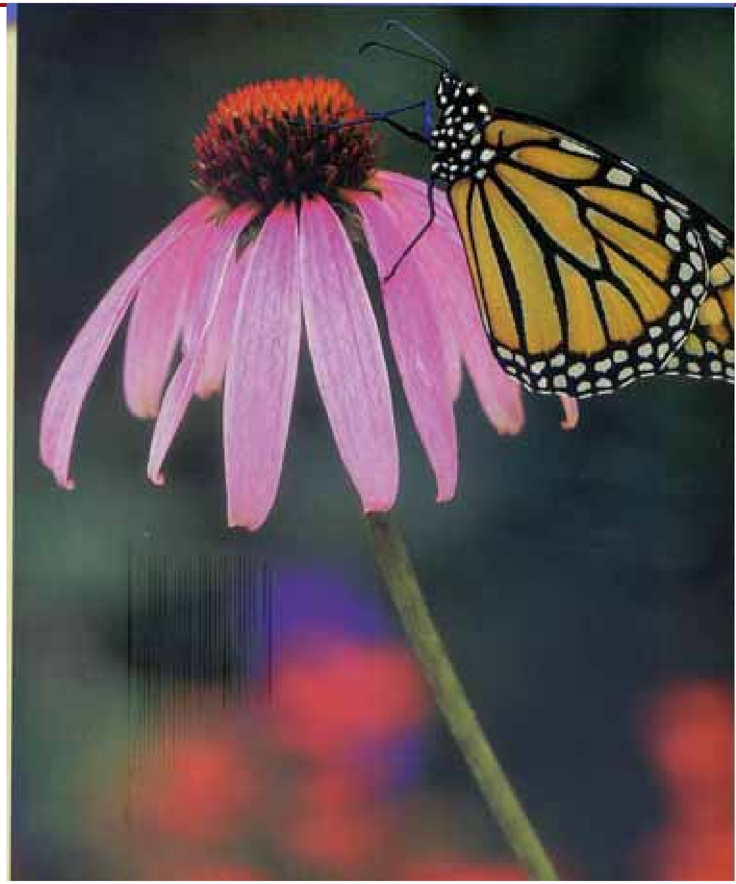
Chemistry

Zumdahl, 7th edition

歐亞書局

CH1 Chemical Foundations



Male Monarch butterflies use the pheromones produced by a gland on their wings to make themselves attractive to females.

Contents

Introduction

- 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.
- Chemistry is also important to historians.

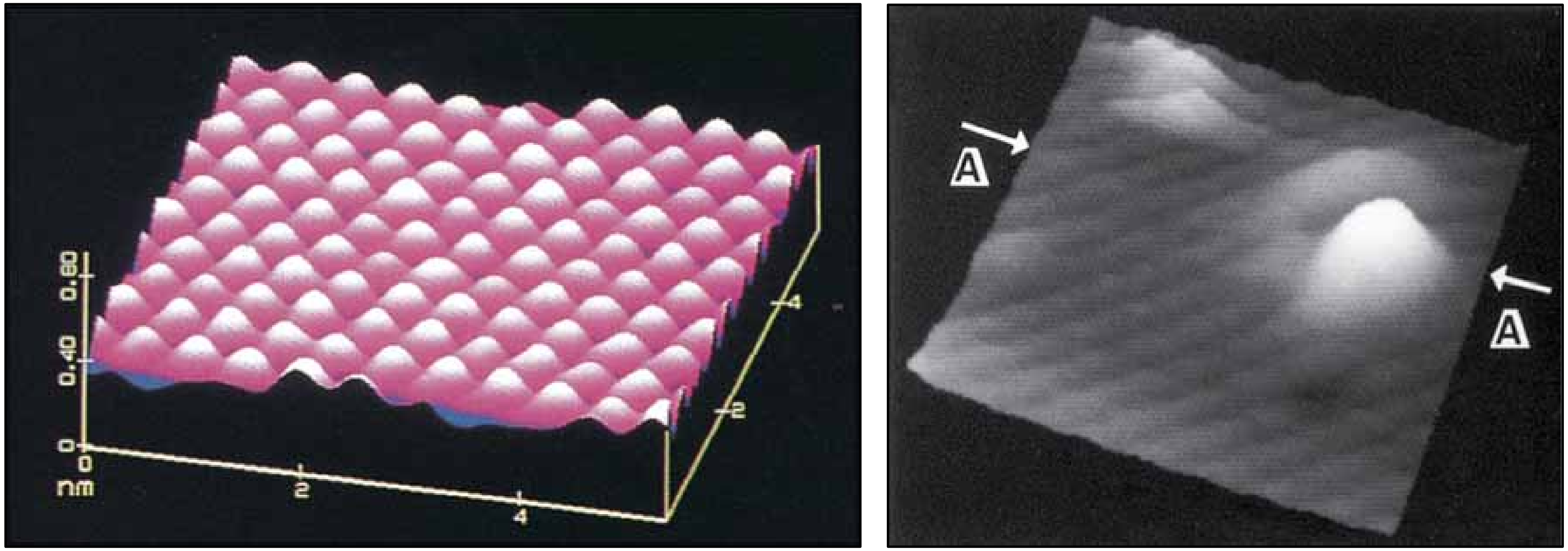
- Chemistry is also apparently very important in determining a person’s behavior.
- The importance of chemistry in the interactions of people should not really surprise us, since we know that insects communicate by emitting and receiving chemical signals via molecules called *pheromones*.

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.
- The STM pictures of several substances are shown in Fig. 1.1.

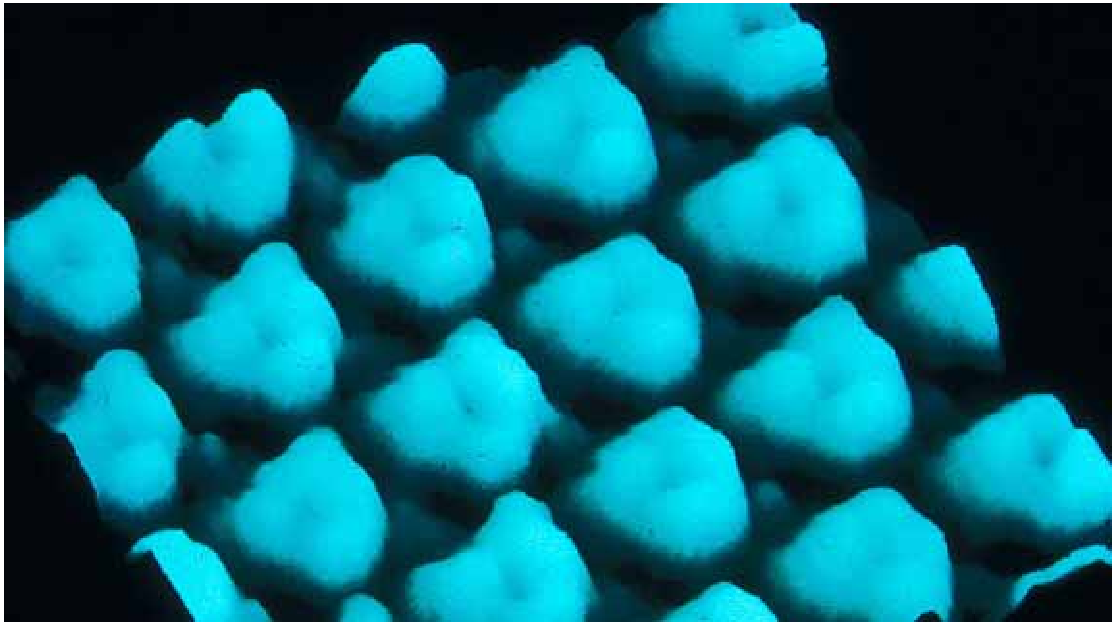
- The tiny white dot in the center of Fig. 1.2 is a single mercury atom that is held in a special trap.

Figure 1.1



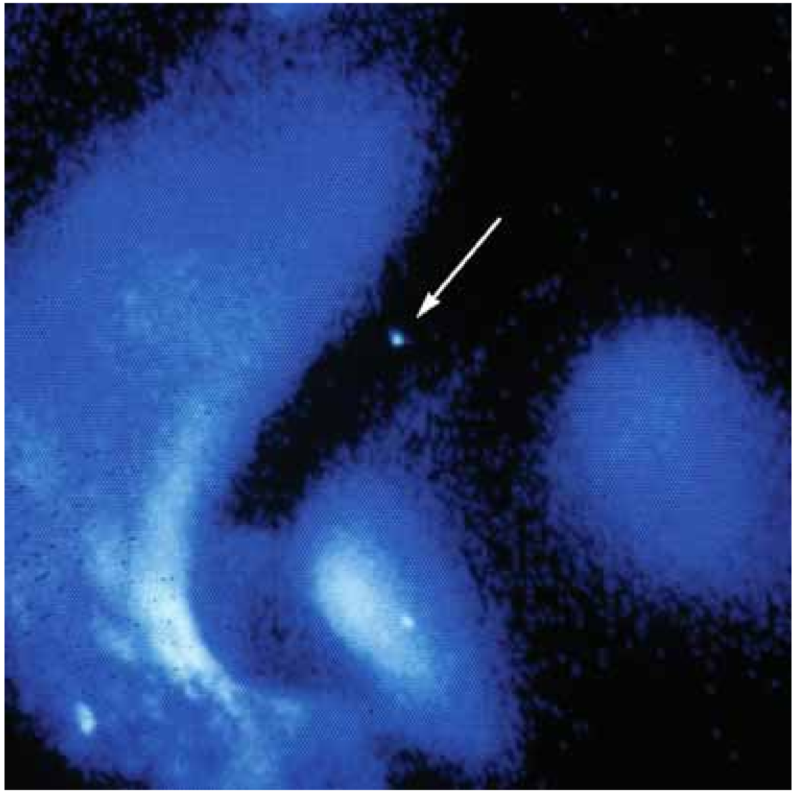
(a) The surface of a single grain of table salt. (b) An oxygen atom (indicated by arrow) on a gallium arsenide surface.

Figure 1.1



(c) Scanning tunneling microscope image showing rows of ring-shaped clusters of benzene molecules on a rhodium surface. Each “doughnut”-shaped image represents a benzene molecule.

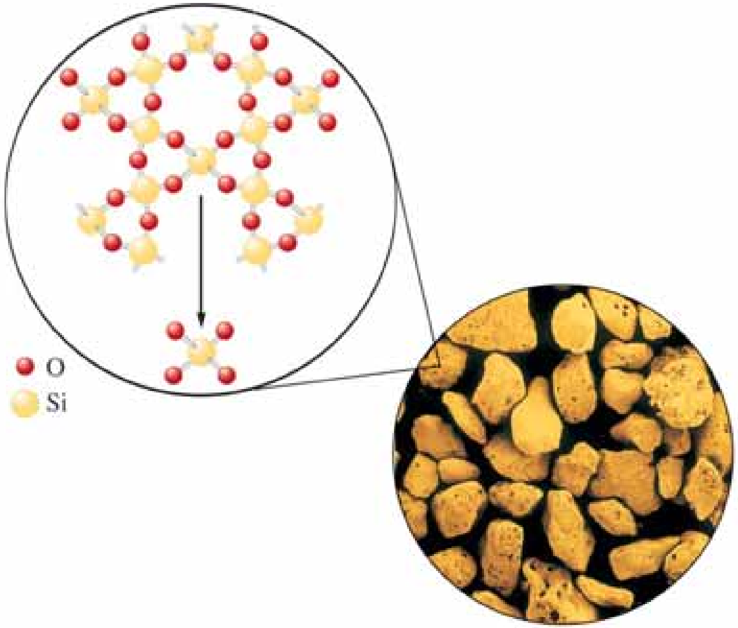
Figure 1.2

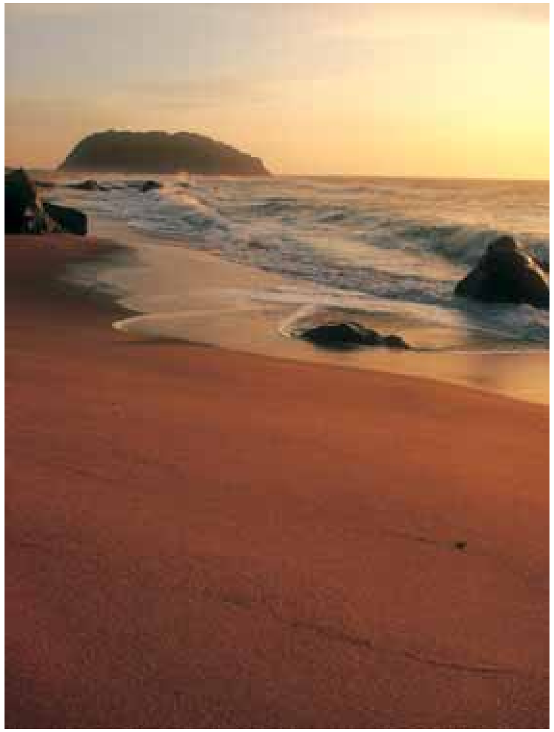


A charged mercury atom shows up as a tiny white dot (indicated by the arrow).

- 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. 1.3).
- One of the main challenges of chemistry is to understand the connection between the macroscopic world that we experience and the *microscopic world* of atoms and molecules.

Figure 1.3





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.

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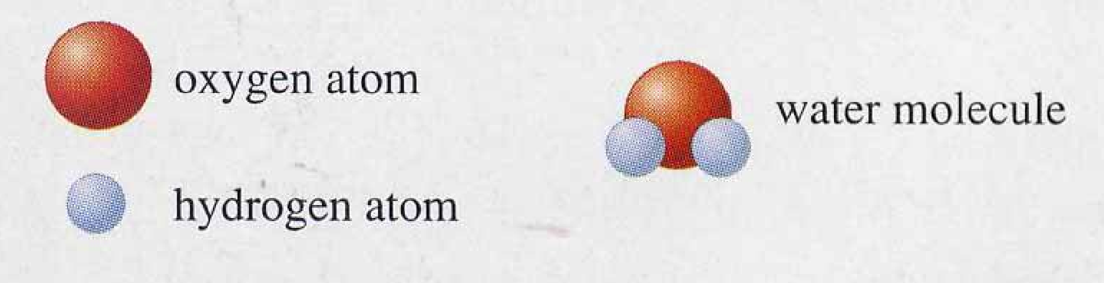
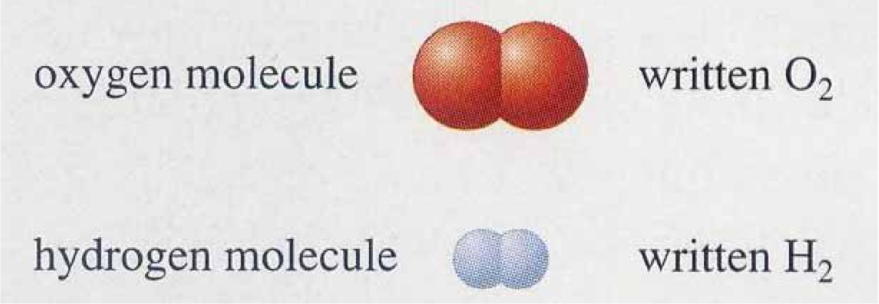
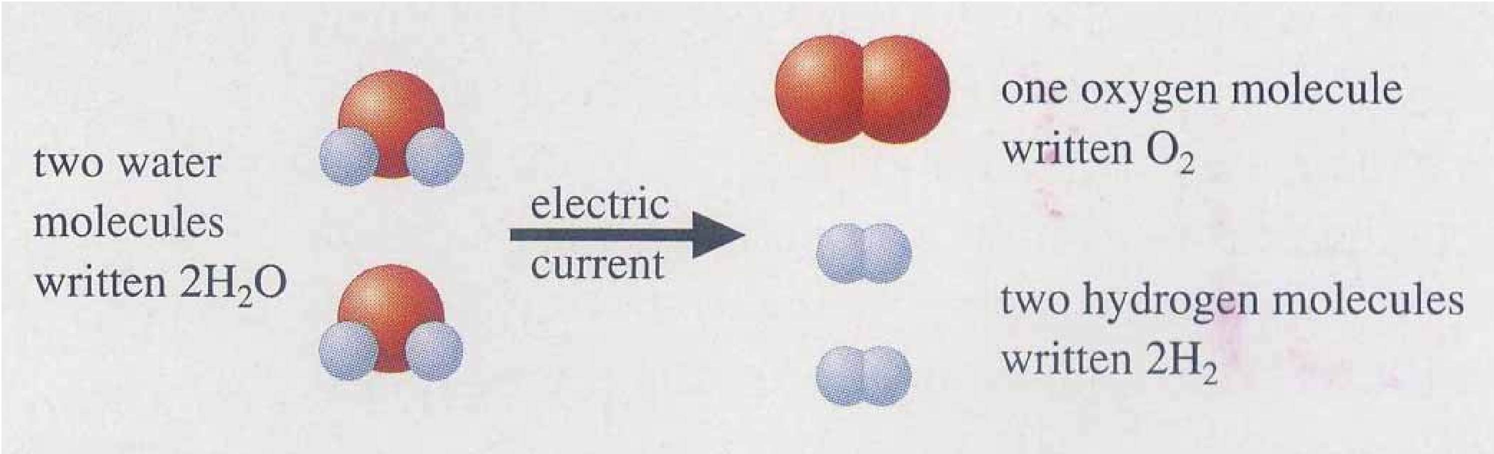
- There are two hydrogen atoms and one oxygen atom bound together to form the water molecule:

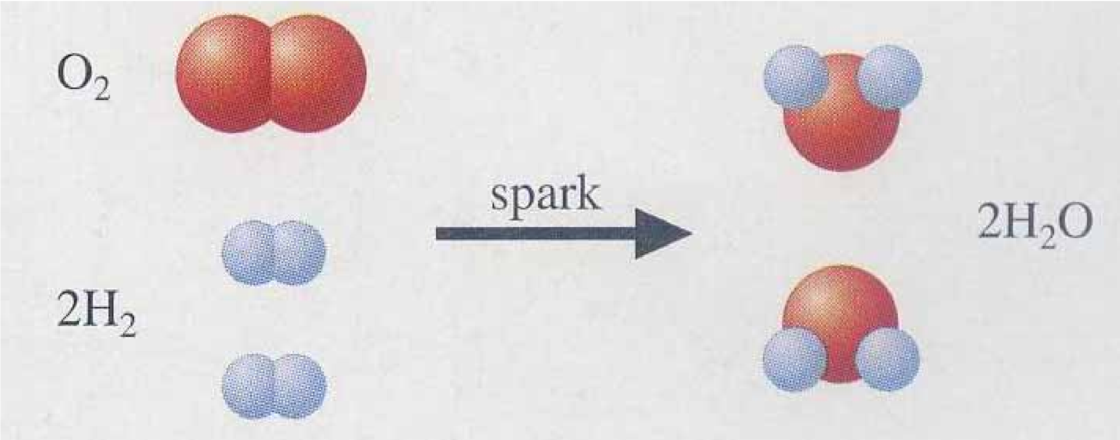


- These chemical elements themselves exist naturally as diatomic (two-atom) molecules:





- 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.

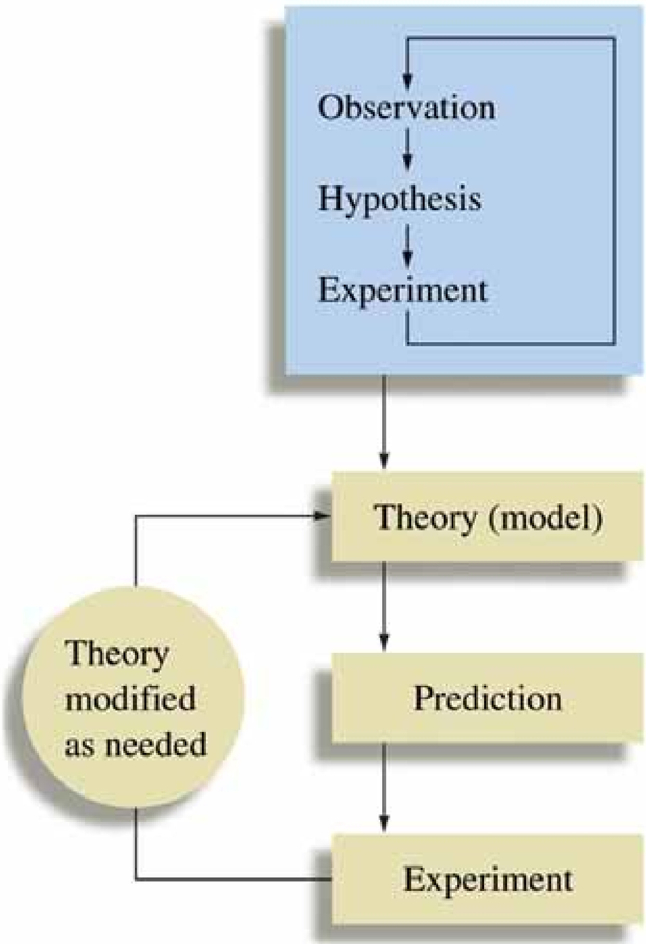
Science: A Process for Understanding Nature and Its Changes

- The *scientific method* :
  1. Making observations (collecting data)
  2. Making a prediction (formulating a hypothesis)
  3. Doing experiments to text the prediction (testing the hypothesis)

1.2 The Scientific Method

- The process that lies at the center of scientific inquiry is called the scientific method.
- It is useful to consider the following general framework for a generic scientific method (see Fig. 1.4):

Figure 1.4



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 quantitative 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.

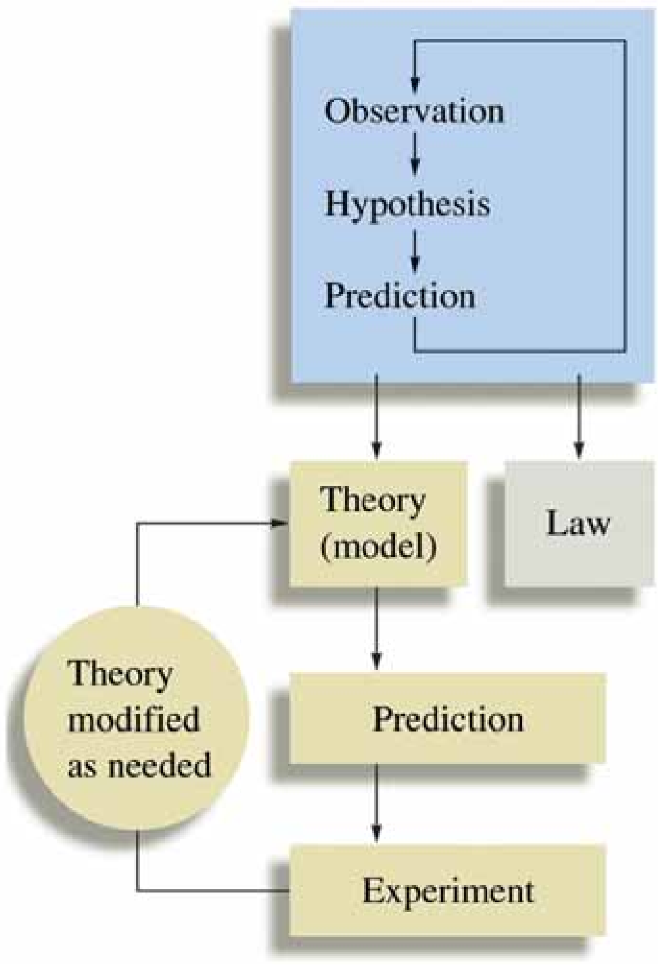
Scientific Models

- 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.

- As scientists observe nature, the 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**.

- 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 is happens.*

Figure 1.5



The various parts of the scientific method.

1.3 Units of Measurement

- A quantitative observation, or measurement, always consists of two parts: a *number* and a scale (called a *unit*).
- 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 Systeme International in French), or the **SI system**.

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

- This system is based on the metric system and units derived from the metric system. The fundamental SI units are listed in Table 1.1.
- 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 1.2. Some common objects and their measurements in SI units are listed in Table 1.3.

TABLE 1.1 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 1.2 The Prefixed 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}

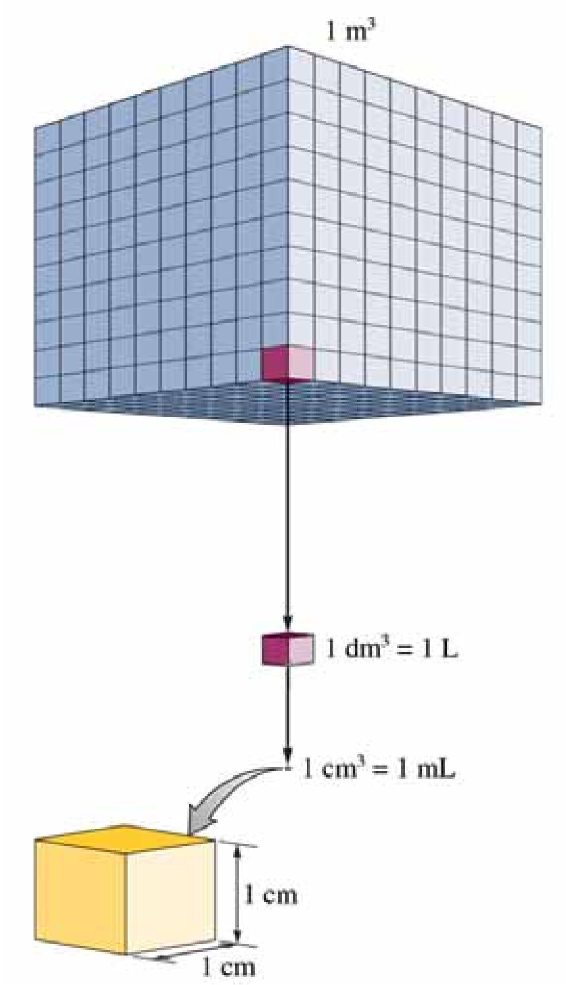
*See Appendix 1.1 if you need a review of exponential notation.

TABLE 1.3 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.

- 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. 1.6.

Figure 1.6



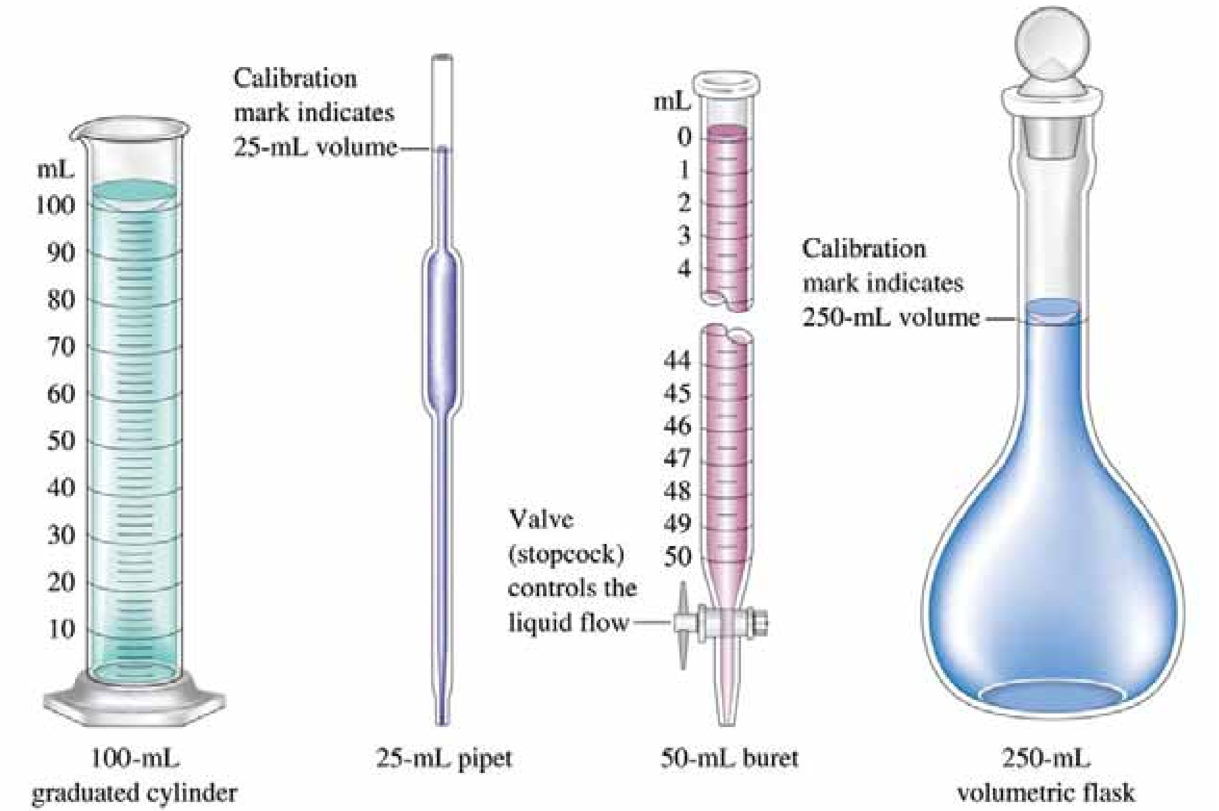
The largest cube has sides 1 m in length and a volume of $1 m^3$. The middle-sized cube has sides 1 dm in length and a volume of $1 dm^3$, or 1 L. The smallest cube has sides 1 cm in length and a volume of $1 cm^3$, or 1 mL.

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

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

- Several devices for the accurate determination of liquid volume are shown in Fig. 1.7.
- An important point concerning measurements is the relationship between mass and weight.

Figure 1.7



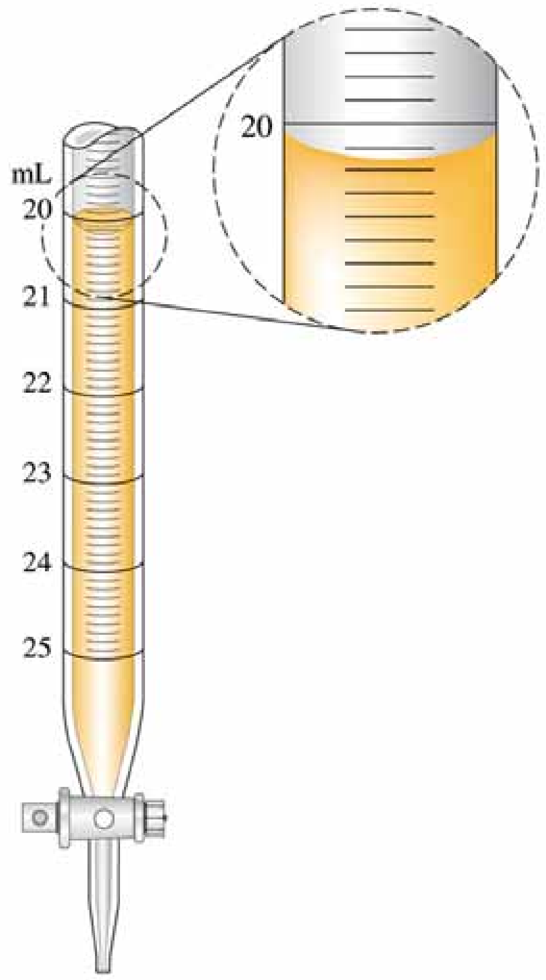
Common types of laboratory equipment used to measure liquid volume.

- ***Mass*** is a measure of the resistance of an object to change in its state of motion.
- On earth we use the force that gravity exerts on an object to measure its mass. We call this force the object's **weight**.
- Because weighing something on a chemical balance (see Fig. 1.8.) 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 1.8

An electronic analytical balance.

Figure 1.9



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

1.4 Uncertainty in Measurement

- The number associated with a measurement is obtained using some measuring device. (shown in Fig. 1.9 with the scale greatly magnified).

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

- 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.

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

- 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.
- 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.

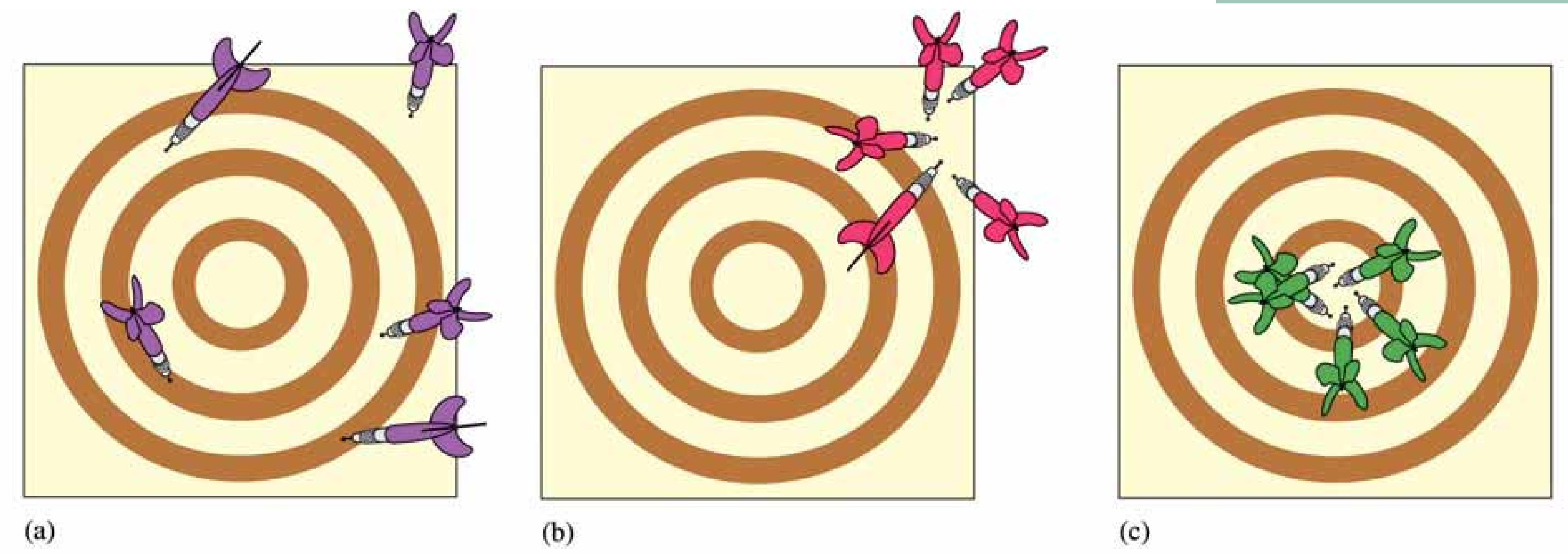
Precision and Accuracy

- Two terms often used to describe the reliability of measurements are precision and accuracy.
- **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. 1.10.
- Two different types of errors are illustrated in Fig. 1.10.
- A **random error** (also called an indeterminate error) means that a measurement has an equal probability of being high or low.
- The second type of error is called **systematic error** (or *determinate error*).

- Figure 1.10(a) indicates large random errors (poor technique).
- Figure 1.10(b) indicates small random errors but a large systematic error, and Figure 1.10(c) indicates small random errors and no systematic error.

Figure 1.10



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).

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

- 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.

1.5 Significant Figures and Calculations

- Calculating the final result for an experiment usually involves adding, subtracting, multiplying, or dividing the results of carious 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.

   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 example of exact numbers are the 2 in $2\pi r$ (the circumference of a circle) and the 4 and the 3 in $4/3\pi r^3$ (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.

- Note that the number 1.00×10^2 above is written in **exponential notation**.

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.

This number must be rounded to two significant figures.

1.6 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**.
- Some equivalents in the English and metric systems are listed in Table 1.4.
- A more complete list of conversion factors given to more significant figures appears in appendix 6.

TABLE 1.4 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	$1\ L = 1.06\ qt$
	$1\ ft^3 = 28.32\ L$

Sample Exercise 1.5 Unit Conversions 1

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 1.37 and 1.38

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 desire units.

Sample Exercise 1.8 Unit Conversion IV

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

Solution

Result obtained by rounding only at the end of the calculation

$$\frac{55\ \cancel{\text{mi}}}{\text{h}} \times \frac{1760\ \cancel{\text{yd}}}{1\ \cancel{\text{mi}}} \times \frac{1\ \cancel{\text{m}}}{1.094\ \cancel{\text{yd}}} \times \frac{1\ \text{km}}{1000\ \cancel{\text{m}}} = 88\ \text{km/h}$$

Sample Exercise 1.8

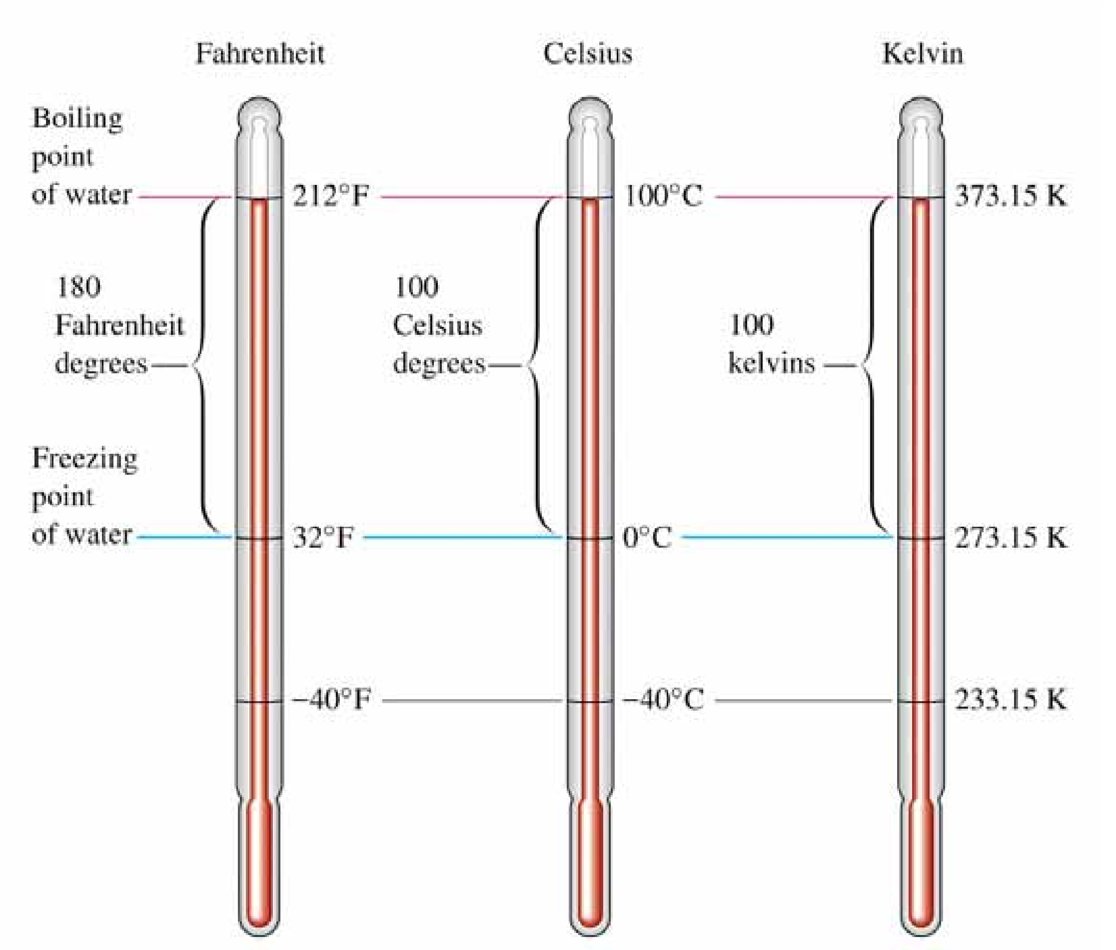
Note that all units cancel except the desired kilometers per hour.

See Exercises 1.43 through 1.45

1.7 Temperature

- Three systems for measuring temperature are widely used: the Celsius scale, the Kelvin scale, and the Fahrenheit scale.
- The three temperature scales are defined and compared in Fig. 1.11.

Figure 1.11



The three major temperature scales.

- The fundamental difference between these two temperature scales is in their zero points.
- Conversion between these two scales simply requires and adjustment for the different zero points.

Temperature (Kelvin) = temperature (Celsius) + 273.15

or

Temperature (Celsius) = temperature (Kelvin) − 273.15

- Notice that since 212°F = 100°C and 32°F = 0°C,

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

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

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

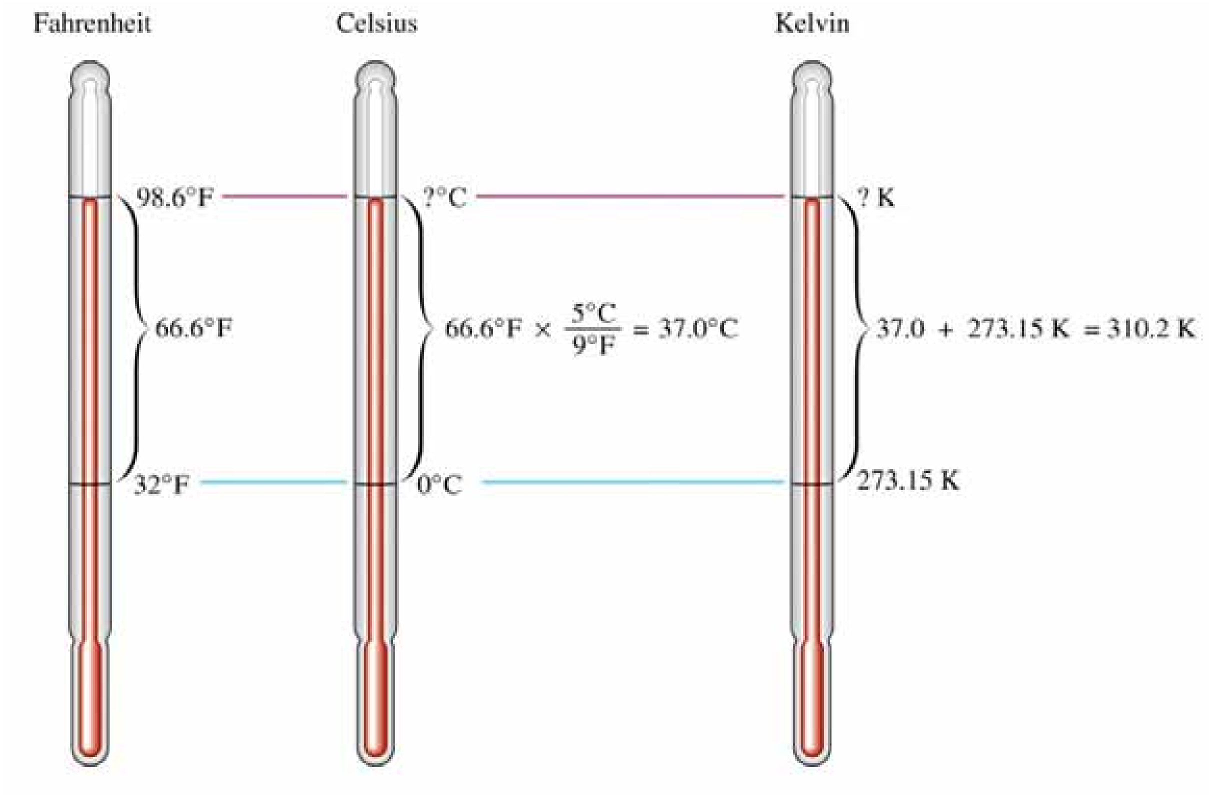
- Since 32°F = 0°C, we obtain the corresponding Celsius temperature by first subtracting 32 from the Fahrenheit temperature to account for the different zero points.
- Then the unit factor is applied to adjust for the difference in the degree size.
- This process is summarized by the equation

$$(T_F - 32°F)\frac{5°C}{9°F} = T_C \qquad (1.1)$$

- where T_F and T_C represent a given temperature on the Fahrenheit and Celsius scales respectively.

$$T_F = T_C \times \frac{9°F}{5°C} + 32°F \tag{1.2}$$

Figure 1.12



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

1.8 Density

- **Density**: the mass of substance per unit volume of the substance:

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

- The densities of various common substances are given in Table 1.5.

TABLE 1.5 Densities of Various Common Substances at 20℃

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

1.9 Classification of Matter

- **Matter**, is the material of the universe.
- 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.
- Molecular-level pictures of the three states of water are given in Fig. 1.13.

Figure 1.13

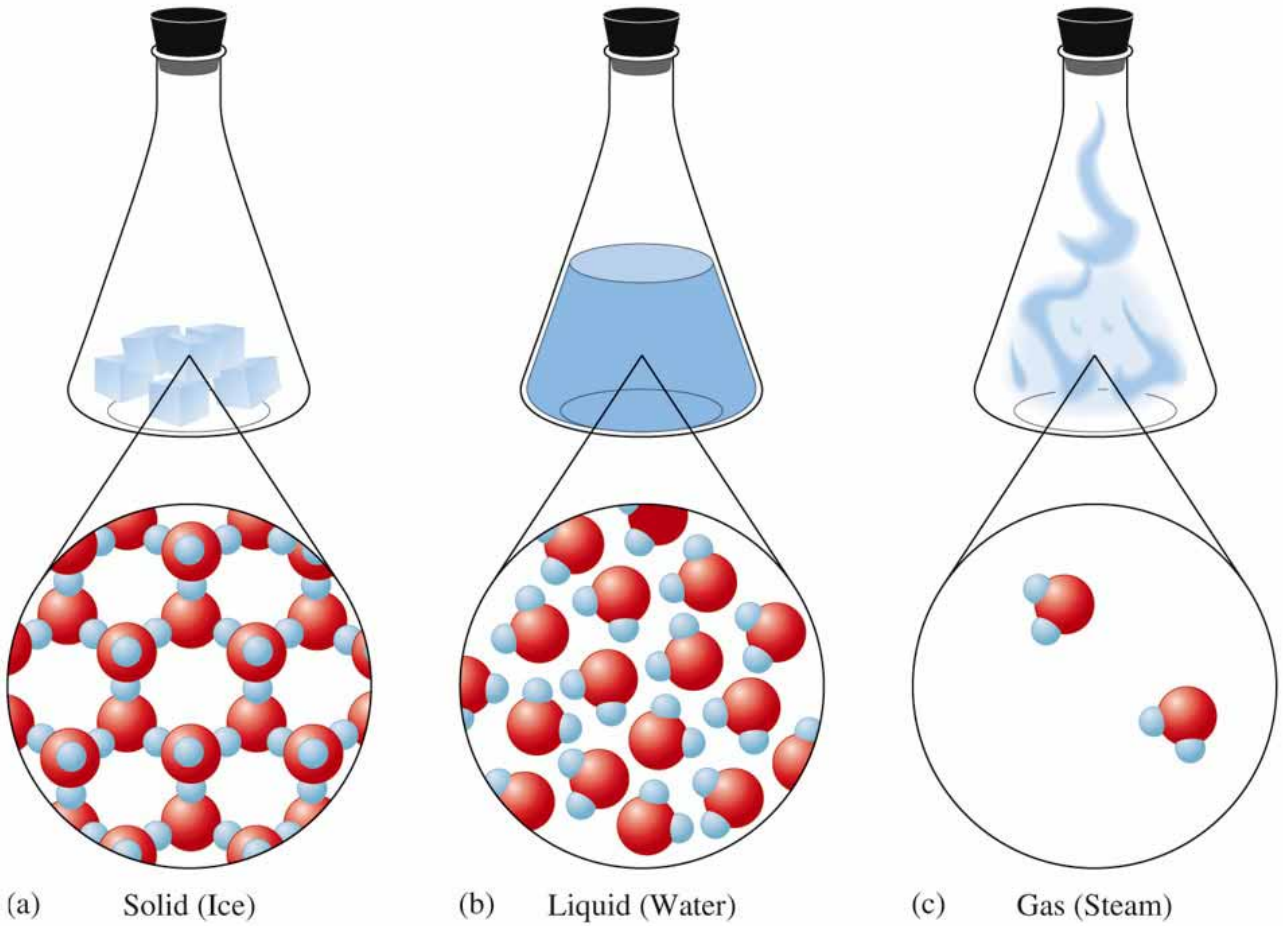


Figure 1.13

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.

- The different properties of ice, liquid water, and steam are determined by the different arrangements of the molecules in theses substance.
- Table 1.5 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.
- The main characteristic of a mixture is that it has *variable composition*.

- Mixtures can be classified as **homogeneous** (having visibly indistinguishable parts) or **heterogeneous** (having visibly distinguishable parts).
- A homogeneous mixture is called a **solution**.
- A **pure substance** is one with constant composition.
- 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.

1. Distillation:

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. 1.14.

Figure 1.14

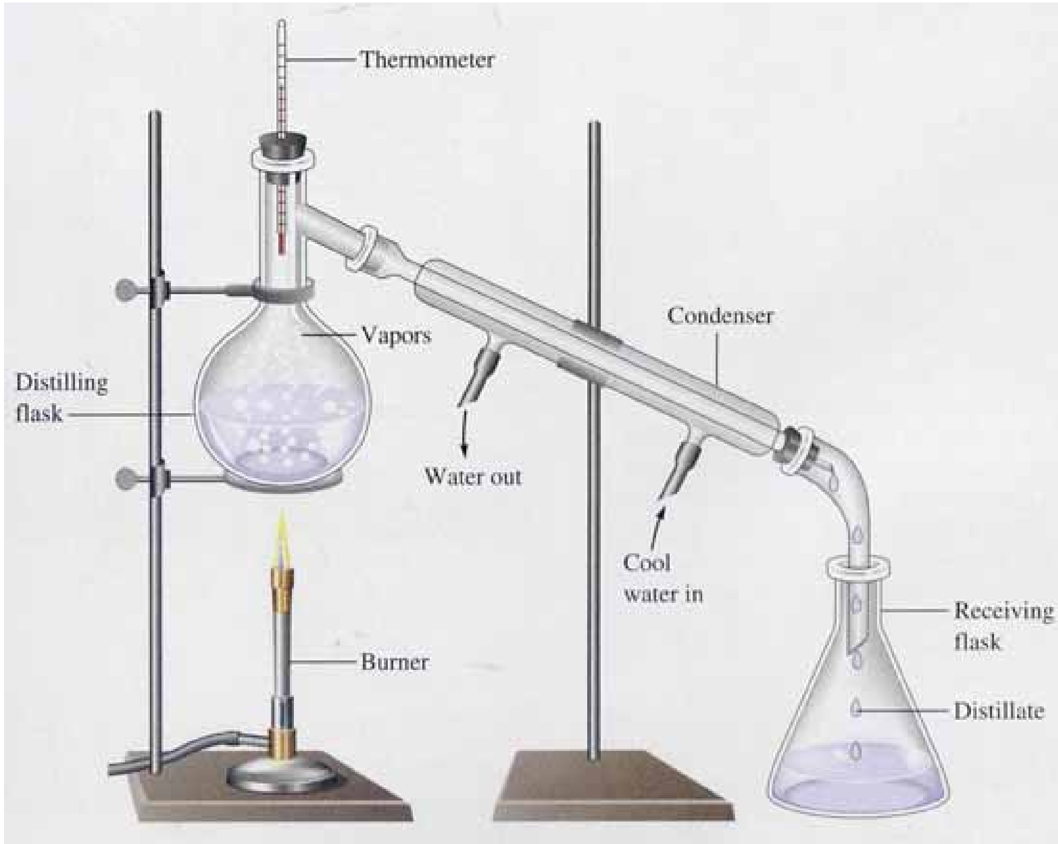

Thermometer
Vapors
Distilling flask
Condenser
Water out
Cool water in
Receiving flask
Burner
Distillate

Figure 1.14

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.

- 2. Filtration

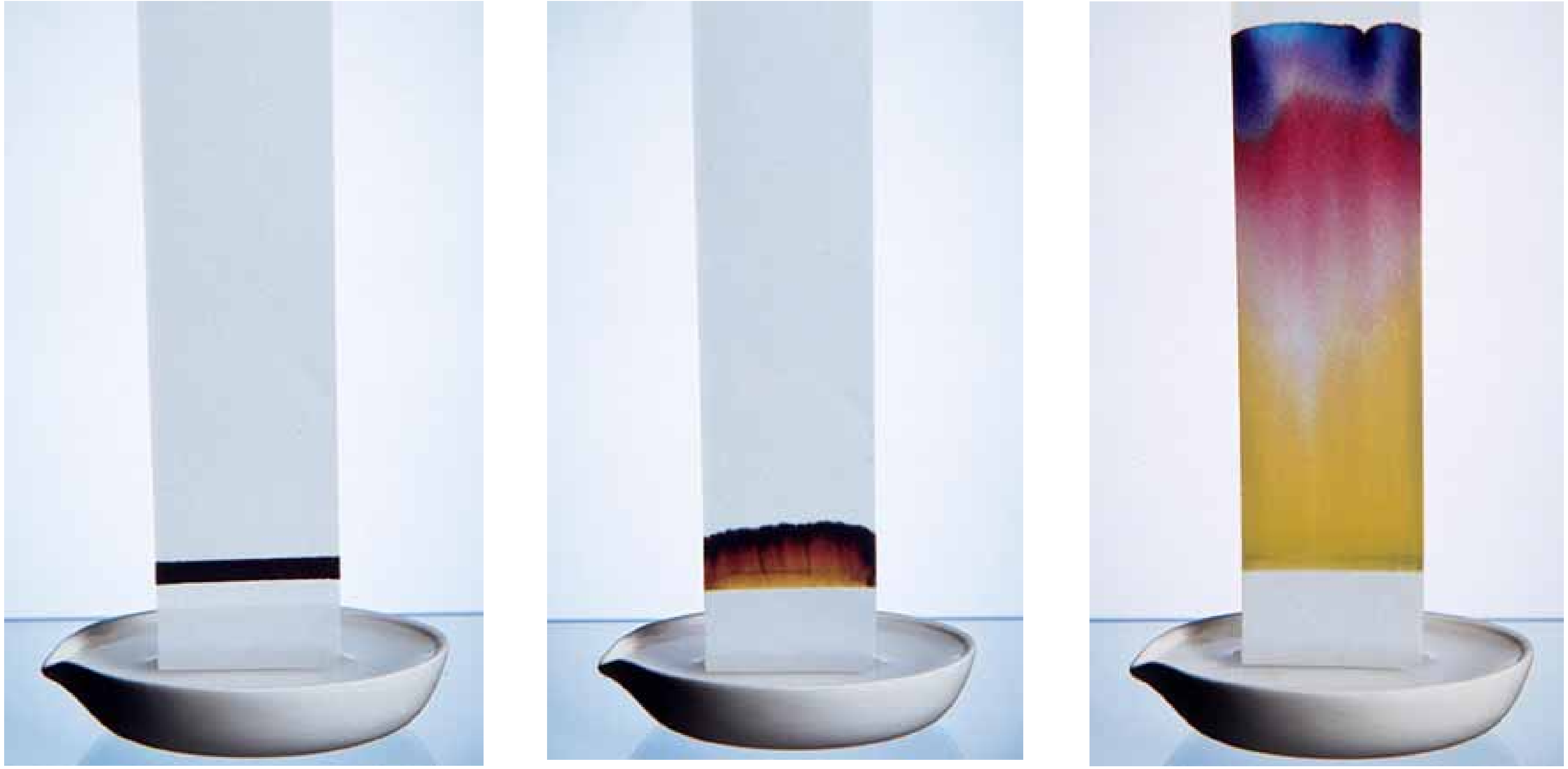
  Another method of separation is simple **filtration**, which is used when a mixture consists of a solid and a liquid. The mixture is poured onto a mesh, such as filter paper, which passes the liquid and leaves the solid behind.

- 3. Chromatography

  A third method of separation is called chromatography. **Chromatography** is the general name applied to a series of methods that employ a system wit two *phase* (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.

- 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(燭心-蠟油往上燃燒） (see fig. 1.15).

Figure 1.15



(a) (b) (c)

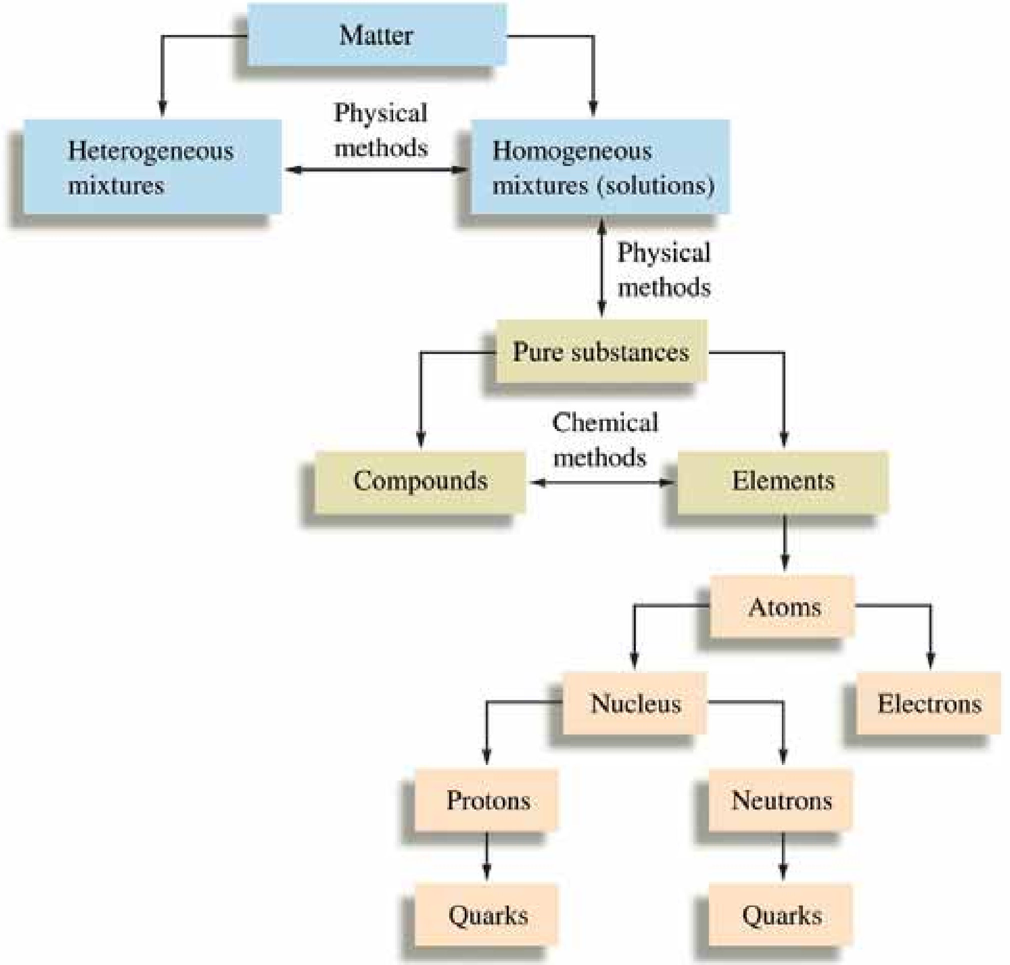
Figure 1.15

Paper chromatography of ink. (a) A line 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. (c) the component with the weakest attraction for the paper travels faster than the components that cling to the paper.

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.

- A **compound** is a substance with constant composition that can be broken down into elements by chemical processes.
- 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.
- Figure 1.16 summarizes our discussion of the organization of matter.

Figure 1.16



The organization of matter.