matter.

OWL and go Chemistry

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Key Terms

Section R.1
SI system
mass
weight

Section R.2
uncertainty
significant figures
accuracy
precision
random error
systematic error

Section R.3
exponential notation

Section R.4
unit factor method
dimensional analysis

Section R.6
density

Section R.7
matter
states (of matter)
homogeneous mixture
heterogeneous mixture
solution
pure substance
physical change
distillation
filtration
chromatography
paper chromatography
compound
chemical change
element

FOR REVIEW

Quantitative observations are called measurements.

- Consist of a number and a unit
- Involve some uncertainty
- Uncertainty is indicated by using significant figures
  - Rules to determine significant figures
  - Calculations using significant figures
- Preferred system is SI

Temperature conversions

- $T_K = T_C + 273$
- $T_C = (T_F - 32°F)\left(\frac{5°C}{9°F}\right)$
- $T_F = T_C\left(\frac{9°F}{5°C}\right) + 32°F$

Density

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

Matter can exist in three states:

- Solid
- Liquid
- Gas

Mixtures can be separated by methods involving only physical changes:

- Distillation
- Filtration
- Chromatography

Compounds can be decomposed to elements only through chemical changes.

A blue question or exercise number indicates that the answer to that question or exercise appears at the back of this book and a solution appears in the *Student Solutions Manual.*

Questions

V denotes Visual Exercises and Questions.

1. What is the difference between random error and systematic error?

2. To determine the volume of a cube, a student measured one of the dimensions of the cube several times. If the true dimension of the cube is 10.62 cm, give an example of four sets of measurements that would illustrate the following.
   a. imprecise and inaccurate data
   b. precise but inaccurate data
   c. precise and accurate data

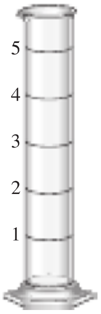
   Give a possible explanation as to why data can be imprecise or inaccurate. What is wrong with saying a set of measurements is imprecise but accurate?

3. A student performed an analysis of a sample for its calcium content and got the following results:

   14.92% 14.91% 14.88% 14.91%

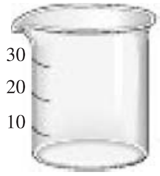
   The actual amount of calcium in the sample is 15.70%. What conclusions can you draw about the accuracy and precision of these results?

4. For each of the following pieces of glassware, provide a sample measurement and discuss the number of significant figures and uncertainty.



a.

b.



c.

5. What are significant figures? Show how to indicate the number one thousand to 1 significant figure, 2 significant figures, 3 significant figures, and 4 significant figures. Why is the answer, to the correct number of significant figures, not 1.0 for the following calculation?

$$\frac{1.5 - 1.0}{0.50} =$$

6. Compare and contrast the multiplication/division significant figure rule to the significant figure rule applied for addition/subtraction mathematical operations.

7. A rule of thumb in designing experiments is to avoid using a result that is the small difference between two large measured quantities. In terms of uncertainties in measurement, why is this good advice?

8. Explain how density can be used as a conversion factor to convert the volume of an object to the mass of the object, and vice versa.

9. When the temperature in degrees Fahrenheit (T_F) is plotted versus the temperature in degrees Celsius (T_C), a straight-line plot results. A straight-line plot also results when T_C is plotted versus T_K (the temperature in kelvins). Reference Appendix A1.3 and determine the slope and y-intercept of each of these two plots.

10. On which temperature scale (°F, °C, or K) does 1 degree represent the smallest change in temperature?

11. Give four examples illustrating each of the following terms.
- **a.** homogeneous mixture
- **b.** heterogeneous mixture
- **c.** compound
- **d.** element
- **e.** physical change
- **f.** chemical change

12. Use molecular-level (microscopic) drawings for each of the following.
- **a.** Show the differences between a gaseous mixture that is a homogeneous mixture of two different compounds, and a gaseous mixture that is a homogeneous mixture of a compound and an element.
- **b.** Show the differences among a gaseous element, a liquid element, and a solid element.

Exercises

Interactive versions of these problems may be assigned in OWL.

In this section, similar exercises are paired.

Significant Figures and Unit Conversions

13. Which of the following are exact numbers?
- **a.** There are *100* cm in 1 m.
- **b.** One meter equals *1.094* yards.
- **c.** We can use the equation

$$°F = \tfrac{9}{5}°C + 32$$

to convert from Celsius to Fahrenheit temperature. Are the numbers $\frac{9}{5}$ and *32* exact or inexact?
- **d.** $\pi = 3.1415927$.

14. Indicate the number of significant figures in each of the following:
- **a.** This book contains more than 1000 pages.
- **b.** A mile is about 5300 ft.
- **c.** A liter is equivalent to 1.059 qt.
- **d.** The population of the United States is approaching 3.1×10^2 million.
- **e.** A kilogram is 1000 g.
- **f.** The Boeing 747 cruises at around 600 mi/h.

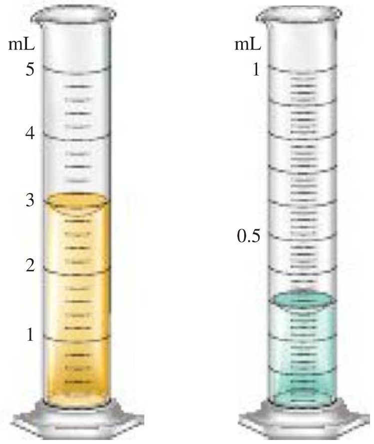
15. How many significant figures are there in each of the following values?
- **a.** 6.07×10^{-15}
- **b.** 0.003840
- **c.** 17.00
- **d.** 8×10^8
- **e.** 463.8052
- **f.** 300
- **g.** 301
- **h.** 300.

16. How many significant figures are in each of the following?
- **a.** 100
- **b.** 1.0×10^2
- **c.** 1.00×10^3
- **d.** 100.
- **e.** 0.0048
- **f.** 0.00480
- **g.** 4.80×10^{-3}
- **h.** 4.800×10^{-3}

17. Round off each of the following numbers to the indicated number of significant digits and write the answer in standard scientific notation.
- **a.** 0.00034159 to three digits
- **b.** 103.351×10^2 to four digits
- **c.** 17.9915 to five digits
- **d.** 3.365×10^5 to three digits

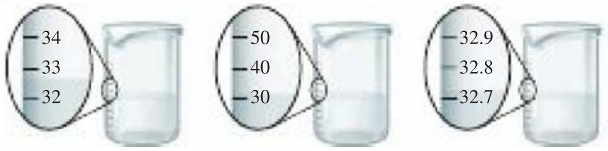
18. Use exponential notation to express the number 385,500 to
- **a.** one significant figure.
- **b.** two significant figures.
- **c.** three significant figures.
- **d.** five significant figures.

V **19.** You have water in each graduated cylinder shown:



You then add both samples to a beaker. How would you write the number describing the total volume? What limits the precision of this number?

V **20.** The beakers shown below have different precisions.



- **a.** Label the amount of water in each of the three beakers to the correct number of significant figures.
- **b.** Is it possible for each of the three beakers to contain the exact same amount of water? If no, why not? If yes, did you report the volumes as the same in part a? Explain.
- **c.** Suppose you pour the water from these three beakers into one container. What should be the volume in the container reported to the correct number of significant figures?

21. Evaluate each of the following and write the answer to the appropriate number of significant figures.
- **a.** $212.2 + 26.7 + 402.09$
- **b.** $1.0028 + 0.221 + 0.10337$
- **c.** $52.331 + 26.01 - 0.9981$
- **d.** $2.01 \times 10^2 + 3.014 \times 10^3$
- **e.** $7.255 - 6.8350$

22. Perform the following mathematical operations, and express each result to the correct number of significant figures.
- **a.** $\dfrac{0.102 \times 0.0821 \times 273}{1.01}$
- **b.** $0.14 \times 6.022 \times 10^{23}$
- **c.** $4.0 \times 10^4 \times 5.021 \times 10^{-3} \times 7.34993 \times 10^2$
- **d.** $\dfrac{2.00 \times 10^6}{3.00 \times 10^{-7}}$

23. Perform the following mathematical operations and express the result to the correct number of significant figures.
- **a.** $\dfrac{2.526}{3.1} + \dfrac{0.470}{0.623} + \dfrac{80.705}{0.4326}$
- **b.** $(6.404 \times 2.91)/(18.7 - 17.1)$
- **c.** $6.071 \times 10^{-5} - 8.2 \times 10^{-6} - 0.521 \times 10^{-4}$
- **d.** $(3.8 \times 10^{-12} + 4.0 \times 10^{-13})/(4 \times 10^{12} + 6.3 \times 10^{13})$
- **e.** $\dfrac{9.5 + 4.1 + 2.8 + 3.175}{4}$

  (Assume that this operation is taking the average of four numbers. Thus 4 in the denominator is exact.)
- **f.** $\dfrac{8.925 - 8.905}{8.925} \times 100$

  (This type of calculation is done many times in calculating a percentage error. Assume that this example is such a calculation; thus 100 can be considered to be an exact number.)

24. Perform the following mathematical operations, and express the result to the correct number of significant figures.
- **a.** $6.022 \times 10^{23} \times 1.05 \times 10^2$
- **b.** $\dfrac{6.6262 \times 10^{-34} \times 2.998 \times 10^8}{2.54 \times 10^{-9}}$
- **c.** $1.285 \times 10^{-2} + 1.24 \times 10^{-3} + 1.879 \times 10^{-1}$
- **d.** $\dfrac{(1.00866 - 1.00728)}{6.02205 \times 10^{23}}$
- **e.** $\dfrac{9.875 \times 10^2 - 9.795 \times 10^2}{9.875 \times 10^2} \times 100$ (100 is exact)
- **f.** $\dfrac{9.42 \times 10^2 + 8.234 \times 10^2 + 1.625 \times 10^3}{3}$ (3 is exact)

25. Perform each of the following conversions.
- **a.** 8.43 cm to millimeters
- **b.** 2.41×10^2 cm to meters
- **c.** 294.5 nm to centimeters
- **d.** 1.445×10^4 m to kilometers
- **e.** 235.3 m to millimeters
- **f.** 903.3 nm to micrometers

26.
- **a.** How many kilograms are in one teragram?
- **b.** How many nanometers are in 6.50×10^2 terameters?
- **c.** How many kilograms are in 25 femtograms?
- **d.** How many liters are in 8.0 cubic decimeters?
- **e.** How many microliters are in one milliliter?
- **f.** How many picograms are in one microgram?

27. Perform the following unit conversions.
- **a.** Congratulations! You and your spouse are the proud parents of a new baby, born while you are studying in a country that uses the metric system. The nurse has informed you that the baby weighs 3.91 kg and measures 51.4 cm. Convert your baby's weight to pounds and ounces and her length to inches (rounded to the nearest quarter inch).
- **b.** The circumference of the earth is 25,000 mi at the equator. What is the circumference in kilometers? in meters?
- **c.** A rectangular solid measures 1.0 m by 5.6 cm by 2.1 dm. Express its volume in cubic meters, liters, cubic inches, and cubic feet.

28. Perform the following unit conversions.
- **a.** 908 oz to kilograms
- **b.** 12.8 L to gallons
- **c.** 125 mL to quarts
- **d.** 2.89 gal to milliliters
- **e.** 4.48 lb to grams
- **f.** 550 mL to quarts

29. Use the following exact conversion factors to perform the stated calculations:

$$5\tfrac{1}{2} \text{ yards} = 1 \text{ rod}$$
$$40 \text{ rods} = 1 \text{ furlong}$$
$$8 \text{ furlongs} = 1 \text{ mile}$$

a. The Kentucky Derby race is 1.25 miles. How long is the race in rods, furlongs, meters, and kilometers?
b. A marathon race is 26 miles, 385 yards. What is this distance in rods, furlongs, meters, and kilometers?

30. Although the preferred SI unit of area is the square meter, land is often measured in the metric system in hectares (ha). One hectare is equal to 10,000 m^2. In the English system, land is often measured in acres (1 acre = 160 rod^2). Use the exact conversions and those given in Exercise 29 to calculate the following.
a. 1 ha = ________ km^2.
b. The area of a 5.5-acre plot of land in hectares, square meters, and square kilometers.
c. A lot with dimensions 120 ft by 75 ft is to be sold for $6500. What is the price per acre? What is the price per hectare?

31. Precious metals and gems are measured in troy weights in the English system:

24 grains = 1 pennyweight (exact)
20 pennyweight = 1 troy ounce (exact)
12 troy ounces = 1 troy pound (exact)
1 grain = 0.0648 gram
1 carat = 0.200 gram

a. The most common English unit of mass is the pound avoirdupois. What is one troy pound in kilograms and in pounds?
b. What is the mass of a troy ounce of gold in grams and in carats?
c. The density of gold is 19.3 g/cm^3. What is the volume of a troy pound of gold?

32. Apothecaries (druggists) use the following set of measures in the English system:

20 grains ap = 1 scruple (exact)
3 scruples = 1 dram ap (exact)
8 dram ap = 1 oz ap (exact)
1 dram ap = 3.888 g

a. Is an apothecary grain the same as a troy grain? (See Exercise 31.)
b. 1 oz ap = ________ oz troy.
c. An aspirin tablet contains 5.00×10^2 mg of active ingredient. What mass in grains ap of active ingredient does it contain? What mass in scruples?
d. What is the mass of 1 scruple in grams?

33. Science fiction often uses nautical analogies to describe space travel. If the starship *U.S.S. Enterprise* is traveling at warp factor 1.71, what is its speed in knots and in miles per hour? (Warp 1.71 = 5.00 times the speed of light; speed of light = 3.00×10^8 m/s; 1 knot = 2000 yd/h, exactly.)

34. The world record for the hundred meter dash is 9.58 s. What is the corresponding average speed in units of m/s, km/h, ft/s, and mi/h? At this speed, how long would it take to run 1.00×10^2 yards?

35. Would a car traveling at a constant speed of 65 km/h violate a 40 mi/h speed limit?

36. You pass a road sign saying "New York 112 km." If you drive at a constant speed of 65 mi/h, how long should it take you to reach New York? If your car gets 28 miles to the gallon, how many liters of gasoline are necessary to travel 112 km?

37. You are in Paris, and you want to buy some peaches for lunch. The sign in the fruit stand indicates that peaches cost 2.45 euros per kilogram. Given that 1 euro is equivalent to approximately $1.32, calculate what a pound of peaches will cost in dollars.

38. Carbon monoxide (CO) detectors sound an alarm when peak levels of carbon monoxide reach 100 parts per million (ppm). This level roughly corresponds to a composition of air that contains 400,000 μg carbon monoxide per cubic meter of air (400,000 μg/m^3). Assuming the dimensions of a room are 18 ft × 12 ft × 8 ft, estimate the mass of carbon monoxide in the room that would register 100 ppm on a carbon monoxide detector.

Temperature

39. Convert the following Fahrenheit temperatures to the Celsius and Kelvin scales.
a. −459°F, an extremely low temperature
b. −40.°F, the answer to a trivia question
c. 68°F, room temperature
d. 7×10^7 °F, temperature required to initiate fusion reactions in the sun

40. A thermometer gives a reading of 96.1°F ± 0.2°F. What is the temperature in °C? What is the uncertainty?

41. Convert the following Celsius temperatures to Kelvin and to Fahrenheit degrees.
a. the temperature of someone with a fever, 39.2°C
b. a cold wintery day, −25°C
c. the lowest possible temperature, −273°C
d. the melting-point temperature of sodium chloride, 801°C

42. Convert the following Kelvin temperatures to Celsius and Fahrenheit degrees.
a. the temperature that registers the same value on both the Fahrenheit and Celsius scales, 233 K
b. the boiling point of helium, 4 K
c. the temperature at which many chemical quantities are determined, 298 K
d. the melting point of tungsten, 3680 K

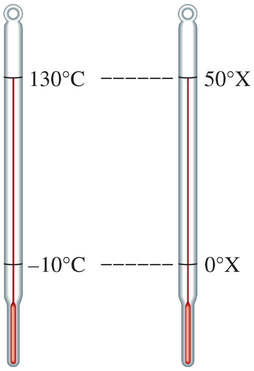
43. At what temperature is the temperature in degrees Fahrenheit equal to twice the temperature in degrees Celsius?

44. The average daytime temperatures on earth and Jupiter are 72°F and 313 K, respectively. Calculate the difference in temperature, in °C, between these two planets.

45. Ethylene glycol is the main component in automobile antifreeze. To monitor the temperature of an auto cooling system, you intend to use a meter that reads from 0 to 100. You devise a new temperature scale based on the approximate melting and boiling points of a typical antifreeze solution (−45°C and 115°C). You wish these points to correspond to 0°A and 100°A, respectively.
a. Derive an expression for converting between °A and °C.
b. Derive an expression for converting between °F and °A.

c. At what temperature would your thermometer and a Celsius thermometer give the same numerical reading?
d. Your thermometer reads 86°A. What is the temperature in °C and in °F?
e. What is a temperature of 45°C in °A?

V **46.** Use the figure below to answer the following questions.



a. Derive the relationship between °C and °X.
b. If the temperature outside is 22.0°C, what is the temperature in units of °X?
c. Convert 58.0°X to units of °C, K, and °F.

Density

47. A material will float on the surface of a liquid if the material has a density less than that of the liquid. Given that the density of water is approximately 1.0 g/mL, will a block of material having a volume of 1.2×10^4 in^3 and weighing 350 lb float or sink when placed in a reservoir of water?

48. For a material to float on the surface of water, the material must have a density less than that of water (1.0 g/mL) and must not react with the water or dissolve in it. A spherical ball has a radius of 0.50 cm and weighs 2.0 g. Will this ball float or sink when placed in water? (*Note*: Volume of a sphere $= \frac{4}{3}\pi r^3$.)

49. A star is estimated to have a mass of 2×10^{36} kg. Assuming it to be a sphere of average radius 7.0×10^5 km, calculate the average density of the star in units of grams per cubic centimeter.

50. A rectangular block has dimensions 2.9 cm × 3.5 cm × 10.0 cm. The mass of the block is 615.0 g. What are the volume and density of the block?

51. Diamonds are measured in carats, and 1 carat = 0.200 g. The density of diamond is 3.51 g/cm^3.
a. What is the volume of a 5.0-carat diamond?
b. What is the mass in carats of a diamond measuring 2.8 mL?

52. Ethanol and benzene dissolve in each other. When 100. mL of ethanol is dissolved in 1.00 L of benzene, what is the mass of the mixture? (See Table R.5.)

53. A sample containing 33.42 g of metal pellets is poured into a graduated cylinder initially containing 12.7 mL of water, causing the water level in the cylinder to rise to 21.6 mL. Calculate the density of the metal.

54. The density of pure silver is 10.5 g/cm^3 at 20°C. If 5.25 g of pure silver pellets is added to a graduated cylinder containing 11.2 mL of water, to what volume level will the water in the cylinder rise?

55. In each of the following pairs, which has the greater mass? (See Table R.5.)
a. 1.0 kg of feathers or 1.0 kg of lead
b. 1.0 mL of mercury or 1.0 mL of water
c. 19.3 mL of water or 1.00 mL of gold
d. 75 mL of copper or 1.0 L of benzene

56. a. Calculate the mass of ethanol in 1.50 qt of ethanol. (See Table R.5.)
b. Calculate the mass of mercury in 3.5 in^3 of mercury. (See Table R.5.)

57. In each of the following pairs, which has the greater volume?
a. 1.0 kg of feathers or 1.0 kg of lead
b. 100 g of gold or 100 g of water
c. 1.0 L of copper or 1.0 L of mercury

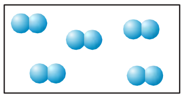
58. Using Table R.5, calculate the volume of 25.0 g of each of the following substances at 1 atm.
a. hydrogen gas
b. water
c. iron
Chapter 7 discusses the properties of gases. One property unique to gases is that they contain mostly empty space. Explain, using the results of your calculations.

59. The density of osmium (the densest metal) is 22.57 g/cm^3. If a 1.00-kg rectangular block of osmium has two dimensions of 4.00 cm × 4.00 cm, calculate the third dimension of the block.

60. A copper wire (density = 8.96 g/cm^3) has a diameter of 0.25 mm. If a sample of this copper wire has a mass of 22 g, how long is the wire?

Classification and Separation of Matter

V **61.** Match each description below with the following microscopic pictures. More than one picture may fit each description. A picture may be used more than once or not used at all.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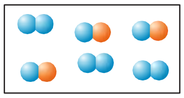
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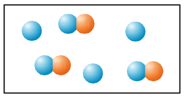
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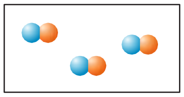
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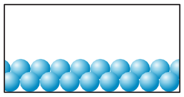
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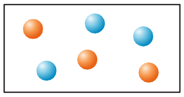

vi

a. a gaseous compound
b. a mixture of two gaseous elements
c. a solid element
d. a mixture of a gaseous element and a gaseous compound

62. Define the following terms: solid, liquid, gas, pure substance, element, compound, homogeneous mixture, heterogeneous mixture, solution, chemical change, physical change.

63. What is the difference between homogeneous and heterogeneous matter? Classify each of the following as homogeneous or heterogeneous.
    - **a.** a door
    - **b.** the air you breathe
    - **c.** a cup of coffee (black)
    - **d.** the water you drink
    - **e.** salsa
    - **f.** your lab partner

64. Classify the following mixtures as homogeneous or as heterogeneous.
    - **a.** potting soil
    - **b.** white wine
    - **c.** your sock drawer
    - **d.** window glass
    - **e.** granite

65. Classify each of the following as a mixture or a pure substance.
    - **a.** water
    - **b.** blood
    - **c.** the oceans
    - **d.** iron
    - **e.** brass
    - **f.** uranium
    - **g.** wine
    - **h.** leather
    - **i.** table salt

    Of the pure substances, which are elements and which are compounds?

66. Suppose a teaspoon of magnesium filings and a teaspoon of powdered sulfur are placed together in a metal beaker. Would this constitute a mixture or a pure substance? Suppose the magnesium filings and sulfur are heated so they react with each other, forming magnesium sulfide. Would this still be a "mixture"? Why or why not?

67. If a piece of hard white blackboard chalk is heated strongly in a flame, the mass of the piece of chalk will decrease, and eventually the chalk will crumble into a fine white dust. Does this change suggest that the chalk is composed of an element or a compound?

68. During a very cold winter, the temperature may remain below freezing for extended periods. However, fallen snow can still disappear, even though it cannot melt. This is possible because a solid can vaporize directly, without passing through the liquid state. Is this process (sublimation) a physical or a chemical change?

69. Classify the following as physical or chemical changes.
    - **a.** Moth balls gradually vaporize in a closet.
    - **b.** Hydrofluoric acid attacks glass and is used to etch calibration marks on glass laboratory utensils.
    - **c.** A French chef making a sauce with brandy is able to boil off the alcohol from the brandy, leaving just the brandy flavoring.
    - **d.** Chemistry majors sometimes get holes in the cotton jeans they wear to lab because of acid spills.

70. The properties of a mixture are typically averages of the properties of its components. The properties of a compound may differ dramatically from the properties of the elements that combine to produce the compound. For each process described below, state whether the material being discussed is most likely a mixture or a compound, and state whether the process is a chemical change or a physical change.
    - **a.** An orange liquid is distilled, resulting in the collection of a yellow liquid and a red solid.
    - **b.** A colorless, crystalline solid is decomposed, yielding a pale yellow-green gas and a soft, shiny metal.
    - **c.** A cup of tea becomes sweeter as sugar is added to it.

› Additional Exercises

71. A children's pain relief elixir contains 80. mg acetaminophen per 0.50 teaspoon. The dosage recommended for a child who weighs between 24 and 35 lb is 1.5 teaspoons. What is the range of acetaminophen dosages, expressed in mg acetaminophen/kg body weight, for children who weigh between 24 and 35 lb?

72. The active ingredient of aspirin tablets is acetylsalicylic acid, which has a density of 1.4 g/cm^3. In a lab class, a student used paper chromatography to isolate another common ingredient of headache remedies. The isolated sample had a mass of 0.384 g and a volume of 0.32 cm^3. Given the data in the following table, what was the other ingredient in the headache remedy?

Density Values for Potential Headache Remedies

Compound	Density (g/cm^3)
White table sugar	0.70
Caffeine	1.2
Acetylsalicylic acid	1.4
Sodium chloride	2.2

73. Mercury poisoning is a debilitating disease that is often fatal. In the human body, mercury reacts with essential enzymes, leading to irreversible inactivity of these enzymes. If the amount of mercury in a polluted lake is 0.4 μg Hg/mL, what is the total mass in kilograms of mercury in the lake? (The lake has a surface area of 100 mi^2 and an average depth of 20 ft.)

74. Which of the following are chemical changes? Which are physical changes?
    - **a.** the cutting of food
    - **b.** interaction of food with saliva and digestive enzymes
    - **c.** proteins being broken down into amino acids
    - **d.** complex sugars being broken down into simple sugars
    - **e.** making maple syrup by heating maple sap to remove water through evaporation
    - **f.** DNA unwinding

75. The contents of one 40-lb bag of topsoil will cover 10. square feet of ground to a depth of 1.0 inch. How many bags are needed to cover a plot that measures 200. by 300. m to a depth of 4.0 cm?

76. In the opening scenes of the movie *Raiders of the Lost Ark*, Indiana Jones tries to remove a gold idol from a booby-trapped pedestal. He replaces the idol with a bag of sand of approximately equal volume. (Density of gold = 19.32 g/cm^3; density of sand ≈ 2 g/cm^3.)
    - **a.** Did he have a reasonable chance of not activating the mass-sensitive booby trap?
    - **b.** In a later scene he and an unscrupulous guide play catch with the idol. Assume that the volume of the idol is about 1.0 L. If it were solid gold, what mass would the idol have? Is playing catch with it plausible?

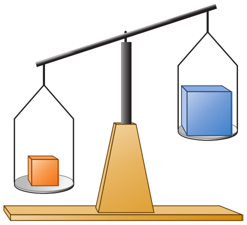
77. A parsec is an astronomical unit of distance where 1 parsec = 3.26 light years (1 light year equals the distance traveled by light in one year). If the speed of light is 186,000 mi/s, calculate the distance in meters of an object that travels 9.6 parsecs.

78. You are driving 65 mi/h and take your eyes off the road for "just a second." What distance (in feet) do you travel in this time?

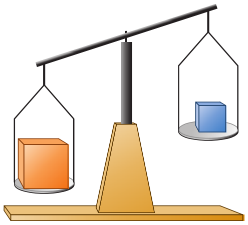
79. A column of liquid is found to expand linearly on heating 5.25 cm for a 10.0°F rise in temperature. If the initial temperature of the liquid is 98.6°F, what will the final temperature be in °C if the liquid has expanded by 18.5 cm?

80. A 25.00-g sample of a solid is placed in a graduated cylinder and then the cylinder is filled to the 50.0-mL mark with benzene. The mass of benzene and solid together is 58.80 g. Assuming that the solid is insoluble in benzene and that the density of benzene is 0.880 g/cm^3, calculate the density of the solid.

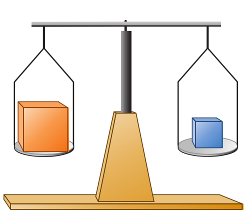
V 81. For each of the following, decide whether the orange block is more dense, the blue block is more dense, or it cannot be determined. Explain your answers.



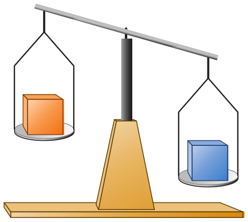
a.



b.



c.



d.

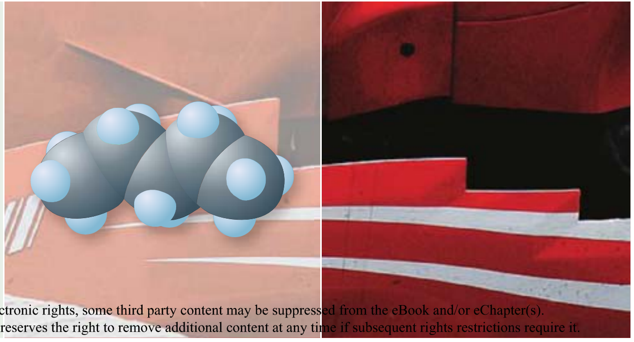
82. According to the *Official Rules of Baseball*, a baseball must have a circumference not more than 9.25 in or less than 9.00 in and a mass not more than 5.25 oz or less than 5.00 oz. What range of densities can a baseball be expected to have? Express this range as a single number with an accompanying uncertainty limit.

83. The density of an irregularly shaped object was determined as follows. The mass of the object was found to be 28.90 g ± 0.03 g. A graduated cylinder was partially filled with water. The reading of the level of the water was 6.4 cm^3 ± 0.1 cm^3. The object was dropped in the cylinder, and the level of the water rose to 9.8 cm^3 ± 0.1 cm^3. What is the density of the object, with appropriate error limits? (See Appendix 1.5.)

84. The chemist in Example R.14 did some further experiments. She found that the pipet used to measure the volume of the cleaner is accurate to ±0.03 cm^3. The mass measurement is accurate to ±0.002 g. Are these measurements sufficiently precise for the chemist to distinguish between isopropyl alcohol and ethanol?

chapter one

Chemical Foundations



Chemistry is involved in the structure, tires, and fuel of a high-performance race car. The molecule shown is pentane, which is a component of gasoline.

Alamy

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go Chemistry Download mini lecture videos for key concept review and exam prep from OWL or purchase them from **www.cengagebrain.com**

When you start your car, do you think about chemistry? Probably not, but you should. The power to start your car is furnished by a lead storage battery. How does this battery work, and what does it contain? When a battery goes dead, what does that mean? If you use a friend's car to "jump start" your car, did you know that your battery could explode? How can you avoid such an unpleasant possibility? What is in the gasoline that you put in your tank, and how does it furnish the energy to drive to school? What is the vapor that comes out of the exhaust pipe, and why does it cause air pollution? Your car's air conditioner might have a substance in it that is leading to the destruction of the ozone layer in the upper atmosphere. What are we doing about that? And why is the ozone layer important anyway?

All these questions can be answered by understanding some chemistry. In fact, we'll consider the answers to all these questions in this text.

Chemistry is around you all the time. You are able to read and understand this sentence because chemical reactions are occurring in your brain. The food you ate for breakfast or lunch is now furnishing energy through chemical reactions. Trees and grass grow because of chemical changes.

Chemistry also crops up in some unexpected places. When archaeologist Luis Alvarez was studying in college, he probably didn't realize that the chemical elements iridium and niobium would make him very famous when they helped him solve the problem of the disappearing dinosaurs. For decades scientists had wrestled with the mystery of why the dinosaurs, after ruling the earth for millions of years, suddenly became extinct 65 million years ago. In studying core samples of rocks dating back to that period, Alvarez and his coworkers recognized unusual levels of iridium and niobium in these samples—levels much more characteristic of extraterrestrial bodies than of the earth. On the basis of these observations, Alvarez hypothesized that a large meteor hit the earth 65 million years ago, changing atmospheric conditions so much that the dinosaurs' food couldn't grow, and they died—almost instantly in the geologic timeframe.

Chemistry is also important to historians. Did you realize that lead poisoning probably was a significant contributing factor to the decline of the Roman Empire? The Romans had high exposure to lead from lead-glazed pottery, lead water pipes, and a sweetening syrup called *sapa* that was prepared by boiling down grape juice in lead-lined vessels. It turns out that one reason for sapa's sweetness was lead acetate ("sugar of lead") that formed as the juice was cooked down. Lead poisoning, with its symptoms of lethargy and mental malfunctions, certainly could have contributed to the demise of the Roman society.

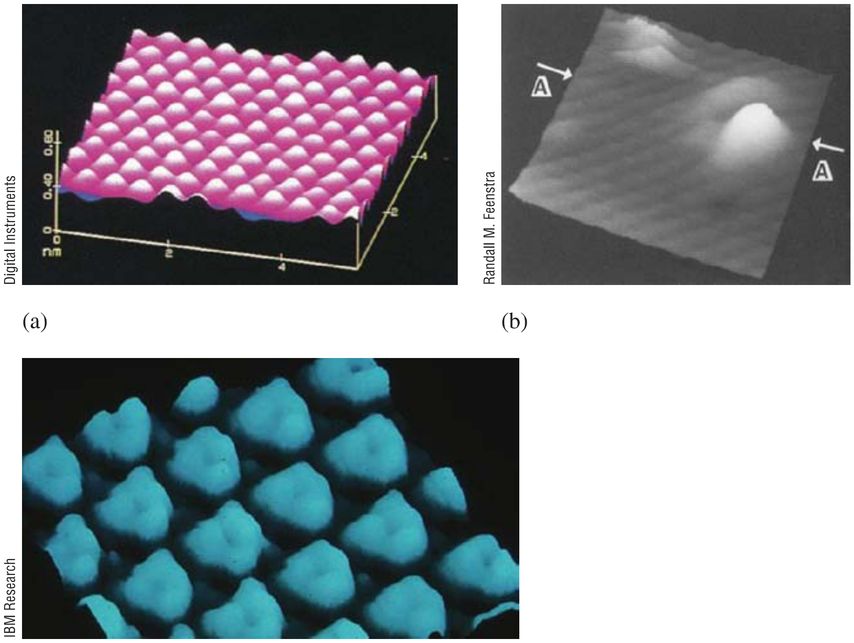
Chemistry is also apparently very important in determining a person's behavior. Various studies have shown that many personality disorders can be linked directly to imbalances of trace elements in the body. For example, studies on the inmates at Stateville Prison in Illinois have linked low cobalt levels with violent behavior. Lithium salts have been shown to be very effective in controlling the effects of manic depressive disease, and you've probably at some time in your life felt a special "chemistry" for another person. Studies suggest there is literally chemistry going on between two people who are attracted to each other. "Falling in love" apparently causes changes in the chemistry of the brain; chemicals are produced that give that "high" associated with a new relationship. Unfortunately, these chemical effects seem to wear off over time, even if the relationship persists and grows.

The importance of chemistry in the interactions of people should not really surprise us, since we know that insects communicate by emitting and receiving chemical signals via molecules called *pheromones*. For example, ants have a very complicated set of chemical signals to signify food sources, danger, and so forth. Also, various female sex attractants have been isolated and used to lure males into traps to control insect populations. It would not be surprising if humans also emitted chemical signals that we were not aware of on a conscious level. Thus chemistry is pretty interesting and pretty important. The main goal of this text is to help you understand the concepts of chemistry so that you can better appreciate the world around you and can be more effective in whatever career you choose.

1.1 › Chemistry: An Overview

Since the time of the ancient Greeks, people have wondered about the answer to the question, What is matter made of? For a long time humans have believed that matter is composed of atoms, and in the previous three centuries we have collected much indirect evidence to support this belief. Very recently, something exciting has happened—for the first time we can "see" individual atoms. Of course, we cannot see atoms with the naked eye but must use a special microscope called a *scanning tunneling microscope* (STM). Although we will not consider the details of its operation here, the STM uses an electron current from a tiny needle to probe the surface of a substance. The STM pictures of several substances are shown in Fig. 1.1. Notice how the atoms are connected to one another by "bridges," which, as we will see, represent the electrons that interconnect atoms.

So, at this point, we are fairly sure that matter consists of individual atoms. The nature of these atoms is quite complex, and the components of atoms don't behave much like the objects we see in the world of our experience. We call this world the *macroscopic world*—the world of cars, tables, baseballs, rocks, oceans, and so forth. One of the main jobs of a scientist is to delve into the macroscopic world and discover its "parts." For example, when you view a beach from a distance, it looks like a continuous solid substance. As you get closer, you see that the beach is really made up of individual grains of sand. As we examine these grains of sand, we find they are composed of silicon and oxygen atoms connected to each other to form intricate shapes (see Fig. 1.2). One of the main challenges of chemistry is to understand the connection between the macroscopic



Digital Instruments

(a)

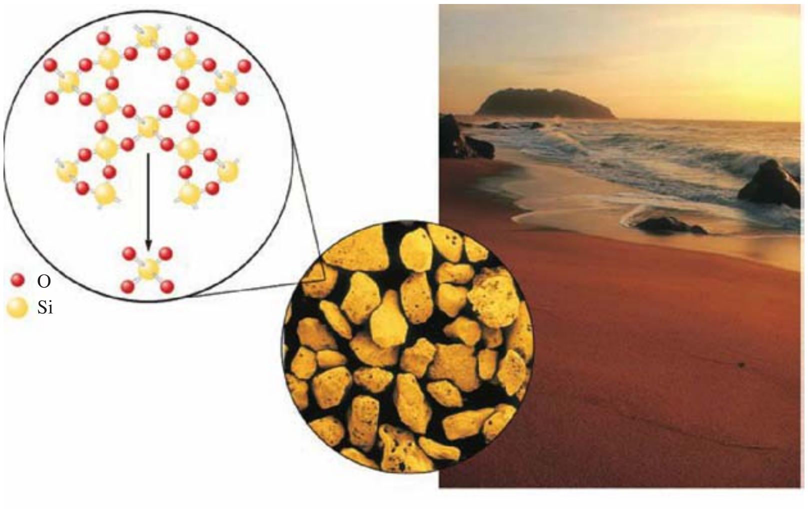
Randall M. Feenstra

(b)

IBM Research

(c)

FIGURE 1.1
(a) The surface of a single grain of table salt. (b) An oxygen atom (indicated by arrow) on a gallium arsenide surface. (c) Scanning tunneling microscope image showing rows of ring-shaped clusters of benzene molecules on a rhodium surface. Each cluster represents a benzene molecule.



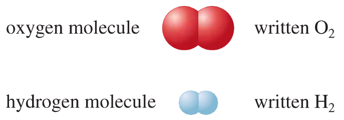
Chuck Place Photography
Inset photo: Jeremy Burgess/SPL/Photo Researchers, Inc.

FIGURE 1.2
Sand on a beach looks uniform from a distance, but up close the irregular sand grains are visible, and each grain is composed of tiny atoms.

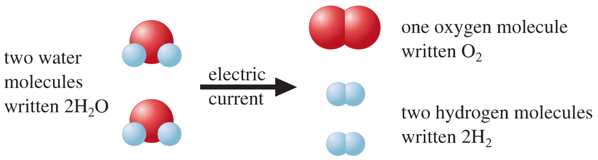
world that we experience and the *microscopic world* of atoms and molecules. To truly understand chemistry, you must learn to think on the atomic level. We will spend much time in this text helping you learn to do that.

One of the amazing things about our universe is that the tremendous variety of substances we find there results from only about 100 different kinds of atoms. You can think of these approximately 100 atoms as the letters in an alphabet, out of which all the "words" in the universe are made. It is the way the atoms are organized in a given substance that determines the properties of that substance. For example, water, one of the most common and important substances on earth, is composed of two types of atoms: hydrogen and oxygen. There are two hydrogen atoms and one oxygen atom bound together to form the water molecule:

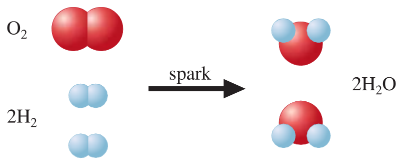
When an electric current passes through it, water is decomposed to hydrogen and oxygen. These *chemical elements* themselves exist naturally as diatomic (two-atom) molecules:



We can represent the decomposition of water to its component elements, hydrogen and oxygen, as follows:



Notice that it takes two molecules of water to furnish the right number of oxygen and hydrogen atoms to allow for the formation of the two-atom molecules. This reaction explains why the battery in your car can explode if you jump start it improperly. When you hook up the jumper cables, current flows through the dead battery, which contains water (and other things), and causes hydrogen and oxygen to form by decomposition of some of the water. A spark can cause this accumulated hydrogen and oxygen to explode, forming water again.



This example illustrates two of the fundamental concepts of chemistry: (1) matter is composed of various types of atoms, and (2) one substance changes to another by reorganizing the way the atoms are attached to each other.

These are core ideas of chemistry, and we will have much more to say about them.

Science: A Process for Understanding Nature and Its Changes

How do you tackle the problems that confront you in real life? Think about your trip to school. If you live in a city, traffic is undoubtedly a problem you confront daily. How do you decide the best way to drive to school? If you are new in town, you first get a map and look at the possible ways to make the trip. Then you might collect information from people who know the area about the advantages and disadvantages of various routes. On the basis of this information, you probably try to predict the best route. However, you can find the best route only by trying several of them and comparing the results. After a few experiments with the various possibilities, you probably will be able to select the best way. What you are doing in solving this everyday problem is applying the same process that scientists use to study nature. The first thing you did was collect relevant data. Then you made a prediction, and then you tested it by trying it out. This process contains the fundamental elements of science:

1. Making observations (collecting data)
2. Suggesting a possible explanation (formulating a hypothesis)
3. Doing experiments to test the possible explanation (testing the hypothesis)

Scientists call this process the *scientific method*. We will discuss it in more detail in the next section. One of life's most important activities is solving problems—not "plug and chug" exercises, but real problems—problems that have new facets to them, that involve things you may have never confronted before. The more creative you are at solving these problems, the more effective you will be in your career and your personal life. Part of the reason for learning chemistry, therefore, is to become a better problem solver. Chemists are usually excellent problem solvers, because to master chemistry, you have to master the scientific approach. Chemical problems are frequently very complicated—there is usually no neat and tidy solution. Often it is difficult to know where to begin.

1.2 › The Scientific Method

Science is a framework for gaining and organizing knowledge. Science is not simply a set of facts but also a plan of action—a *procedure* for processing and understanding certain types of information. Scientific thinking is useful in all aspects of life, but in this text we will use it to understand how the chemical world operates. As we have said in our previous discussion, the process that lies at the center of scientific inquiry is called the **scientific method**. There are actually many scientific methods, depending on the nature of the specific problem under study and on the particular investigator involved. However, it is useful to consider the following general framework for a generic scientific method (see Fig. 1.3).

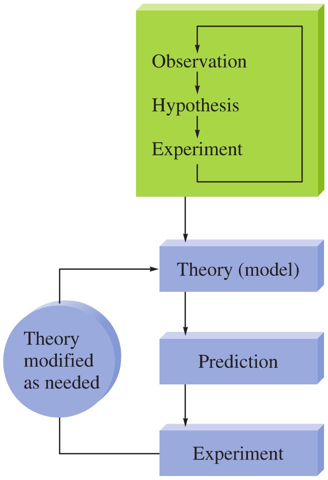


FIGURE 1.3
The fundamental steps of the scientific method.

Steps in the Scientific Method

1. *Making observations*. Observations may be *qualitative* (the sky is blue; water is a liquid) or *quantitative* (water boils at 100°C; a certain chemistry book weighs 2 kilograms). A qualitative observation does not involve a number. A quantitative observation (called a **measurement**) involves both a number and a unit.
2. *Formulating hypotheses*. A **hypothesis** is a *possible* explanation for an observation.
3. *Performing experiments*. An experiment is carried out to test a hypothesis. This involves gathering new information that enables a scientist to decide whether the hypothesis is valid—that is, whether it is supported by the new information learned from the experiment. Experiments always produce new observations, and this brings the process back to the beginning again.

To understand a given phenomenon, an investigator repeats these steps many times, gradually accumulating the knowledge necessary to provide a possible explanation of the phenomenon.

Scientific Models

Once a set of hypotheses that agree with the various observations is obtained, the hypotheses are assembled into a theory. A **theory**, which is often called a **model**, is a set of tested hypotheses that gives an overall explanation of some natural phenomenon.

It is very important to distinguish between observations and theories. An observation is something that is witnessed and can be recorded. A theory is an *interpretation*—a possible explanation of *why* nature behaves in a particular way. Theories inevitably change as more information becomes available. For example, the motions of the sun and stars have remained virtually the same over the thousands of years during which humans have been observing them, but our explanations—our theories—for these motions have changed greatly since ancient times.

The point is that scientists do not stop asking questions just because a given theory seems to account satisfactorily for some aspect of natural behavior. They continue doing experiments to refine or replace the existing theories. This is generally done by using the currently accepted theory to make a prediction and then performing an experiment (making a new observation) to see whether the results bear out this prediction.

Always remember that theories (models) are human inventions. They represent attempts to explain observed natural behavior in terms of human experiences. A theory is actually an educated guess. We must continue to do experiments and to refine our theories (making them consistent with new knowledge) if we hope to approach a more nearly complete understanding of nature.

As scientists observe nature, they often see that the same observation applies to many different systems. For example, studies of innumerable chemical changes have shown that the total observed mass of the materials involved is the same before and after the change. Such generally observed behavior is formulated into a statement called a **natural law**. For example, the observation that the total mass of materials is not affected by a chemical change in those materials is called the **law of conservation of mass**.

Note the difference between a natural law and a theory. A natural law is a summary of observed (measurable) behavior, whereas a theory is an explanation of behavior. A law summarizes what happens; a theory (model) is an attempt to explain why it happens.

In this section we have described the scientific method as it might ideally be applied (see Fig. 1.4). However, it is important to remember that science does not always progress smoothly and efficiently. For one thing, hypotheses and observations are not totally independent of each other, as we have assumed in the description of the idealized scientific method. The coupling of observations and hypotheses occurs because once we begin to proceed down a given theoretical path, our hypotheses are unavoidably couched in the

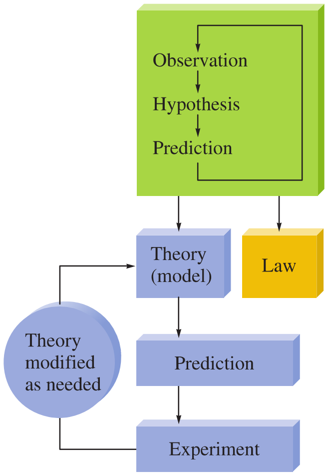
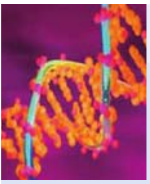


FIGURE 1.4
The various parts of the scientific method.



CHEMICAL CONNECTIONS

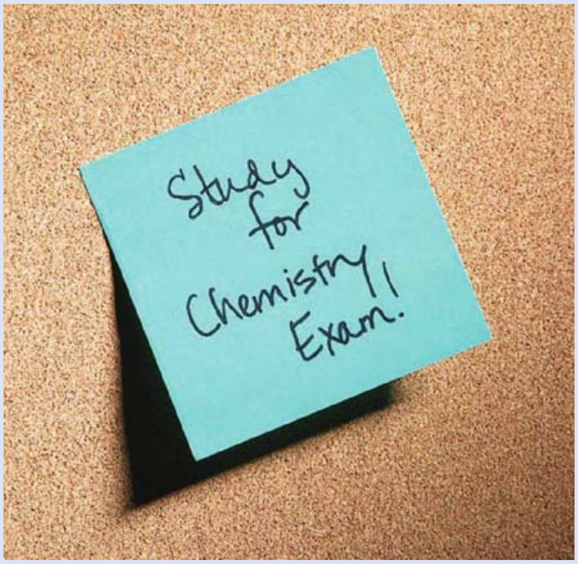
A Note-able Achievement

Post-it Notes, a product of the 3M Corporation, revolutionized casual written communications and personal reminders. Introduced in the United States in 1980, these sticky-but-not-too-sticky notes have now found countless uses in offices, cars, and homes throughout the world.

The invention of sticky notes occurred over a period of about 10 years and involved a great deal of serendipity. The adhesive for Post-it Notes was discovered by Dr. Spencer F. Silver of 3M in 1968. Silver found that when an acrylate polymer material was made in a particular way, it formed cross-linked microspheres. When suspended in a solvent and sprayed on a sheet of paper, this substance formed a "sparse monolayer" of adhesive after the solvent evaporated. Scanning electron microscope images of the adhesive show that it has an irregular surface, a little like the surface of a gravel road. In contrast, the adhesive on cellophane tape looks smooth and uniform, like a superhighway. The bumpy surface of Silver's adhesive caused it to be sticky but not so sticky to produce permanent adhesion, because the number of contact points between the binding surfaces was limited.

When he invented this adhesive, Silver had no specific ideas for its use, so he spread the word of his discovery to his fellow employees at 3M to see if anyone had an application for it. In addition, over the next several years development was carried out to improve the adhesive's properties. It was not until 1974 that the idea for Post-it Notes popped up. One Sunday, Art Fry, a chemical engineer for 3M, was singing in his church choir when he became annoyed that the bookmark in his hymnal kept falling out. He thought to himself that it would be nice if the bookmark were sticky enough to stay in place but not so sticky that it couldn't be moved. Luckily, he remembered Silver's glue—and the Post-it Note was born.

For the next three years Fry worked to overcome the manufacturing obstacles associated with the product. By 1977 enough Post-it Notes were being produced to supply 3M's corporate headquarters, where the employees quickly became addicted to their many uses. Post-it Notes are now available in 62 colors and 25 shapes.



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In the years since their introduction, 3M has heard some remarkable stories connected to the use of these notes. For example, a Post-it Note was applied to the nose of a corporate jet, where it was intended to be read by the plane's Las Vegas ground crew. Someone forgot to remove it, however. The note was still on the nose of the plane when it landed in Minneapolis, having survived a take-off and landing and speeds of 500 miles per hour at temperatures as low as −56°F. Stories on the 3M Web site also describe how a Post-it Note on the front door of a home survived the 140-mile-per-hour winds of Hurricane Hugo and how a foreign official accepted Post-it Notes in lieu of cash when a small bribe was needed to cut through bureaucratic hassles.

Post-it Notes have definitely changed the way we communicate and remember things.

language of that theory. In other words, we tend to see what we expect to see and often fail to notice things that we do not expect. Thus the theory we are testing helps us because it focuses our questions. However, at the very same time, this focusing process may limit our ability to see other possible explanations.

It is also important to keep in mind that scientists are human. They have prejudices; they misinterpret data; they become emotionally attached to their theories and thus lose objectivity; and they play politics. Science is affected by profit motives, budgets, fads, wars, and religious beliefs. Galileo, for example, was forced to recant his astronomical observations in the face of strong religious resistance. Lavoisier, the father of modern chemistry, was beheaded because of his political affiliations. Great progress in the chem-

The Granger Collection, New York

Robert Boyle (1627–1691) was born in Ireland. He became especially interested in experiments involving air and developed an air pump with which he produced evacuated cylinders. He used these cylinders to show that a feather and a lump of lead fall at the same rate in the absence of air resistance and that sound cannot be produced in a vacuum. His most famous experiments involved careful measurements of the volume of a gas as a function of pressure. In his book *The Skeptical Chymist*, Boyle urged that the ancient view of elements as mystical substances should be abandoned and that an element should instead be defined as anything that cannot be broken down into simpler substances. This conception was an important step in the development of modern chemistry.

istry of nitrogen fertilizers resulted from the desire to produce explosives to fight wars. The progress of science is often affected more by the frailties of humans and their institutions than by the limitations of scientific measuring devices. The scientific methods are only as effective as the humans using them. They do not automatically lead to progress.

1.3 › The Early History of Chemistry

Chemistry has been important since ancient times. The processing of natural ores to produce metals for ornaments and weapons and the use of embalming fluids are just two applications of chemical phenomena that were utilized prior to 1000 B.C.

The Greeks were the first to try to explain why chemical changes occur. By about 400 B.C. they had proposed that all matter was composed of four fundamental substances: fire, earth, water, and air. The Greeks also considered the question of whether matter is continuous, and thus infinitely divisible into smaller pieces, or composed of small, indivisible particles. Supporters of the latter position were Demokritos* of Abdera (*c*. 460–*c*. 370 B.C.) and Leucippos, who used the term *atomos* (which later became *atoms*) to describe these ultimate particles. However, because the Greeks had no experiments to test their ideas, no definitive conclusion could be reached about the divisibility of matter.

The next 2000 years of chemical history were dominated by a pseudoscience called *alchemy*. Some alchemists were mystics and fakes who were obsessed with the idea of turning cheap metals into gold. However, many alchemists were serious scientists, and this period saw important advances: the alchemists discovered several elements and learned to prepare the mineral acids.

The foundations of modern chemistry were laid in the sixteenth century with the development of systematic metallurgy (extraction of metals from ores) by a German, Georg Bauer (1494–1555), and the medicinal application of minerals by a Swiss alchemist/physician known as Paracelsus (full name: Philippus Theophrastus Bombastus von Hohenheim [1493–1541]).

The first "chemist" to perform truly quantitative experiments was Robert Boyle (1627–1691), who carefully measured the relationship between the pressure and volume of air. When Boyle published his book *The Skeptical Chymist* in 1661, the quantitative sciences of physics and chemistry were born. In addition to his findings on the quantitative behavior of gases, Boyle's other major contribution to chemistry consisted of his ideas about the chemical elements. Boyle held no preconceived notion about the number of elements. In his view, a substance was an element unless it could be broken down into two or more simpler substances. As Boyle's experimental definition of an element became generally accepted, the list of known elements began to grow, and the Greek system of four elements finally died. Although Boyle was an excellent scientist, he was not always right. For example, he clung to the alchemists' views that metals were not true elements and that a way would eventually be found to change one metal into another.

The phenomenon of combustion evoked intense interest in the seventeenth and eighteenth centuries. The German chemist Georg Stahl (1660–1734) suggested that a substance he called "phlogiston" flowed out of burning material. Stahl postulated that a substance burning in a closed container eventually stopped burning because the air in the

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*Democritus is an alternate spelling.

Roald Hoffman

FIGURE 1.5
The Priestley Medal is the highest honor given by the American Chemical Society. It is named for Joseph Priestley, who was born in England on March 13, 1733. He performed many important scientific experiments, and among his discoveries was a gas, later identified as carbon dioxide, that could be dissolved in water to produce *seltzer*. Also, as a result of meeting Benjamin Franklin in London in 1766, Priestley became interested in electricity and was the first to observe that graphite was an electrical conductor. However, his greatest discovery occurred in 1774, when he isolated oxygen by heating mercuric oxide.

Because of his nonconformist political views, Priestley was forced to leave England. He died in the United States in 1804.

container became saturated with phlogiston. Oxygen gas, discovered by Joseph Priestley (1733–1804),* an English clergyman and scientist (Fig. 1.5), was found to support vigorous combustion and was thus supposed to be low in phlogiston. In fact, oxygen was originally called "dephlogisticated air."

1.4 ❯ Fundamental Chemical Laws

By the late eighteenth century, combustion had been studied extensively; the gases carbon dioxide, nitrogen, hydrogen, and oxygen had been discovered; and the list of elements continued to grow. However, it was Antoine Lavoisier (1743–1794), a French chemist (Fig. 1.6), who finally explained the true nature of combustion, thus clearing the way for the tremendous progress that was made near the end of the eighteenth century. Lavoisier, like Boyle, regarded measurement as the essential operation of chemistry. His experiments, in which he carefully weighed the reactants and products of various reactions, suggested that *mass is neither created nor destroyed*. Lavoisier's verification of this **law of conservation of mass** was the basis for the developments in chemistry in the nineteenth century.

Mass is neither created nor destroyed in a chemical reaction.

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*Oxygen gas was actually first observed by the Swedish chemist Karl W. Scheele (1742–1786), but because his results were published after Priestley's, the latter is commonly credited with the discovery of oxygen.

FIGURE 1.6
Antoine Lavoisier was born in Paris on August 26, 1743. Although Lavoisier's father wanted his son to follow him into the legal profession, young Lavoisier was fascinated by science. From the beginning of his scientific career, Lavoisier recognized the importance of accurate measurements. His careful weighings showed that mass is conserved in chemical reactions and that combustion involves reaction with oxygen. Also, he wrote the first modern chemistry textbook. It is not surprising that Lavoisier is often called the father of modern chemistry.

To help support his scientific work, Lavoisier invested in a private tax-collecting firm and married the daughter of one of the company executives. His connection to the tax collectors proved fatal, for radical French revolutionaries demanded his execution, which occurred on May 8, 1794.

Antoine Lavoisier and His Wife (Detail) by Jacques Louis David. The Metropolitan Museum of Art, Purchase, Mr. and Mrs. Charles Wrightsman. Gift, in honor of Everett Fahy, 1977

Oxygen is from the French *oxygène*, meaning "generator of acid," because it was initially considered to be an integral part of all acids.

Lavoisier's quantitative experiments showed that combustion involved oxygen (which Lavoisier named), not phlogiston. He discovered that life was supported by a process that also involved oxygen and was similar in many ways to combustion. In 1789 Lavoisier published the first modern chemistry textbook, *Elementary Treatise on Chemistry*, in which he presented a unified picture of the chemical knowledge assembled up to that time. Unfortunately, in the same year the text was published, the French Revolution broke out. Lavoisier, who had been associated with collecting taxes for the government, was executed on the guillotine as an enemy of the people in 1794.

After 1800, chemistry was dominated by scientists who, following Lavoisier's lead, performed careful weighing experiments to study the course of chemical reactions and to determine the composition of various chemical compounds. One of these chemists, a Frenchman, Joseph Proust (1754–1826), showed that *a given compound always contains exactly the same proportion of elements by mass*. For example, Proust found that the substance copper carbonate is always 5.3 parts copper to 4 parts oxygen to 1 part carbon (by mass). The principle of the constant composition of compounds, originally called "Proust's law," is now known as the **law of definite proportion**. A given compound always contains exactly the same proportion of elements by mass.

Manchester Literary and Philosophical Society

FIGURE 1.7

John Dalton (1766–1844), an Englishman, began teaching at a Quaker school when he was 12. His fascination with science included an intense interest in meteorology, which led to an interest in the gases of the air and their ultimate components, atoms. Dalton is best known for his atomic theory, in which he postulated that the fundamental differences among atoms are their masses. He was the first to prepare a table of relative atomic weights.

Dalton was a humble man with several apparent disabilities: He was not articulate and he was color-blind, a terrible problem for a chemist. Despite these disadvantages, he helped to revolutionize the science of chemistry.

Proust's discovery stimulated John Dalton (1766–1844), an English schoolteacher (Fig. 1.7), to think about atoms as the particles that might compose elements. Dalton reasoned that if elements were composed of tiny individual particles, a given compound should always contain the same combination of these atoms. This concept explained why the same relative masses of elements were always found in a given compound.

But Dalton discovered another principle that convinced him even more of the existence of atoms. He noted, for example, that carbon and oxygen form two different compounds that contain different relative amounts of carbon and oxygen, as shown by the following data:

	Mass of Oxygen That Combines with 1 g of Carbon
Compound I	1.33 g
Compound II	2.66 g

Dalton noted that compound II contains twice as much oxygen per gram of carbon as compound I, a fact that could easily be explained in terms of atoms. Compound I might be CO, and compound II might be CO_2.* This principle, which was found to apply to compounds of other elements as well, became known as the **law of multiple proportions**: When two elements form a series of compounds, the ratios of the masses of the second element that combine with 1 gram of the first element can always be reduced to small whole numbers.

To make sure the significance of this observation is clear, in Example 1.1 we will consider data for a series of compounds consisting of nitrogen and oxygen.

EXAMPLE 1.1 Illustrating the Law of Multiple Proportions

The following data were collected for several compounds of nitrogen and oxygen:

	Mass of Nitrogen That Combines with 1 g of Oxygen
Compound A	1.750 g
Compound B	0.8750 g
Compound C	0.4375 g

Show how these data illustrate the law of multiple proportions.

---

*Subscripts are used to show the numbers of atoms present. The number 1 is understood (not written). The symbols for the elements and the writing of chemical formulas will be illustrated further in Section 3.4.

Solution

For the law of multiple proportions to hold, the ratios of the masses of nitrogen combining with 1 gram of oxygen in each pair of compounds should be small whole numbers. We therefore compute the ratios as follows:

$$\frac{A}{B} = \frac{1.750}{0.8750} = \frac{2}{1}$$

$$\frac{B}{C} = \frac{0.8750}{0.4375} = \frac{2}{1}$$

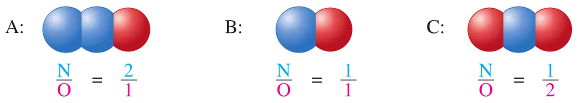
$$\frac{A}{C} = \frac{1.750}{0.4375} = \frac{4}{1}$$

These results support the law of multiple proportions.

See Exercises 1.29 and 1.30

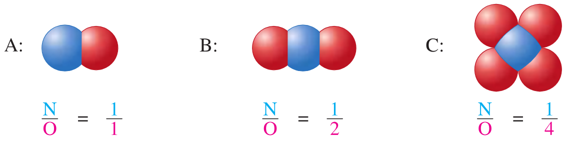
The significance of the data in Example 1.1 is that compound A contains twice as much nitrogen (N) per gram of oxygen (O) as does compound B and that compound B contains twice as much nitrogen per gram of oxygen as does compound C.

These data can be explained readily if the substances are composed of molecules made up of nitrogen atoms and oxygen atoms. For example, one set of possibilities for compounds A, B, and C is



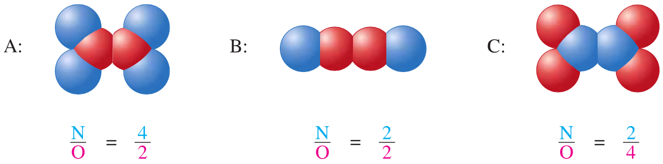
Now we can see that compound A contains two atoms of N for every atom of O, whereas compound B contains one atom of N per atom of O. That is, compound A contains twice as much nitrogen per given amount of oxygen as does compound B. Similarly, since compound B contains one N per O and compound C contains one N per *two* O's, the nitrogen content of compound C per given amount of oxygen is half that of compound B.

Another set of compounds that fits the data in Example 1.1 is



Verify for yourself that these compounds satisfy the requirements.

Still another set that works is



See if you can come up with still another set of compounds that satisfies the data in Example 1.1. How many more possibilities are there?

In fact, an infinite number of other possibilities exists. Dalton could not deduce absolute formulas from the available data on relative masses. However, the data on the composition of compounds in terms of the relative masses of the elements supported his hypothesis that each element consisted of a certain type of atom and that compounds were formed from specific combinations of atoms.

1.5 › Dalton's Atomic Theory

In 1808 Dalton published *A New System of Chemical Philosophy*, in which he presented his theory of atoms.

These statements are a modern paraphrase of Dalton's ideas.

Dalton's Atomic Theory

1. Each element is made up of tiny particles called atoms.
2. The atoms of a given element are identical; the atoms of different elements are different in some fundamental way or ways.
3. Chemical compounds are formed when atoms of different elements combine with each other. A given compound always has the same relative numbers and types of atoms.
4. Chemical reactions involve reorganization of the atoms—changes in the way they are bound together. The atoms themselves are not changed in a chemical reaction.

It is instructive to consider Dalton's reasoning on the relative masses of the atoms of the various elements. In Dalton's time water was known to be composed of the elements hydrogen and oxygen, with 8 grams of oxygen present for every 1 gram of hydrogen. If the formula for water were OH, an oxygen atom would have to have 8 times the mass of a hydrogen atom. However, if the formula for water were H_2O (two atoms of hydrogen for every oxygen atom), this would mean that each atom of oxygen is 16 times as massive as *each* atom of hydrogen (since the ratio of the mass of one oxygen to that of *two* hydrogens is 8 to 1). Because the formula for water was not then known, Dalton could not specify the relative masses of oxygen and hydrogen unambiguously. To solve the problem, Dalton made a fundamental assumption: He decided that nature would be as simple as possible. This assumption led him to conclude that the formula for water should be OH. He thus assigned hydrogen a mass of 1 and oxygen a mass of 8.

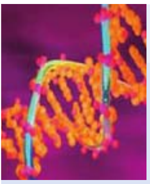
Using similar reasoning for other compounds, Dalton prepared the first table of **atomic masses** (sometimes called **atomic weights** by chemists, since mass is often determined by comparison to a standard mass—a process called *weighing*). Many of the masses were later proved to be wrong because of Dalton's incorrect assumptions about the formulas of certain compounds, but the construction of a table of masses was an important step forward.


The Granger Collection, New York

Joseph Louis Gay-Lussac, a French physicist and chemist, was remarkably versatile. Although he is now known primarily for his studies on the combining of volumes of gases, Gay-Lussac was instrumental in the studies of many of the other properties of gases. Some of Gay-Lussac's motivation to learn about gases arose from his passion for ballooning. In fact, he made ascents to heights of over 4 miles to collect air samples, setting altitude records that stood for about 50 years. Gay-Lussac also was the codiscoverer of boron and the developer of a process for manufacturing sulfuric acid. As chief assayer of the French mint, Gay-Lussac developed many techniques for chemical analysis and invented many types of glassware now used routinely in labs. Gay-Lussac spent his last 20 years as a lawmaker in the French government.

Although not recognized as such for many years, the keys to determining absolute formulas for compounds were provided in the experimental work of the French chemist Joseph Gay-Lussac (1778–1850) and by the hypothesis of an Italian chemist named Amadeo Avogadro (1776–1856). In 1809 Gay-Lussac performed experiments in which he measured (under the same conditions of temperature and pressure) the volumes of gases that reacted with each other. For example, Gay-Lussac found that 2 volumes of hydrogen react with 1 volume of oxygen to form 2 volumes of gaseous water and that 1 volume of hydrogen reacts with 1 volume of chlorine to form 2 volumes of hydrogen chloride. These results are represented schematically in Fig. 1.8.

In 1811 Avogadro interpreted these results by proposing that *at the same temperature and pressure, equal volumes of different gases contain the same number of particles*. This assumption (called **Avogadro's hypothesis**) makes sense if the distances between the particles in a gas are very great compared with the sizes of the particles. Under these conditions, the volume of a gas is determined by the number of molecules present, not by the size of the individual particles.



CHEMICAL CONNECTIONS

Berzelius, Selenium, and Silicon

Jöns Jakob Berzelius was probably the best experimental chemist of his generation and, given the crudeness of his laboratory equipment, maybe the best of all time. Unlike Lavoisier, who could afford to buy the best laboratory equipment available, Berzelius worked with minimal equipment in very plain surroundings. One of Berzelius's students described the Swedish chemist's workplace: "The laboratory consisted of two ordinary rooms with the very simplest arrangements; there were neither furnaces nor hoods, neither water system nor gas. Against the walls stood some closets with the chemicals, in the middle the mercury trough and the blast lamp table. Beside this was the sink consisting of a stone water holder with a stopcock and a pot standing under it. [Next door in the kitchen] stood a small heating furnace."

Comparison of Several of Berzelius's Atomic Masses with the Modern Values

Element	Atomic Mass Berzelius's Value	 Current Value
Chlorine	35.41	35.45
Copper	63.00	63.55
Hydrogen	1.00	1.01
Lead	207.12	207.2
Nitrogen	14.05	14.01
Oxygen	16.00	16.00
Potassium	39.19	39.10
Silver	108.12	107.87
Sulfur	32.18	32.07

In these simple facilities Berzelius performed more than 2000 experiments over a 10-year period to determine accurate atomic masses for the 50 elements then known. His success can be seen from the data in the table at left. These remarkably accurate values attest to his experimental skills and patience.

Besides his table of atomic masses, Berzelius made many other major contributions to chemistry. The most important of these was the invention of a simple set of symbols for the elements along with a system for writing the formulas of compounds to replace the awkward symbolic representations of the alchemists. Although some chemists, including Dalton, objected to the new system, it was gradually adopted and forms the basis of the system we use today.

In addition to these accomplishments, Berzelius discovered the elements cerium, thorium, selenium, and silicon. Of these elements, selenium and silicon are particularly important in today's world. Berzelius discovered selenium in 1817 in connection with his studies of sulfuric acid. For years selenium's toxicity has been known, but only recently have we become aware that it may have a positive effect on human health.

If Avogadro's hypothesis is correct, Gay-Lussac's result,

2 volumes of hydrogen react with 1 volume of oxygen $\longrightarrow$ 2 volumes of water vapor

can be expressed as follows:

2 molecules* of hydrogen react with 1 molecule of oxygen $\longrightarrow$ 2 molecules of water

There are seven elements that occur as diatomic molecules:
H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2

These observations can best be explained by assuming that gaseous hydrogen, oxygen, and chlorine are all composed of diatomic (two-atom) molecules: H_2, O_2, and Cl_2, respec-

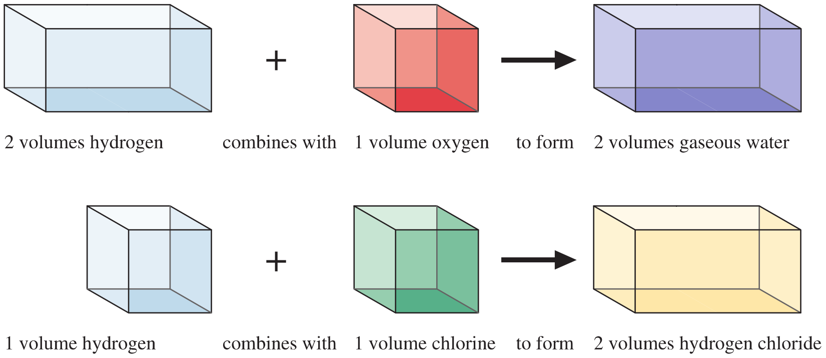


FIGURE 1.8
A representation of some of Gay-Lussac's experimental results on combining gas volumes.

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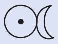
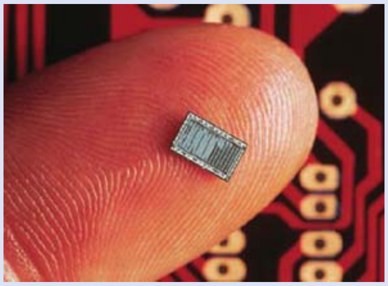
*A *molecule* is a collection of atoms (see Section 3.4).

The Alchemists' Symbols for Some Common Elements and Compounds

Substance	Alchemists' Symbol
Silver	
Lead	
Tin	
Platinum	
Sulfuric acid	
Alcohol	
Sea salt	

Studies have shown that trace amounts of selenium in the diet may protect people from heart disease and cancer. One study based on data from 27 countries showed an inverse relationship between the cancer death rate and the selenium content of soil in a particular region (low cancer death rate in areas with high selenium content). Another research paper reported an inverse relationship between the selenium content of the blood and the incidence of breast cancer in women. A study reported in 1998 used the toenail clippings of 33,737 men to show that selenium seems to protect against prostate cancer. Selenium is also found in the heart muscle and may play an important role in proper heart function. Because of these and other studies, selenium's reputation has improved, and many scientists are now studying its function in the human body.

Silicon is the second most abundant element in the earth's crust, exceeded only by oxygen. As we will see in Chapter 8, compounds involving silicon bonded to oxygen make up most of the earth's sand, rock, and soil. Berzelius prepared silicon in its pure form in 1824 by heating silicon tetrafluoride (SiF_4) with potassium metal. Today, silicon forms the basis for the modern microelectronics industry centered near San Francisco in a place that has come to be known as "Silicon Valley." The technology of the silicon chip (see figure) with its printed circuits has transformed computers from room-sized monsters with thousands of unreliable vacuum tubes to desktop and notebook-sized units with trouble-free "solid-state" circuitry.


Roger Du Buisson/Corbis

A silicon chip.

See E. J. Holmyard, *Alchemy* (New York: Penguin Books, 1968).

tively. Gay-Lussac's results can then be represented as shown in Fig. 1.9. (Note that this reasoning suggests that the formula for water is H_2O, not OH as Dalton believed.)

Unfortunately, Avogadro's interpretations were not accepted by most chemists, and a half-century of confusion followed, in which many different assumptions were made about formulas and atomic masses.

During the nineteenth century, painstaking measurements were made of the masses of various elements that combined to form compounds. From these experiments a list of relative atomic masses could be determined. One of the chemists involved in contributing to this list was a Swede named Jöns Jakob Berzelius (1779–1848), who discovered the elements cerium, selenium, silicon, and thorium and developed the modern symbols for the elements used in writing the formulas of compounds.

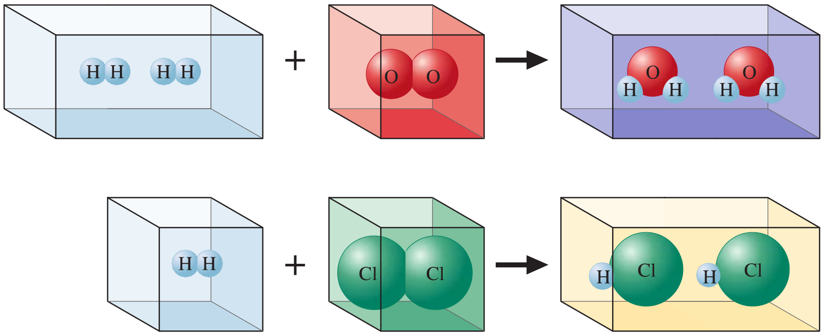
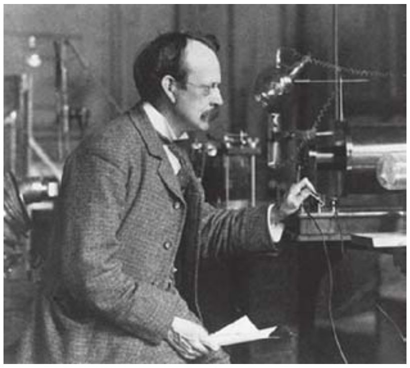


FIGURE 1.9
A representation of combining gases at the molecular level. The spheres represent atoms in the molecules.



The Cavendish Laboratory/University of Cambridge

FIGURE 1.10
J. J. Thomson (1856–1940) was an English physicist at Cambridge University. He received the Nobel Prize in physics in 1906.

1.6 › Early Experiments to Characterize the Atom

On the basis of the work of Dalton, Gay-Lussac, Avogadro, and others, chemistry was beginning to make sense. The concept of atoms was clearly a good idea. Inevitably, scientists began to wonder about the nature of the atom. What is an atom made of, and how do the atoms of the various elements differ?

The Electron

The first important experiments that led to an understanding of the composition of the atom were done by the English physicist J. J. Thomson (Fig. 1.10), who studied electrical discharges in partially evacuated tubes called **cathode-ray tubes** (Fig. 1.11) during the period from 1898 to 1903. Thomson found that when high voltage was applied to the tube, a "ray" he called a *cathode ray* (because it emanated from the negative electrode, or cathode) was produced. Because this ray was produced at the negative electrode and was repelled by the negative pole of an applied electric field (see Fig. 1.12), Thomson postulated that the ray was a stream of negatively charged particles, now called **electrons**. From experiments in which he measured the deflection of the beam of electrons in a magnetic field, Thomson determined the *charge-to-mass ratio* of an electron:

$$\frac{e}{m} = -1.76 \times 10^8 \text{ C/g}$$

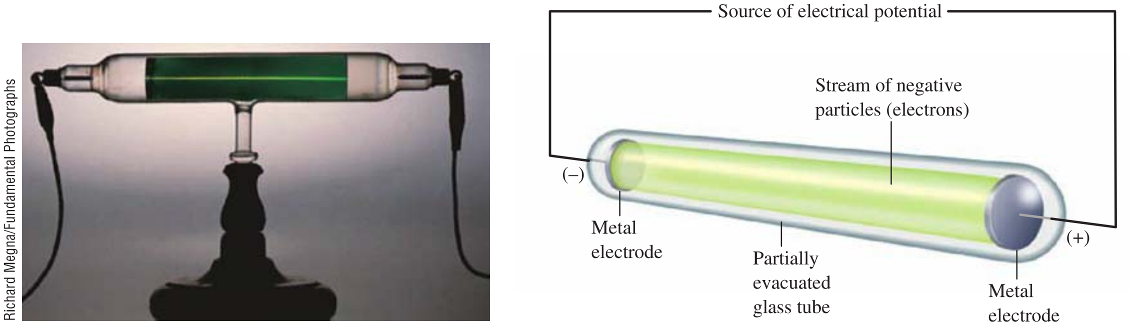
where e represents the charge on the electron in coulombs (C) and m represents the electron mass in grams.

StockFood/Getty Images

A classic English plum pudding, in which the raisins represent the distribution of electrons in the atom.

One of Thomson's primary goals in his cathode-ray tube experiments was to gain an understanding of the structure of the atom. He reasoned that since electrons could be produced from electrodes made of various types of metals, *all* atoms must contain electrons. Since atoms were known to be electrically neutral, Thomson further assumed that atoms also must contain some positive charge. Thomson postulated that an atom consisted of a diffuse cloud of positive charge with the negative electrons embedded randomly in it. This model, shown in Fig. 1.13, is often called the *plum pudding model* because the electrons are like raisins dispersed in a pudding (the positive charge cloud), as in plum pudding, a favorite English dessert.

In 1909 Robert Millikan (1868–1953), working at the University of Chicago, performed very clever experiments involving charged oil drops. These experiments allowed him to determine the magnitude of the electron charge (see Fig. 1.14). With this value



Richard Megna/Fundamental Photographs

FIGURE 1.11
A cathode-ray tube. The fast-moving electrons excite the gas in the tube, causing a glow between the electrodes. The green color in the photo is due to the response of the screen (coated with zinc sulfide) to the electron beam.

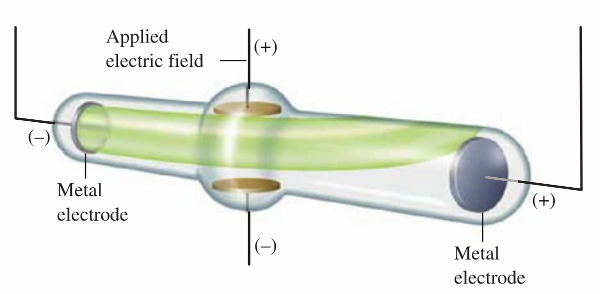


FIGURE 1.12
Deflection of cathode rays by an applied electric field.

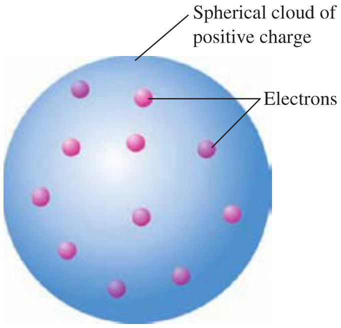


FIGURE 1.13
The plum pudding model of the atom.

and the charge-to-mass ratio determined by Thomson, Millikan was able to calculate the mass of the electron as 9.11×10^{-31} kilogram.

Radioactivity

In the late nineteenth century scientists discovered that certain elements produce high-energy radiation. For example, in 1896 the French scientist Henri Becquerel found accidentally that a piece of a mineral containing uranium could produce its image on a photographic plate in the absence of light. He attributed this phenomenon to a spontaneous emission of radiation by the uranium, which he called **radioactivity**. Studies in the early twentieth century demonstrated three types of radioactive emission: gamma (γ) rays, beta (β) particles, and alpha (α) particles. A γ ray is high-energy "light"; a β particle is a high-speed electron; and an α particle has a 2+ charge, that is, a charge twice that of the electron and with the opposite sign. The mass of an α particle is 7300 times that of the electron. More modes of radioactivity are now known, and we will discuss them in Chapter 19. Here we will consider only α particles because they were used in some crucial early experiments.

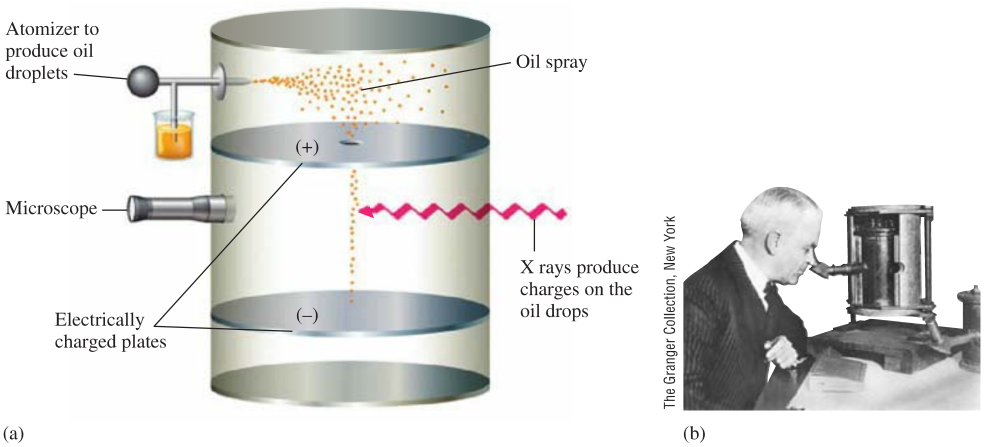
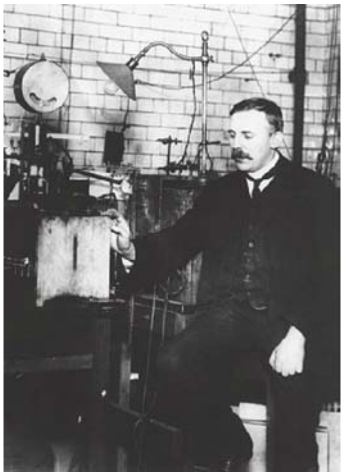


FIGURE 1.14
(a) A schematic representation of the apparatus Millikan used to determine the charge on the electron. The fall of charged oil droplets due to gravity can be halted by adjusting the voltage across the two plates. This voltage and the mass of the oil drop can then be used to calculate the charge on the oil drop. Millikan's experiments showed that the charge on an oil drop is always a whole-number multiple of the electron charge. (b) Robert Millikan using his apparatus.



Topham Picture Library/The Image Works

FIGURE 1.15
Ernest Rutherford (1871–1937) was born on a farm in New Zealand. In 1895 he placed second in a scholarship competition to attend Cambridge University but was awarded the scholarship when the winner decided to stay home and get married. As a scientist in England, Rutherford did much of the early work on characterizing radioactivity. He named the α and β particles and the γ ray and coined the term *half-life* to describe an important attribute of radioactive elements. His experiments on the behavior of α particles striking thin metal foils led him to postulate the nuclear atom. He also invented the name *proton* for the nucleus of the hydrogen atom. He received the Nobel Prize in chemistry in 1908.

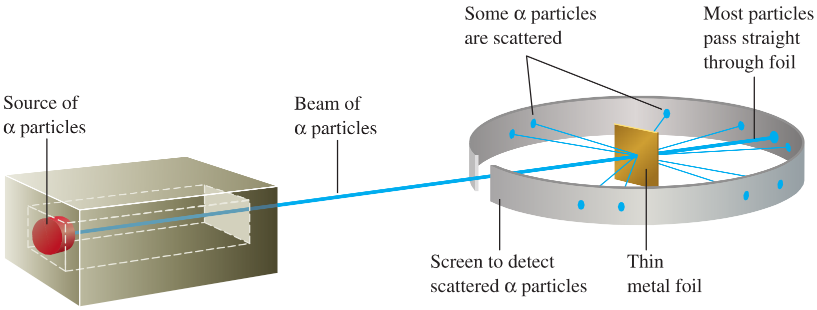


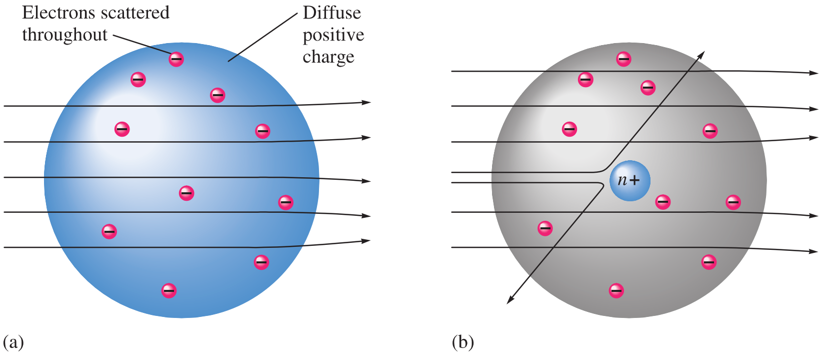
FIGURE 1.16
Rutherford's experiment on α-particle bombardment of metal foil.

The Nuclear Atom

In 1911 Ernest Rutherford (Fig. 1.15), who performed many of the pioneering experiments to explore radioactivity, carried out an experiment to test Thomson's plum pudding model. The experiment involved directing α particles at a thin sheet of metal foil, as illustrated in Fig. 1.16. Rutherford reasoned that if Thomson's model were accurate, the massive α particles should crash through the thin foil like cannonballs through gauze, as shown in Fig. 1.17(a). He expected the α particles to travel through the foil with, at the most, very minor deflections in their paths. The results of the experiment were very different from those Rutherford anticipated. Although most of the α particles passed straight through, many of the particles were deflected at large angles, as shown in Fig. 1.17(b), and some were reflected, never hitting the detector. This outcome was a great surprise to Rutherford. (He wrote that this result was comparable with shooting a howitzer at a piece of paper and having the shell reflected back.)

Rutherford knew from these results that the plum pudding model for the atom could not be correct. The large deflections of the α particles could be caused only by a center of concentrated positive charge that contains most of the atom's mass, as illustrated in Fig. 1.17(b). Most of the α particles pass directly through the foil because the atom is mostly open space. The deflected α particles are those that had a "close encounter" with the massive positive center of the atom, and the few reflected α particles are those that made a "direct hit" on the much more massive positive center.

In Rutherford's mind these results could be explained only in terms of a **nuclear atom**—an atom with a dense center of positive charge (the **nucleus**) with electrons moving around the nucleus at a distance that is large relative to the nuclear radius.



(a) (b)

FIGURE 1.17
(a) The expected results of the metal foil experiment if Thomson's model were correct. (b) Actual results.

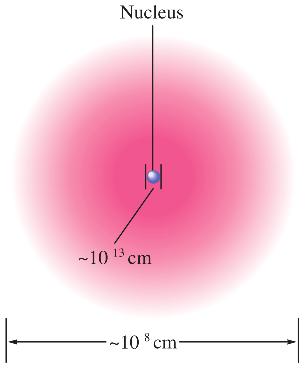


FIGURE 1.18
A nuclear atom viewed in cross section. Note that this drawing is not to scale.

1.7 › The Modern View of Atomic Structure: An Introduction

In the years since Thomson and Rutherford, a great deal has been learned about atomic structure. Because much of this material will be covered in detail in later chapters, only an introduction will be given here. The simplest view of the atom is that it consists of a tiny nucleus (with a diameter of about 10^{-13} cm) and electrons that move about the nucleus at an average distance of about 10^{-8} cm from it (see Fig. 1.18).

As we will see later, the chemistry of an atom mainly results from its electrons. For this reason, chemists can be satisfied with a relatively crude nuclear model. The nucleus is assumed to contain **protons**, which have a positive charge equal in magnitude to the electron's negative charge, and **neutrons**, which have virtually the same mass as a proton but no charge. The masses and charges of the electron, proton, and neutron are shown in Table 1.1.

Two striking things about the nucleus are its small size compared with the overall size of the atom and its extremely high density. The tiny nucleus accounts for almost all the atom's mass. Its great density is dramatically demonstrated by the fact that a piece of nuclear material about the size of a pea would have a mass of 250 million tons!

An important question to consider at this point is, "*If all atoms are composed of these same components, why do different atoms have different chemical properties*?" The answer to this question lies in the number and the arrangement of the electrons. The electrons constitute most of the atomic volume and thus are the parts that "intermingle" when atoms combine to form molecules. Therefore, the number of electrons possessed by a given atom greatly affects its ability to interact with other atoms. As a result, the atoms of different elements, which have different numbers of protons and electrons, show different chemical behavior.

The *chemistry* of an atom arises from its electrons.

A sodium atom has 11 protons in its nucleus. Since atoms have no net charge, the number of electrons must equal the number of protons. Therefore, a sodium atom has 11 electrons moving around its nucleus. It is *always* true that a sodium atom has 11 protons and 11 electrons. However, each sodium atom also has neutrons in its nucleus, and different types of sodium atoms exist that have different numbers of neutrons. For example, consider the sodium atoms represented in Fig. 1.19. These two atoms are **isotopes**, or *atoms with the same number of protons but different numbers of neutrons*. Note that the symbol for one particular type of sodium atom is written

Mass number → $^{23}_{11}\text{Na}$ ← Element symbol
Atomic number → $^{23}_{11}\text{Na}$

Mass number → $^{A}_{Z}X$ ← Element symbol
Atomic number → $^{A}_{Z}X$

TABLE 1.1 › The Mass and Charge of the Electron, Proton, and Neutron

Particle	Mass	Charge*
Electron	9.109×10^{-31} kg	1−
Proton	1.673×10^{-27} kg	1+
Neutron	1.675×10^{-27} kg	None

*The magnitude of the charge of the electron and the proton is 1.60×10^{-19} C.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If the atomic nucleus were the size of this ball bearing, a typical atom would be the size of this stadium.

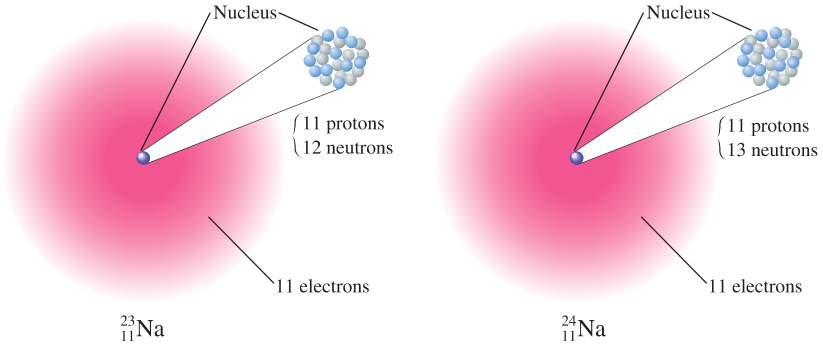

FIGURE 1.19
Two isotopes of sodium. Both have 11 protons and 11 electrons, but they differ in the number of neutrons in their nuclei.

where the **atomic number** Z (number of protons) is written as a subscript, and the **mass number** A (the total number of protons and neutrons) is written as a superscript. (The particular atom represented here is called "sodium twenty-three." It has 11 electrons, 11 protons, and 12 neutrons.) Because the chemistry of an atom is due to its electrons, isotopes show almost identical chemical properties. In nature most elements contain mixtures of isotopes.

INTERACTIVE EXAMPLE 1.2

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Writing the Symbols for Atoms

Write the symbol for the atom that has an atomic number of 9 and a mass number of 19. How many electrons and how many neutrons does this atom have?

Solution

The atomic number 9 means the atom has 9 protons. This element is called *fluorine*, symbolized by F. The atom is represented as

$$^{19}_{9}\text{F}$$

and is called *fluorine nineteen*. Since the atom has 9 protons, it also must have 9 electrons to achieve electrical neutrality. The mass number gives the total number of protons and neutrons, which means that this atom has 10 neutrons.

See Exercises 1.43 through 1.46

1.8 › Introduction to Energy

Although energy is a familiar concept, it is difficult to define precisely. For our purposes we will define **energy** as the ability to do work or produce heat. **Work** is defined as force acting over a distance. Heat is best defined as energy that flows from one object to another because of a temperature difference between the two objects.

One of the most important characteristics of energy is that it can be converted from one form to another, and in such a process no energy is created or destroyed. This principle is summarized by the Law of Conservation of Energy, which states that energy can be converted from one form to another but the total quantity of energy is not changed; that is, the energy content of the universe is constant.

Energy can be divided into two categories: kinetic energy (KE) and potential energy (PE). **Kinetic energy** is the energy of motion. A baseball thrown by a pitcher at 90 miles per hour has more kinetic energy than one thrown at 80 miles per hour. **Potential energy** can be defined as stored energy due to position. For example, the water stored behind a

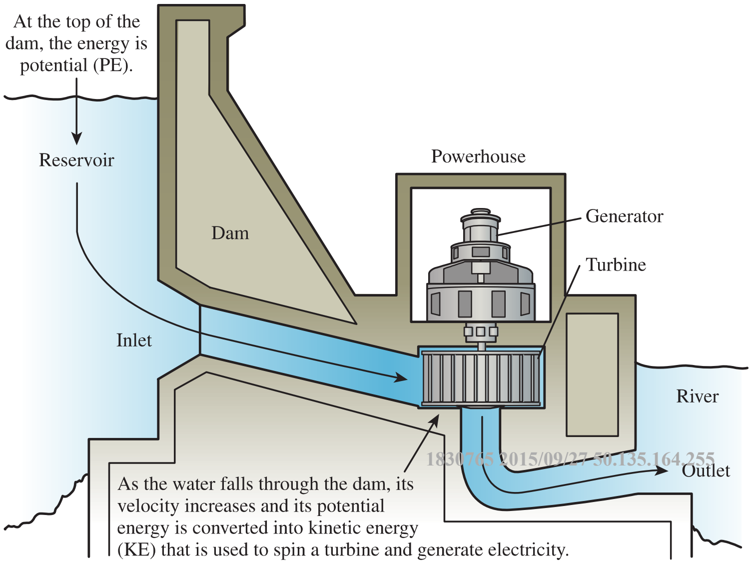

FIGURE 1.20
Using a dam to convert potential energy to kinetic energy.

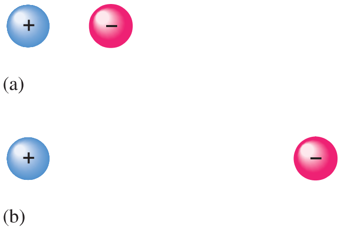

FIGURE 1.21
Two positrons of opposite charges.

dam has potential energy that can be converted to electrical energy, as shown in Fig. 1.20. In this case the water behind the dam is at a higher level than the outlet water in the river and thus has a higher PE. This PE can be changed to KE as it runs through the generator turbine, which produces electricity.

Another type of PE important in chemical processes is that due to the attraction of opposite charges. For example, consider two relative positions of a positively charged proton and a negatively charged electron, shown in Fig. 1.21. The potential energies of the two cases are different. Does the situation shown in Fig. 1.21(a) have a higher or lower PE than that shown in Fig. 1.21(b)? We can answer this question by devising an imaginary tiny machine, as illustrated in Fig. 1.22. Work is done by this machine as the negative charge moves toward the positive charge. Which situation [Fig. 1.21(a) or Fig. 1.21(b)] would produce the most work (the greatest lifting of the weight)? We can see that the situation in Fig. 1.21(b), with the charges furthest apart, can do the most work because the negative charge moves the greatest distance and thus lifts the weight the greatest distance. Thus, because Fig. 1.21(b) has more stored energy than Fig. 1.21(a), Fig. 1.21(b) has a higher PE. In general, the farther apart two opposite charges are, the greater the stored energy and the greater the PE.

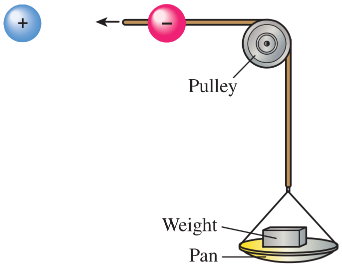

FIGURE 1.22
Work is done as the charges move together and the weight is lifted.

The significance of charged-based PE in chemistry has to do with the energies of electrons in atoms and the energies of chemical reactions, as we will see in Chapters 2 and 6, respectively.

The fundamental unit of energy in science is the joule (pronounced jewel). A **joule** is defined as a kilogram meter squared per second squared:

$$1 \text{ joule} = \frac{\text{kg} \cdot \text{m}^2}{\text{s}^2}.$$

Often the kilojoule (kJ) is used to describe energies.

$$1 \text{ kJ} = 10^3 \text{ J}$$

We will cover energy in much more detail in Chapter 6.

1.9 › The Mole: An Introduction

Have you ever wondered how many water molecules are in a glass of water (H_2O)? How can we find out? The molecules are so tiny we can't see them with the naked eye, so we can't count them directly. As we saw in section 1.1, recent technology, such as the scanning tunneling microscope, allows us to "see" individual atoms and molecules. However, this technology is very recent. How did we count atoms and molecules before this? The answer is we count atoms and molecules by weighing samples containing them.

How do we count by weighing? Consider a sample of jelly beans. What do we need to know about these jelly beans to count them by weighing samples of them? The answer is we need to know the average mass of a jelly bean. For example, if the average mass is 5.0 g, then a 500 g sample would contain 100 jelly beans.

We count atoms and molecules by weighing in the same way we count jelly beans. If we know the average mass of a water molecule, we can determine the number of water molecules in a glass of water by weighing the sample of water in the glass.

We won't detail here how we determine the mass of a water molecule (we will do that in Chapter 3). At this point we ask you to trust us that we have determined the average masses of all the atoms and molecules and can use these masses to count atoms and molecules by weighing samples of them.

When we count the water molecules in a glass of water (by weighing the water) we find that the number present is unimaginably large. The number of molecules in a glass of water is much larger than the age of the earth in seconds (4.32×10^{17} seconds) and larger than the number of milliliters of water in the earth's oceans (1.3×10^{24} mL). Therefore, because of the huge numbers of atoms and molecules in normal-sized samples of matter, we need to invent a unit for describing the number of atoms and molecules present. This unit has to be very large to be convenient. The unit of one dozen (12) works fine for eggs, but it wouldn't help much for describing numbers of atoms or molecules. The unit we have chosen is called the **mole**, which for our purposes we will define as 6.022×10^{23}. Thus, a glass of water that contains 9.0×10^{24} water molecules contains

$$9.0 \times 10^{24}\ \text{H}_2\text{O molecules} \times \frac{1\ \text{mol}}{6.022 \times 10^{23}\ \text{molecules}} = 15\ \text{mol H}_2\text{O molecules.}$$

Note that 15 moles is much more convenient than 9.0×10^{24} water molecules.

We will discuss the mole and its use in chemistry in much more detail in Chapter 3. For the present we simply need to know that the mole is a unit for counting atoms and molecules. For example, when energy terms are given, they are usually given per mole. Examples are the bond energies we will consider in Chapter 4. The bond energy for the hydrogen molecule (H_2) is 432 kJ/mol (where mol is the abbreviation for mole). This means that 432 kJ of energy is required to break one mole (6.022×10^{23}) of H—H bonds. Just remember that when you encounter "per mole" on a unit, it means that 6.022×10^{23} events are represented by the quantity in question.

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› Key Terms

Section 1.2

scientific method
measurement
hypothesis
theory
model
natural law
law of conservation of mass

Section 1.4

law of conservation of mass
law of definite proportion
law of multiple proportions

Section 1.5

atomic masses
atomic weights
Avogadro's hypothesis

Section 1.6

cathode-ray tubes
electrons
radioactivity
nuclear atom
nucleus

Section 1.7

protons
neutrons
isotopes
atomic number
mass number

Section 1.8

energy
work
kinetic energy
potential energy
joule

Section 1.9

mole

FOR REVIEW

Scientific method

- Make observations
- Formulate hypotheses
- Perform experiments

Models (theories) are explanations of why nature behaves in a particular way.

- They are subject to modification over time and sometimes fail.

Quantitative observations are called measurements.

- Consist of a number and a unit
- Involve some uncertainty
- Uncertainty is indicated by using significant figures
    - Rules to determine significant figures
    - Calculations using significant figures
- Preferred system is SI

Fundamental laws

- Conservation of mass
- Definite proportion
- Multiple proportions

Dalton's atomic theory

- All elements are composed of atoms.
- All atoms of a given element are identical.
- Chemical compounds are formed when atoms combine.
- Atoms are not changed in chemical reactions, but the way they are bound together changes.

Early atomic experiments and models

- Thomson model
- Millikan experiment
- Rutherford experiment
- Nuclear model

Atomic structure

- Small dense nucleus contains protons and neutrons.
    - Protons—positive charge
    - Neutrons—no charge
- Electrons reside outside the nucleus in the relatively large remaining atomic volume.
    - Electrons—negative charge, small mass (1/1840 of proton)
- Isotopes have the same atomic number but different mass numbers.

Energy

- The ability to do work or release heat.
    - Kinetic—energy of motion
    - Potential—stored energy of position
    - Joule—SI unit

Mole

- 6.022×10^{23} units

REVIEW QUESTIONS

1. Define and explain the differences between the following terms.
    - a. law and theory
    - b. theory and experiment
    - c. qualitative and quantitative
    - d. hypothesis and theory
2. Is the scientific method suitable for solving problems only in the sciences? Explain.
3. Use Dalton's atomic theory to account for each of the following.
    - a. the law of conservation of mass
    - b. the law of definite proportion
    - c. the law of multiple proportions
4. What evidence led to the conclusion that cathode rays had a negative charge?
5. What discoveries were made by J. J. Thomson, Henri Becquerel, and Lord Rutherford? How did Dalton's model of the atom have to be modified to account for these discoveries?
6. Consider Ernest Rutherford's α-particle bombardment experiment illustrated in Fig. 1.16. How did the results of this experiment lead Rutherford away from the plum pudding model of the atom to propose the nuclear model of the atom?
7. Do the proton and the neutron have exactly the same mass? How do the masses of the proton and neutron compare to the mass of the electron? Which particles make the greatest contribution to the mass of an atom? Which particles make the greatest contribution to the chemical properties of an atom?
8. What is the distinction between atomic number and mass number? Between mass number and atomic mass?
9. What is the Law of Conservation of Energy? Differentiate between kinetic energy and potential energy.
10. Explain the concept of counting by weighing. What is a mole and why do we use the mole unit?

› Active Learning Questions

These questions are designed to be used by groups of students in class. V denotes Visual Exercises and Questions.

1. Paracelsus, a sixteenth-century alchemist and healer, adopted as his slogan: "The patients are your textbook, the sickbed is your study." Is this view consistent with using the scientific method?

2. What is wrong with the following statement? "The results of the experiment do not agree with the theory. Something must be wrong with the experiment."

3. Which of the following is true about an individual atom? Explain.
- **a.** An individual atom should be considered to be a solid.
- **b.** An individual atom should be considered to be a liquid.
- **c.** An individual atom should be considered to be a gas.
- **d.** The state of the atom depends on which element it is.
- **e.** An individual atom cannot be considered to be a solid, liquid, or gas.

Justify your choice, and for choices you did not pick, explain what is wrong with them.

4. These questions concern the work of J. J. Thomson.
- **a.** From Thomson's work, which particles do you think he would feel are most important for the formation of compounds (chemical changes), and why?
- **b.** Of the remaining two subatomic particles, which do you place second in importance for forming compounds, and why?
- **c.** Propose three models that explain Thomson's findings and evaluate them. To be complete you should include Thomson's findings.

5. Heat is applied to an ice cube in a closed container until only steam is present. Draw a representation of this process, assuming you can see it at an extremely high level of magnification. What happens to the size of the molecules? What happens to the total mass of the sample?

V **6.** You have a chemical in a sealed glass container filled with air. The setup is sitting on a balance as shown below. The chemical is ignited by means of a magnifying glass focusing sunlight on the reactant. After the chemical has completely burned, which of the following is true? Explain your answer.

- **a.** The balance will read less than 250.0 g.
- **b.** The balance will read 250.0 g.

c. The balance will read greater than 250.0 g.
d. The scale's reading cannot be determined without knowing the identity of the chemical.

7. You may have noticed that when water boils, you can see bubbles that rise to the surface of the water. Which of the following is inside these bubbles? Explain.
a. air
b. hydrogen and oxygen gas
c. oxygen gas
d. water vapor
e. carbon dioxide gas

8. One of the best indications of a useful theory is that it raises more questions for further experimentation than it originally answered. Does this apply to Dalton's atomic theory? Give examples.

9. Dalton assumed that all atoms of the same element were identical in all their properties. Explain why this assumption is not valid.

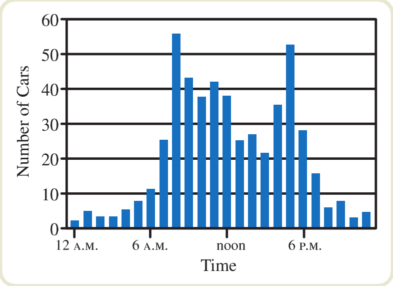
10. Which (if any) of the following can be determined by knowing the number of protons in a neutral element? Explain your answer.
a. the number of neutrons in the neutral element
b. the number of electrons in the neutral element
c. the name of the element

A blue question or exercise number indicates that the answer to that question or exercise appears at the back of this book and a solution appears in the *Student Solutions Manual.*

Questions

11. The difference between a *law* and a *theory* is the difference between *what* and *why*. Explain.

V 12. As part of a science project, you study traffic patterns in your city at an intersection in the middle of downtown. You set up a device that counts the cars passing through this intersection for a 24-hour period during a weekday. The graph of hourly traffic looks like this.



a. At what time(s) does the highest number of cars pass through the intersection?
b. At what time(s) does the lowest number of cars pass through the intersection?
c. Briefly describe the trend in numbers of cars over the course of the day.
d. Provide a hypothesis explaining the trend in numbers of cars over the course of the day.
e. Provide a possible experiment that could test your hypothesis.

13. Explain the fundamental steps of the scientific method. The scientific method is a dynamic process. What does this mean?

14. When hydrogen is burned in oxygen to form water, the composition of water formed does not depend on the amount of oxygen reacted. Interpret this in terms of the law of definite proportion.

15. Explain the law of conservation of mass, the law of definite proportion, and the law of multiple proportions.

16. Section 1.5 describes the postulates of Dalton's atomic theory. With some modifications, these postulates hold up very well regarding how we view elements, compounds, and chemical reactions today. Answer the following questions concerning Dalton's atomic theory and the modifications made today.
a. The atom can be broken down into smaller parts. What are the smaller parts?
b. How are atoms of hydrogen identical to each other and how can they be different from each other?
c. How are atoms of hydrogen different from atoms of helium? How can H atoms be similar to He atoms?
d. How is water different from hydrogen peroxide (H_2O_2) even though both compounds are composed of only hydrogen and oxygen?
e. What happens in a chemical reaction and why is mass conserved in a chemical reaction?

17. The contributions of J. J. Thomson and Ernest Rutherford led the way to today's understanding of the structure of the atom. What were their contributions?

18. What is the modern view of the structure of the atom?

19. The number of protons in an atom determines the identity of the atom. What do the number and arrangement of the electrons in an atom determine? What does the number of neutrons in an atom determine?

20. If the volume of a proton is similar to the volume of an electron, how will the densities of these two particles compare to each other?

21. For lighter, stable isotopes, the ratio of the mass number to the atomic number is close to a certain value. What is the value? What happens to the value of the mass number to atomic number ratio as stable isotopes become heavier?

22. What refinements had to be made in Dalton's atomic theory to account for Gay-Lussac's results on the combining volumes of gases?

23. Why is the energy of the universe constant?

24. What is the importance of the mole concept?

Exercises

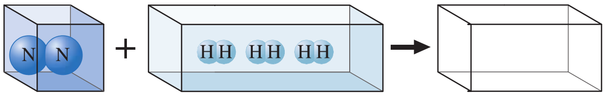
OWL Interactive versions of these problems may be assigned in OWL.

In this section, similar exercises are paired.

Development of the Atomic Theory

25. When mixtures of gaseous H_2 and gaseous Cl_2 react, a product forms that has the same properties regardless of the relative amounts of H_2 and Cl_2 used.
a. How is this result interpreted in terms of the law of definite proportion?
b. When a volume of H_2 reacts with an equal volume of Cl_2 at the same temperature and pressure, what volume of product having the formula HCl is formed?

V **26.** Observations of the reaction between nitrogen gas and hydrogen gas show us that 1 volume of nitrogen reacts with 3 volumes of hydrogen to make 2 volumes of gaseous product, as shown below:



Determine the formula of the product and justify your answer.

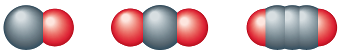
27. A sample of chloroform is found to contain 12.0 g of carbon, 106.4 g of chlorine, and 1.01 g of hydrogen. If a second sample of chloroform is found to contain 30.0 g of carbon, what is the total mass of chloroform in the second sample?

28. A sample of H_2SO_4 contains 2.02 g of hydrogen, 32.07 g of sulfur, and 64.00 g of oxygen. How many grams of sulfur and grams of oxygen are present in a second sample of H_2SO_4 containing 7.27 g of hydrogen?

29. Hydrazine, ammonia, and hydrogen azide all contain only nitrogen and hydrogen. The mass of hydrogen that combines with 1.00 g of nitrogen for each compound is 1.44×10^{-1} g, 2.16×10^{-1} g, and 2.40×10^{-2} g, respectively. Show how these data illustrate the law of multiple proportions.

30. Consider 100.0-g samples of two different compounds consisting only of carbon and oxygen. One compound contains 27.2 g of carbon and the other has 42.9 g of carbon. How can these data support the law of multiple proportions if 42.9 is not a multiple of 27.2? Show that these data support the law of multiple proportions.

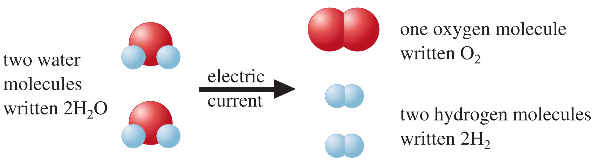
V **31.** The three most stable oxides of carbon are carbon monoxide (CO), carbon dioxide (CO_2), and carbon suboxide (C_3O_2). The molecules can be represented as



Explain how these molecules illustrate the law of multiple proportions.

32. Two elements, R and Q, combine to form two binary compounds. In the first compound, 14.0 g of R combines with 3.00 g of Q. In the second compound, 7.00 g of R combines with 4.50 g of Q. Show that these data are in accord with the law of multiple proportions. If the formula of the second compound is RQ, what is the formula of the first compound?

V **33.** In Section 1.1 of the text, the concept of a chemical reaction was introduced with the example of the decomposition of water, represented as follows:



Use ideas from Dalton's atomic theory to explain how the above representation illustrates the law of conservation of mass.

34. In a combustion reaction, 46.0 g of ethanol reacts with 96.0 g of oxygen to produce water and carbon dioxide. If 54.0 g of water is produced, what mass of carbon dioxide is produced?

35. Early tables of atomic weights (masses) were generated by measuring the mass of a substance that reacts with 1.00 g of oxygen. Given the following data and taking the atomic mass of hydrogen as 1.00, generate a table of relative atomic masses for oxygen, sodium, and magnesium.

Element	Mass That Combines with 1.00 g Oxygen	Assumed Formula
Hydrogen	0.126 g	HO
Sodium	2.875 g	NaO
Magnesium	1.500 g	MgO

36. Indium oxide contains 4.784 g of indium for every 1.000 g of oxygen. In 1869, when Mendeleev first presented his version of the periodic table, he proposed the formula In_2O_3 for indium oxide. Before that time it was thought that the formula was InO. What values for the atomic mass of indium are obtained using these two formulas? Assume that oxygen has an atomic mass of 16.00.

The Nature of the Atom

37. From the information in this chapter on the mass of the proton, the mass of the electron, and the sizes of the nucleus and the atom, calculate the densities of a hydrogen nucleus and a hydrogen atom.

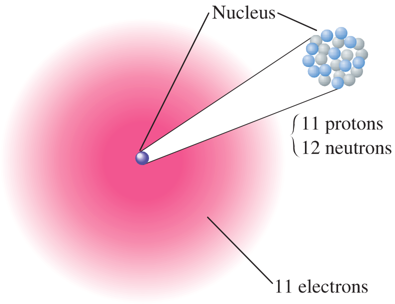
38. If you wanted to make an accurate scale model of the hydrogen atom and decided that the nucleus would have a diameter of 1 mm, what would be the diameter of the entire model?

39. In an experiment it was found that the total charge on an oil drop was 5.93×10^{-18} C. How many negative charges does the drop contain?

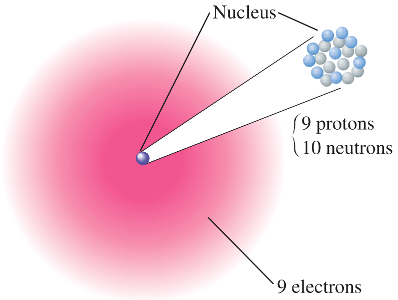
40. A chemist in a galaxy far, far away performed the Millikan oil drop experiment and got the following results for the charges on various drops. Use these data to calculate the charge of the electron in zirkombs.

2.56×10^{-12} zirkombs	7.68×10^{-12} zirkombs
3.84×10^{-12} zirkombs	6.40×10^{-13} zirkombs

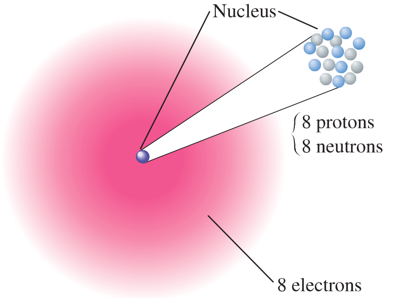
V **41.** Write the symbol of each atom using the ${}^{A}_{Z}X$ format.



a.



b.



c.

42. For carbon-14 and carbon-12, how many protons and neutrons are in each nucleus? Assuming neutral atoms, how many electrons are present in an atom of carbon-14 and in an atom of carbon-12?

43. How many protons and neutrons are in the nucleus of each of the following atoms? In a neutral atom of each element, how many electrons are present?

a. ^{79}Br
b. ^{81}Br
c. ^{239}Pu
d. ^{133}Cs
e. ^{3}H
f. ^{56}Fe

44. What number of protons and neutrons is contained in the nucleus of each of the following atoms? Assuming each atom is uncharged, what number of electrons is present?

a. $^{235}_{92}U$
b. $^{13}_{6}C$
c. $^{57}_{26}Fe$
d. $^{208}_{82}Pb$
e. $^{86}_{37}Rb$
f. $^{41}_{20}Ca$

45. Write the atomic symbol ($^{A}_{Z}X$) for each of the following isotopes.

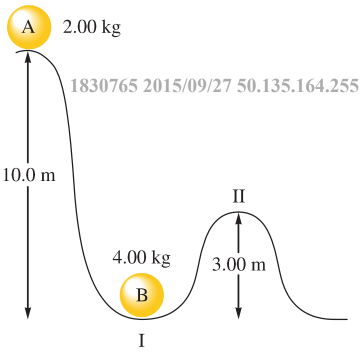
a. $Z = 8$, number of neutrons = 9
b. the isotope of chlorine in which $A = 37$
c. $Z = 27$, $A = 60$
d. number of protons = 26, number of neutrons = 31
e. the isotope of I with a mass number of 131
f. $Z = 3$, number of neutrons = 4

46. Write the atomic symbol ($^{A}_{Z}X$) for each of the isotopes described below.

a. number of protons = 27, number of neutrons = 31
b. the isotope of boron with mass number 10
c. $Z = 12$, $A = 23$
d. atomic number 53, number of neutrons = 79
e. $Z = 9$, number of neutrons = 10
f. number of protons = 29, mass number 65

Additional Exercises

V **47.** Consider the accompanying diagram. Ball A is allowed to fall and strike ball B. Assume that all of ball A's energy is transferred to ball B, at point I, and that there is no loss of energy to other sources. What are the kinetic energy and the potential energy of ball B at point II? The potential energy is given by $PE = mgz$, where m is the mass in kilograms, g is the gravitational constant ($9.81\ m/s^2$), and z is the distance in meters.



48. If you had a mole of U.S. dollar bills and equally distributed the money to all of the people of the world, how rich would every person be? Assume a world population of 6 billion.

49. Chlorine has two natural isotopes: $^{37}_{17}Cl$ and $^{35}_{17}Cl$. Hydrogen reacts with chlorine to form the compound HCl. Would a given amount of hydrogen react with different masses of the two chlorine isotopes? Does this conflict with the law of definite proportion? Why or why not?

50. Which of the following statements is/are *true*? For the false statements, correct them.

a. All particles in the nucleus of an atom are charged.
b. The atom is best described as a uniform sphere of matter in which electrons are embedded.
c. The mass of the nucleus is only a very small fraction of the mass of the entire atom.
d. The volume of the nucleus is only a very small fraction of the total volume of the atom.
e. The number of neutrons in a neutral atom must equal the number of electrons.

51. Identify each of the following elements. Give the number of protons and neutrons in each nucleus.

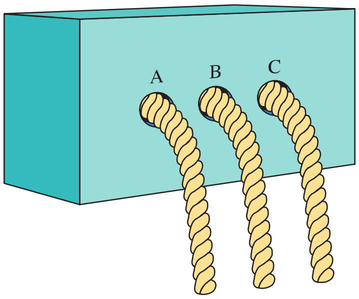
a. $^{31}_{15}X$
b. $^{127}_{53}X$
c. $^{39}_{19}X$
d. $^{173}_{70}X$

52. The early alchemists used to do an experiment in which water was boiled for several days in a sealed glass container. Eventually, some solid residue would appear in the bottom of the flask, which was interpreted to mean that some of the water in the flask had been converted into "earth." When Lavoisier repeated this experiment, he found that the water weighed the same before and after heating and the mass of the flask plus the solid residue equaled the original mass of the flask. Were the alchemists correct? Explain what happened. (This experiment is described in the article by A. F. Scott in *Scientific American,* January 1984.)

53. In a reaction, 34.0 g of chromium(III) oxide reacts with 12.1 g of aluminum to produce chromium and aluminum oxide. If 23.3 g of chromium is produced, what mass of aluminum oxide is produced?

› Challenge Problems

V **54.** Confronted with the box shown in the diagram, you wish to discover something about its internal workings. You have no tools and cannot open the box. You pull on rope B, and it moves rather freely. When you pull on rope A, rope C appears to be pulled slightly into the box. When you pull on rope C, rope A almost disappears into the box.

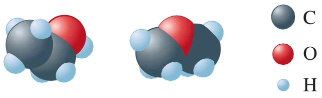


a. Based on these observations, construct a model for the interior mechanism of the box.

b. What further experiments could you do to refine your model?

V 55. Each of the following statements is true, but Dalton might have had trouble explaining some of them with his atomic theory. Give explanations for the following statements.

a. The space-filling models for ethyl alcohol and dimethyl ether are shown below.



These two compounds have the same composition by mass (52% carbon, 13% hydrogen, and 35% oxygen), yet the two have different melting points, boiling points, and solubilities in water.

b. Burning wood leaves an ash that is only a small fraction of the mass of the original wood.

c. Atoms can be broken down into smaller particles.

d. One sample of lithium hydride is 87.4% lithium by mass, while another sample of lithium hydride is 74.9% lithium by mass. However, the two samples have the same chemical properties.

56. Reaction of 2.0 L of hydrogen gas with 1.0 L of oxygen gas yields 2.0 L of water vapor. All gases are at the same temperature and pressure. Show how these data support the idea that oxygen gas is a diatomic molecule. Must we consider hydrogen to be a diatomic molecule to explain these results?

57. A combustion reaction involves the reaction of a substance with oxygen gas. The complete combustion of any hydrocarbon (binary compound of carbon and hydrogen) produces carbon dioxide and water as the only products. Octane is a hydrocarbon that is found in gasoline. Complete combustion of octane produces 8 liters of carbon dioxide for every 9 liters of water vapor (both measured at the same temperature and pressure). What is the ratio of carbon atoms to hydrogen atoms in a molecule of octane?

58. A chemistry instructor makes the following claim: "Consider that if the nucleus were the size of a grape, the electrons would be about 1 *mile* away on average." Is this claim reasonably accurate? Provide mathematical support.

59. You have two distinct gaseous compounds made from element X and element Y. The mass percents are as follows:

Compound I: 30.43% X, 69.57% Y
Compound II: 63.64% X, 36.36% Y

In their natural standard states, element X and element Y exist as gases. (Monatomic? Diatomic? Triatomic? That is for you to determine.) When you react "gas X" with "gas Y" to make the products, you get the following data (all at the same pressure and temperature):

1 volume "gas X" + 2 volumes "gas Y" ⟶ 2 volumes compound I

2 volumes "gas X" + 1 volume "gas Y" ⟶ 2 volumes compound II

Assume the simplest possible formulas for reactants and products in the chemical equations above. Then, determine the relative atomic masses of element X and element Y.

› Marathon Problem

This problem is designed to incorporate several concepts and techniques into one situation.

60. You have gone back in time and are working with Dalton on a table of relative masses. Following are his data.

0.602 g gas A reacts with 0.295 g gas B
0.172 g gas B reacts with 0.401 g gas C
0.320 g gas A reacts with 0.374 g gas C

a. Assuming simplest formulas (AB, BC, and AC), construct a table of relative masses for Dalton.

b. Knowing some history of chemistry, you tell Dalton that if he determines the volumes of the gases reacted at constant temperature and pressure, he need not assume simplest formulas. You collect the following data:

6 volumes gas A + 1 volume gas B → 4 volumes product
1 volume gas B + 4 volumes gas C → 4 volumes product
3 volumes gas A + 2 volumes gas C → 6 volumes product

Write the simplest balanced equations, and find the actual relative masses of the elements. Explain your reasoning.