Chapter 1

Introduction: Matter, Energy, and Measurement

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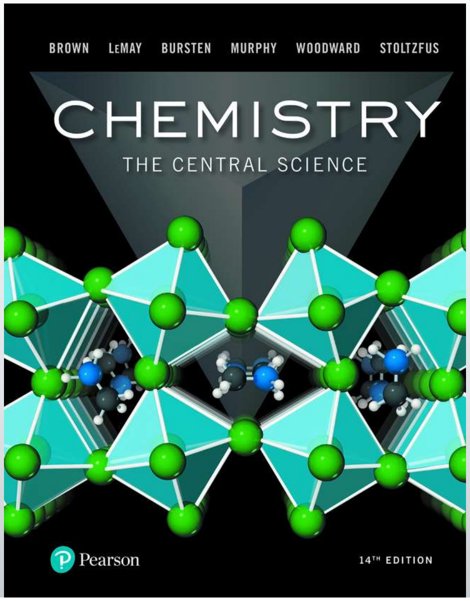
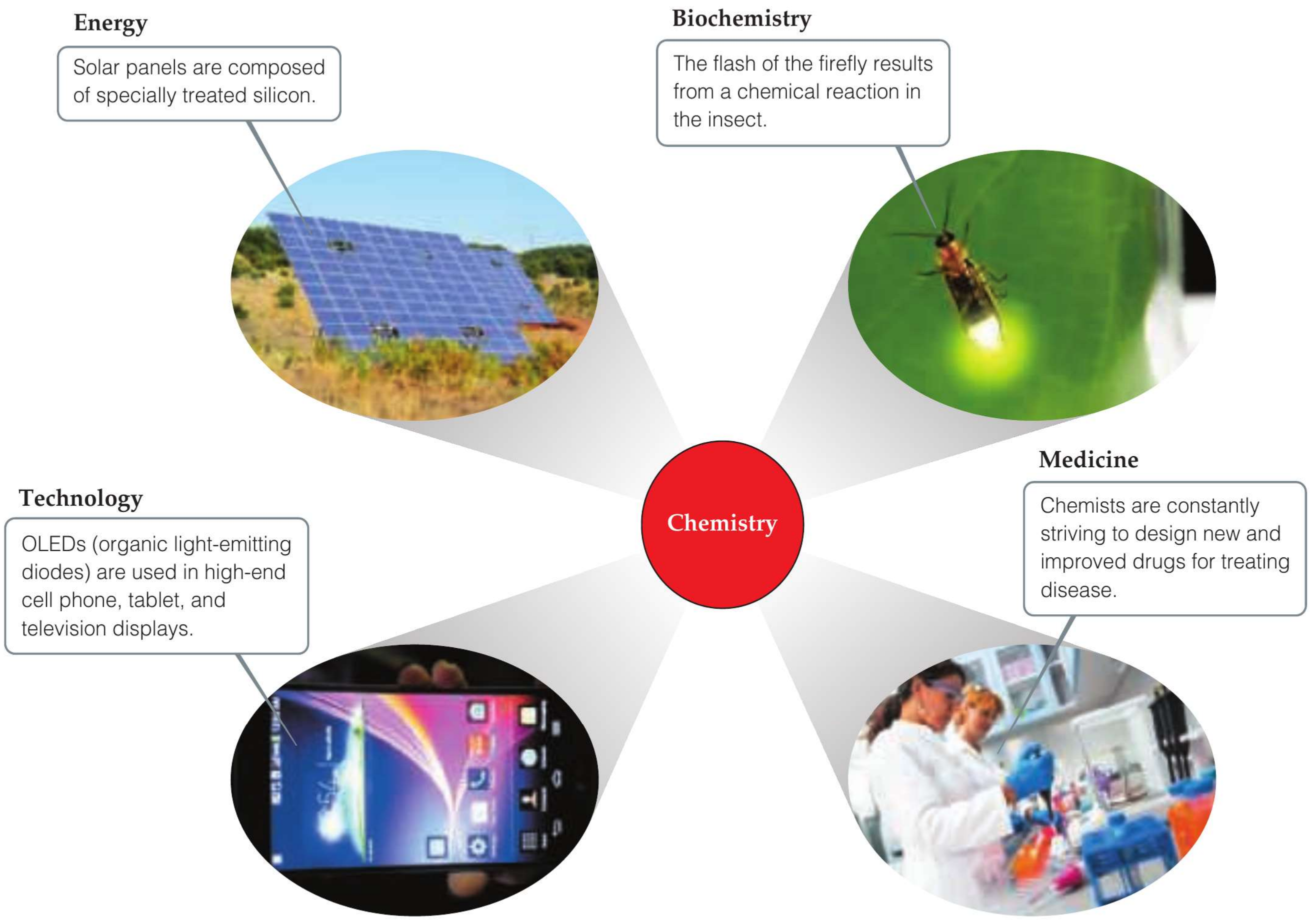


1.1 The Study of Chemistry

Why Study Chemistry?

- You may be studying chemistry because it is an essential part of your curriculum.
- Chemistry is central to a fundamental understanding of governing principles in many science-related fields.



▲ Figure 1.2 Chemistry is central to our understanding of the world around us.

1.1 The Study of Chemistry

- Chemistry is the science that seeks to understand the properties and behavior of matter by studying the properties and behavior of atoms and molecules.

The Atomic and Molecular Perspective of Chemistry

- Matter is anything that has mass and occupies space.
- A property is any characteristic that allows us to recognize a particular type of matter and to distinguish it from other types.
- Atoms are the small building blocks of matter.
- Molecules: two or more atoms are joined in specific shapes.

1.2 Classifications of Matter

States of Matter

- A gas (also known as vapor) has no fixed volume or shape; rather, it uniformly fills its container. A gas can be compressed to occupy a smaller volume, or it can expand to occupy a larger one.
- A liquid has a distinct volume independent of its container, assumes the shape of the portion of the container it occupies, and is not compressible to any appreciable extent.
- A solid has both a definite shape and a definite volume and is not compressible to any appreciable extent.

1.2 Classifications of Matter

Pure Substances

- A pure substance (usually referred to simply as a substance) is matter that has distinct properties and a composition that does not vary from sample to sample.
- All substances are either elements or compounds.

Elements

- Elements are substances that cannot be decomposed into simpler substances. On the molecular level, each element is composed of only one kind of atom.
- Currently, 118 elements are known, though they vary widely in abundance.

1.2 Classifications of Matter

Compounds

- Compounds are substances composed of two or more elements.
- **Example**: water is a compound composed of two elements; hydrogen and oxygen.

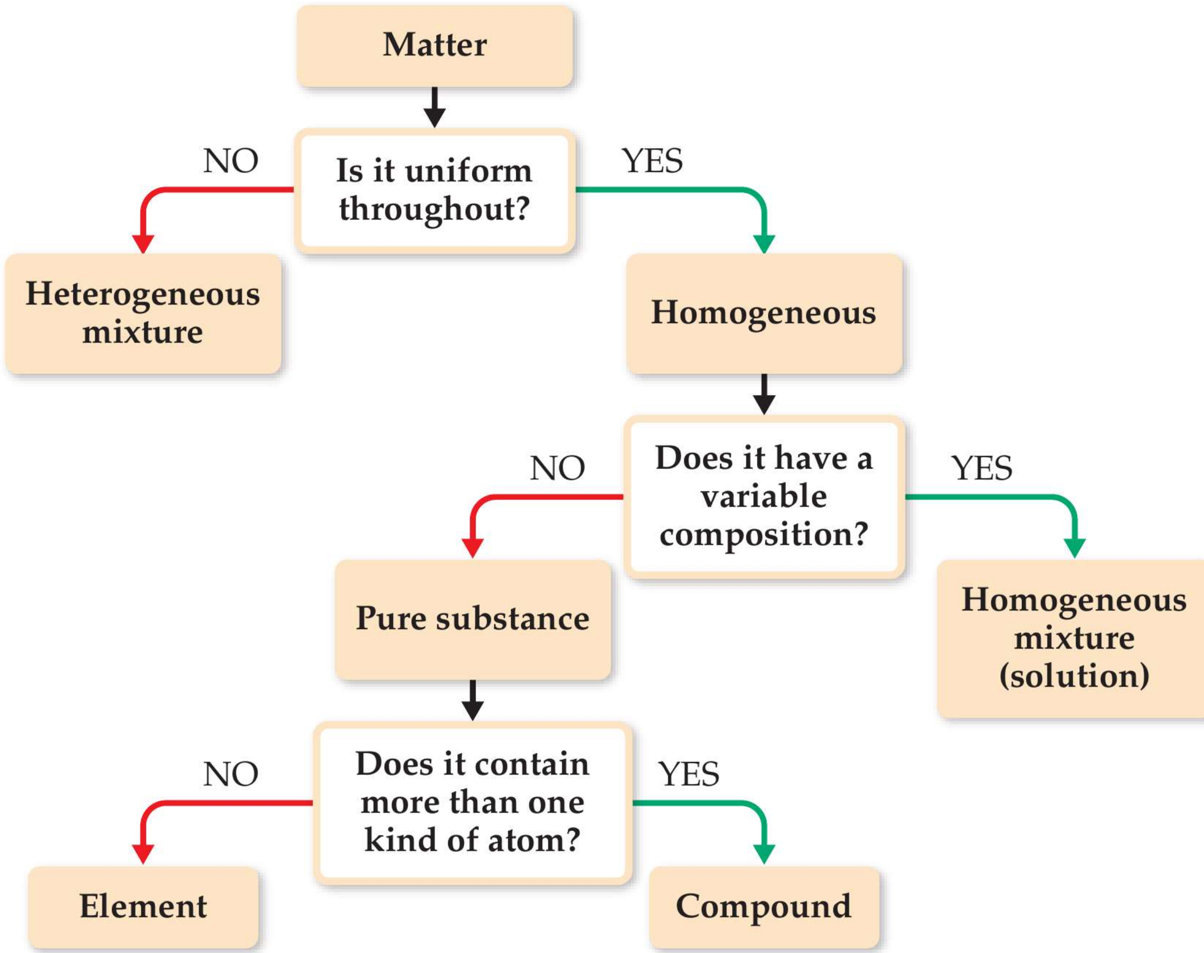
- The observation that the elemental composition of a compound is always the same is known as the law of constant composition (or the law of definite proportions).

1.2 Classifications of Matter

Mixtures

- Mixtures are combinations of two or more substances in which each substance retains its chemical identity.
- The substances making up a mixture are called components of the mixture.
- Heterogeneous mixtures are mixtures that vary in texture and appearance in any typical sample, such as rocks and wood.
- Homogeneous mixtures are mixtures that are uniform throughout, such as air, salt + water, and sugar + water.
- Homogeneous mixtures are also called **solutions**.

1.2 Classifications of Matter



Sample Exercise 1.1

“White gold” contains gold and a “white” metal, such as palladium. Two samples of white gold differ in the relative amounts of gold and palladium they contain. Both samples are uniform in composition throughout. Classify white gold.

- Because the material is uniform throughout, it is homogeneous.
- Because its composition differs for the two samples, it cannot be a compound.
- Instead, it must be a homogeneous mixture.

1.3 Properties of Matter

- Physical properties can be observed without changing the identity and composition of the substance.
- **Examples**: color, odor, density, melting point, boiling point, and hardness.
- Chemical properties describe the way a substance may change, or react, to form other substances.
- **Example**: flammability; the ability of a substance to burn in the presence of oxygen.
- Intensive properties are properties that do not depend on the amount of sample being examined.
- **Examples**: temperature and melting point.

1.3 Properties of Matter

- Extensive properties are properties that depend on the amount of sample being examined.
- **Examples**: mass and volume.

Physical and Chemical Changes

- In a physical change, a substance changes its physical appearance but not its composition, such as changes of state (from liquid to gas, or from liquid to solid, etc).
- In a chemical change (also called a chemical reaction), a substance is transformed into a chemically different substance, such as hydrogen burns in air to form water.

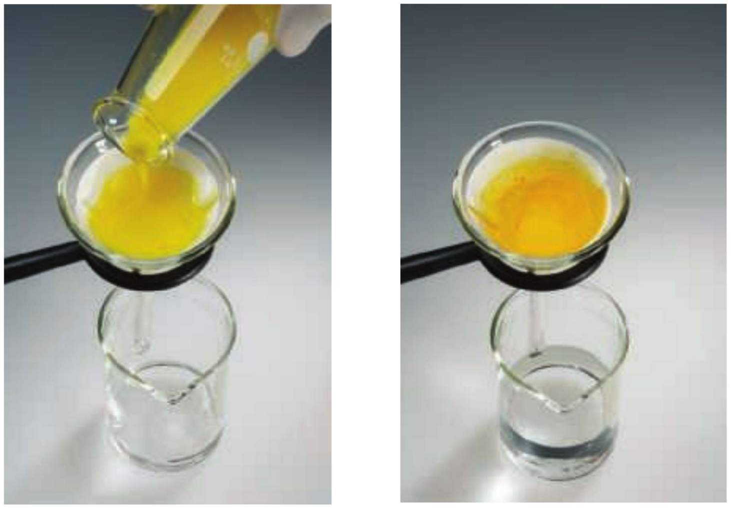
Separation of Mixtures

- We can separate a mixture into its components by taking advantage of differences in their properties.
- **Example**: a heterogeneous mixture of iron filings and gold filings could be sorted by color into iron and gold. A less tedious approach would be to use a magnet to attract the iron filings, leaving the gold ones behind. If we put our mixture into an appropriate acid, the acid would dissolve the iron and the solid gold would be left behind; the two could then be separated by filtration.

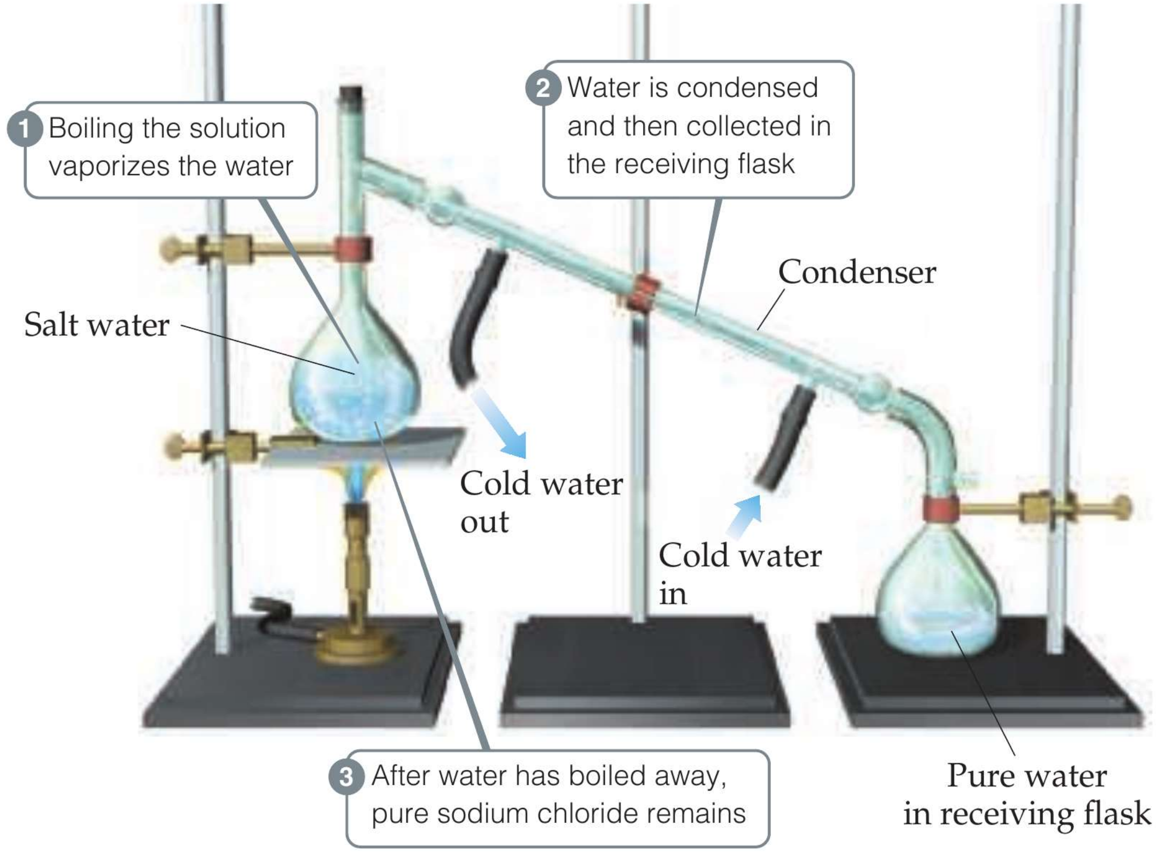
1.3 Properties of Matter

- An important method of separating the components of a homogeneous mixture is **distillation**, a process that depends on the different abilities of substances to form gases.
- **Example**: if we boil a solution of salt and water, the water evaporates, forming a gas, and the salt is left behind. The gaseous water can be converted back to a liquid on the walls of a condenser.
- The differing abilities of substances to adhere to the surfaces of solids can also be used to separate mixtures; this ability is the basis of **chromatography**.

1.3 Properties of Matter



▲ **Figure 1.12 Separation by filtration.** A mixture of a solid and a liquid is poured through filter paper. The liquid passes through the paper while the solid remains on the paper.



▲ **Figure 1.13 Distillation.** Apparatus for separating a sodium chloride solution (salt water) into its components.

1.5 Units of Measurement

- Many properties of matter are **quantitative**, that is, associated with numbers.
- When a number represents a measured quantity, the units of that quantity must be specified.

SI Units

- The metric units used in scientific measurements are called SI units.
- This system has seven base units from which all other units are derived.
- With SI units, **prefixes** are used to indicate decimal fractions or multiples of various units.

1.5 Units of Measurement

TABLE 1.3 SI Base Units

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

1.5 Units of Measurement

TABLE 1.4 Prefixes Used in the Metric System and with SI Units

Prefix	Abbreviation	Meaning	Example	
Peta	P	10^{15}	1 petawatt (PW)	$= 1 \times 10^{15}$ watts[a]
Tera	T	10^{12}	1 terawatt (TW)	$= 1 \times 10^{12}$ watts
Giga	G	10^{9}	1 gigawatt (GW)	$= 1 \times 10^{9}$ watts
Mega	M	10^{6}	1 megawatt (MW)	$= 1 \times 10^{6}$ watts
Kilo	k	10^{3}	1 kilowatt (kW)	$= 1 \times 10^{3}$ watts
Deci	d	10^{-1}	1 deciwatt (dW)	$= 1 \times 10^{-1}$ watt
Centi	c	10^{-2}	1 centiwatt (cW)	$= 1 \times 10^{-2}$ watt
Milli	m	10^{-3}	1 milliwatt (mW)	$= 1 \times 10^{-3}$ watt
Micro	μ[b]	10^{-6}	1 microwatt (μW)	$= 1 \times 10^{-6}$ watt
Nano	n	10^{-9}	1 nanowatt (nW)	$= 1 \times 10^{-9}$ watt
Pico	p	10^{-12}	1 picowatt (pW)	$= 1 \times 10^{-12}$ watt
Femto	f	10^{-15}	1 femtowatt (fW)	$= 1 \times 10^{-15}$ watt
Atto	a	10^{-18}	1 attowatt (aW)	$= 1 \times 10^{-18}$ watt
Zepto	z	10^{-21}	1 zeptowatt (zW)	$= 1 \times 10^{-21}$ watt

[a]The watt (W) is the SI unit of power, which is the rate at which energy is either generated or consumed. The SI unit of energy is the joule (J); $1\ \mathrm{J} = 1\ \mathrm{kg \cdot m^2/s^2}$ and $1\ \mathrm{W} = 1\ \mathrm{J/s}$.

1.5 Units of Measurement

Length and Mass

- Mass is a measure of the amount of material in an object.

Temperature

- Temperature, a measure of the hotness or coldness of an object, is a physical property that determines the direction of heat flow.
- Heat always flows spontaneously from a substance at higher temperature to one at lower temperature.
- The temperature scales commonly employed in science are the Celsius and Kelvin scales.

1.5 Units of Measurement

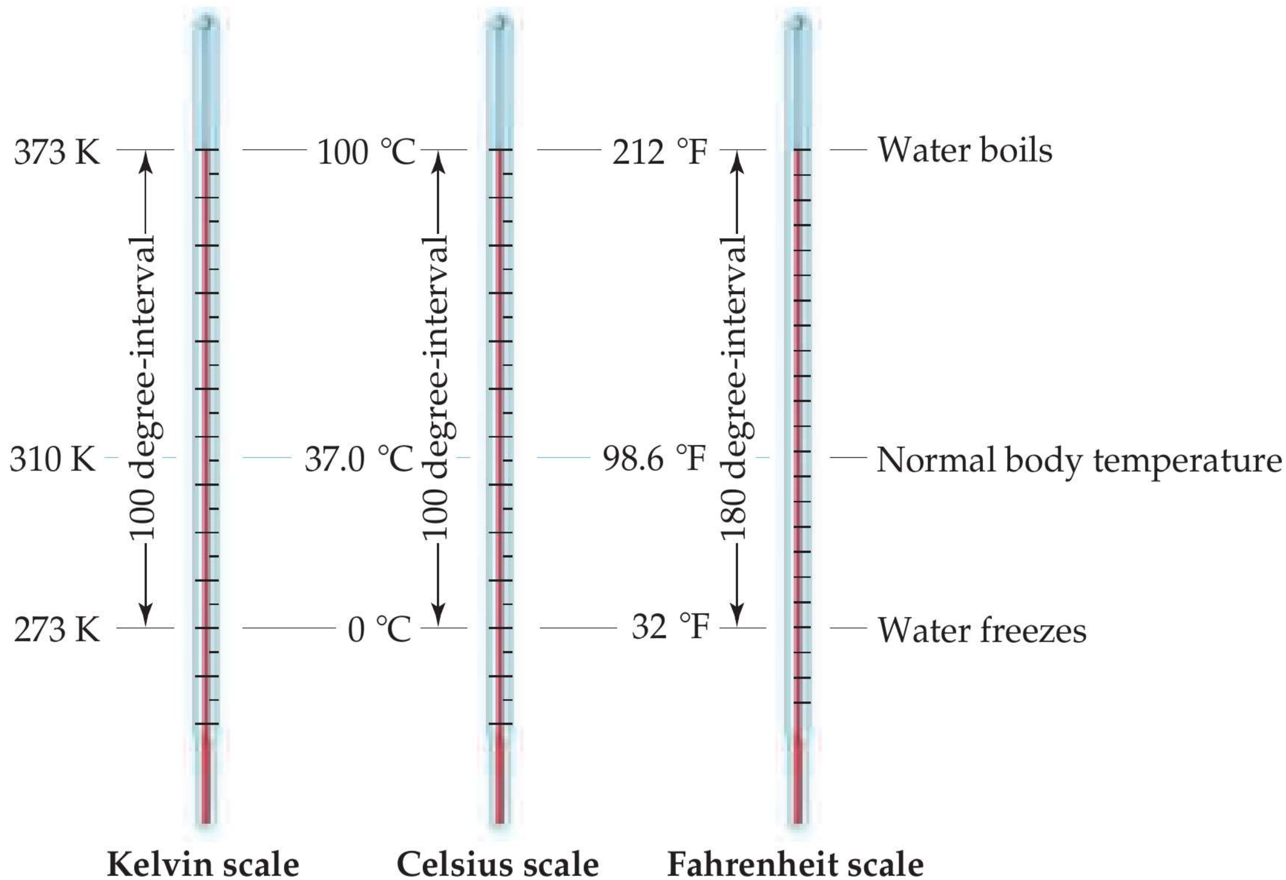
- The **Celsius scale** was originally based on the assignment of 0 °C to the freezing point of water and 100 °C to its boiling point at sea level.
- Zero on the Kelvin scale is the temperature at which all thermal motion ceases, a temperature referred to as absolute zero.
- The common temperature scale in the United States is the **Fahrenheit scale**, which is not generally used in science.

1.5 Units of Measurement

$$t_{K} = t_{^\circ C} + 273.15$$

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

$$t_{^\circ F} = \frac{9}{5}t_{^\circ C} + 32$$



▲ Figure 1.19 Comparison of the Kelvin, Celsius, and Fahrenheit temperature scales.

Sample Exercise 1.2

What is the name of the unit that equals

a. 10^{-9} g → nanogram, ng

b. 10^{-6} second → microsecond, μs

c. 10^{-3} meter → millimeter, mm

Sample Exercise 1.3

A weather forecaster predicts the temperature will reach 31 °C. What is this temperature

a. K

$$t_{\mathrm{K}} = t_{^\circ\mathrm{C}} + 273.15$$

$$t_{\mathrm{K}} = 31 + 273.15 = 304\ \mathrm{K}$$

b. °F

$$t_{^\circ\mathrm{C}} = \frac{5}{9}(t_{^\circ\mathrm{F}} - 32)$$

$$t_{^\circ\mathrm{F}} = \frac{9}{5}(31) + 32 = 88\ ^\circ\mathrm{F}$$

1.5 Units of Measurement

Derived SI Units

- A derived unit is obtained by multiplication or division of one or more of the base units.

Volume

- The volume of a cube is its length cubed, length^3.
- Smaller units, such as cubic centimeters, cm^3 (sometimes written cc), are frequently used in chemistry.
- Another volume unit used in chemistry is the liter (L).

$$1\ L = 1\ dm^3 \qquad\qquad 1\ mL = 1\ cm^3$$

1.5 Units of Measurement

Density

- Density (d) is defined as the amount of mass (m) in a unit volume (V) of a substance:

$$d = \frac{m}{V}$$

- The densities of solids and liquids are commonly expressed in either grams per cubic centimeter (g/cm^3) or grams per milliliter (g/mL).
- Densities are temperature dependent.
- If no temperature is reported, we assume 25 °C, close to normal room temperature.

1.5 Units of Measurement

Units of Energy

- The SI unit for energy is the joule.
- Joule is a derived unit, $1\ \mathrm{J} = 1\ \mathrm{kg \cdot m^2/s^2}$.
- It is still quite common in chemistry, biology, and biochemistry to find energy changes associated with chemical reactions expressed in the non-SI unit of calories.
- A calorie (cal) was originally defined as the amount of energy required to raise the temperature of 1 g of water 1 °C.

$$1\ \mathrm{cal} = 4.184\ \mathrm{J}$$

Sample Exercise 1.4

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

$$d = \frac{m}{V} = \frac{1.00 \times 10^2}{7.36} = 13.6 \ \mathrm{g/cm^3}$$

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

$$V = \frac{m}{d} = \frac{65.0}{0.791} = 82.2 \ \mathrm{mL}$$

Sample Exercise 1.4

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

$$m = dV = (19.32)\left(2.00^3\right) = 155\ \mathrm{g}$$

1.6 Uncertainty in Measurement

- Two kinds of numbers are encountered in scientific work: **exact numbers** (those whose values are known exactly) and **inexact numbers** (those whose values have some uncertainty).
- **Example**: there are exactly 12 eggs in a dozen, exactly 1000 g in a kilogram, and exactly 2.54 cm in an inch.
- Numbers obtained by measurement are always inexact.
- The equipment used to measure quantities always has inherent limitations (equipment errors), and there are differences in how different people make the same measurement (human errors).

1.6 Uncertainty in Measurement

Precision and Accuracy

- Precision is a measure of how closely individual measurements agree with one another.
- Accuracy refers to how closely individual measurements agree with the correct, or true, value.

Good accuracy
Good precision

Poor accuracy
Good precision

Poor accuracy
Poor precision

1.6 Uncertainty in Measurement

Significant Figures

- Suppose you determine the mass of a dime on a balance capable of measuring to the nearest 0.0001 g.
- You could report the mass as 2.2405 ± 0.0001 g.
- The ± notation expresses the magnitude of the uncertainty of your measurement.
- All digits of a measured quantity, including the uncertain one, are called significant figures.
- The greater the number of significant figures, the greater the precision implied for the measurement.

1.6 Uncertainty in Measurement

- To determine the number of significant figures in a reported measurement, read the number from left to right, counting the digits starting with the first digit that is not zero.
- In any measurement that is properly reported, all nonzero digits are significant.
- Because zeros can be used either as part of the measured value or merely to locate the decimal point, they may or may not be significant:
  - Zeros between nonzero digits are always significant:
  - **Examples**: 1005 kg (four significant figures); 7.03 cm (three significant figures).

1.6 Uncertainty in Measurement

- Zeros at the beginning of a number are never significant.
- **Examples**: 0.02 g (one significant figure); 0.0026 cm (two significant figures).
- Zeros at the end of a number are significant if the number contains a decimal point:
- **Examples**: 0.0200 g (three significant figures); 3.0 cm (two significant figures).

❖ A problem arises when a number ends with zeros but contains no decimal point.

❖ In such cases, it is normally assumed that the zeros are not significant.

1.6 Uncertainty in Measurement

- Exponential notation can be used to indicate whether end zeros are significant.

1.03×10^4 g	(three significant figures)
1.030×10^4 g	(four significant figures)
1.0300×10^4 g	(five significant figures)

Sample Exercise 1.7

How many significant figures are in each of the following numbers (assume that each number is a measured quantity)?

a. 4.003 → 4 significant figures

b. 6.023×10^{23} → 4 significant figures

c. 5000 → 1 significant figure

1.7 Dimensional Analysis

- Because measured quantities have units associated with them, it is important to keep track of units as well as numerical values when using the quantities in calculations.
- In **dimensional analysis**, units are multiplied together or divided into each other along with the numerical values.

1.7 Dimensional Analysis

Conversion Factors

- A conversion factor is a fraction whose numerator and denominator are the same quantity expressed in different units.
- **Example**:

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

Desired unit

$$\text{Number of centimeters} = (8.50\ \cancel{\text{in.}})\ \frac{2.54\ \text{cm}}{1\ \cancel{\text{in.}}} = 21.6\ \text{cm}$$

Given unit

1.7 Dimensional Analysis

- Because the numerator and denominator of a conversion factor are equal, multiplying any quantity by a conversion factor is equivalent to multiplying by the number 1 and so does not change the intrinsic value of the quantity.
- In general, we begin any conversion by examining the units of the given data and the units we desire.
- We then ask ourselves what conversion factors we have available to take us from the units of the given quantity to those of the desired one.

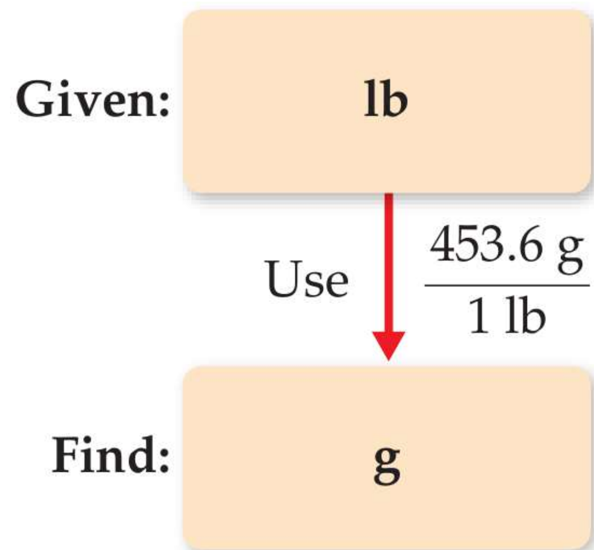
$$\cancel{\text{Given unit}} \times \frac{\text{desired unit}}{\cancel{\text{given unit}}} = \text{desired unit}$$

Sample Exercise 1.10

If a woman has a mass of 115 lb, what is her mass in grams?
(1 lb = 453.6 g)

$$\text{Mass} = 115\ \text{lb}\left(\frac{453.6\ \text{g}}{1\ \text{lb}}\right)$$

$$= 5.22 \times 10^{4}\ \text{g}$$



1.7 Dimensional Analysis

Using Two or More Conversion Factors

Given: **Find:**

$$\mathbf{m} \xrightarrow[\frac{1\ \text{cm}}{10^{-2}\ \text{m}}]{\text{Use}} \mathbf{cm} \xrightarrow[\frac{1\ \text{in.}}{2.54\ \text{cm}}]{\text{Use}} \mathbf{in.}$$

$$\text{Number of inches} = (8.00\ \cancel{\text{m}})\left(\frac{1\ \cancel{\text{cm}}}{10^{-2}\ \cancel{\text{m}}}\right)\left(\frac{1\ \text{in.}}{2.54\ \cancel{\text{cm}}}\right) = 315\ \text{in.}$$

Sample Exercise 1.11

The average speed of a nitrogen molecule in air at 25 °C is 515 m/s. Convert this speed to miles per hour. (1 mi = 1.6093 km)

Given: **Find:**

$$\text{m/s} \xrightarrow[\frac{1\ \text{km}}{10^3\ \text{m}}]{\text{Use}} \text{km/s} \xrightarrow[\frac{1\ \text{mi}}{1.6093\ \text{km}}]{\text{Use}} \text{mi/s} \xrightarrow[\frac{60\ \text{s}}{1\ \text{min}}]{\text{Use}} \text{mi/min} \xrightarrow[\frac{60\ \text{min}}{1\ \text{hr}}]{\text{Use}} \text{mi/hr}$$

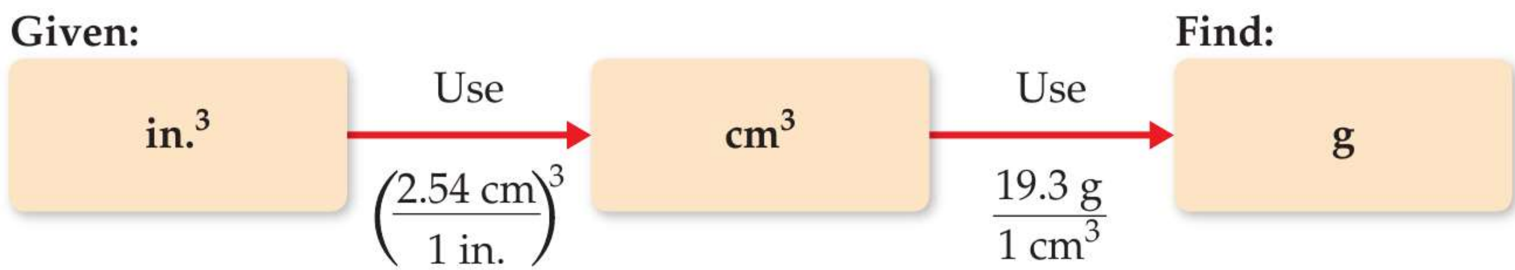
$$\text{Speed} = 515\,\frac{\text{m}}{\text{s}}\left(\frac{1\ \text{km}}{10^3\ \text{m}}\right)\left(\frac{1\ \text{mi}}{1.6093\ \text{km}}\right)\left(\frac{60\ \text{s}}{1\ \text{min}}\right)\left(\frac{60\ \text{min}}{1\ \text{h}}\right)$$

$$= 1.15 \times 10^3\ \text{mi/h}$$

1.7 Dimensional Analysis

Conversions Involving Volume

- Suppose we want to know the mass in grams of 2.00 cubic inches of gold, which has a density of 19.3 g/cm^3. (1 in. = 2.54 cm)



$$\text{Mass} = 2.00\ \text{in.}^3 \left(\frac{2.54\ \text{cm}}{1\ \text{in.}}\right)^3 \left(\frac{19.3\ \text{g}}{1\ \text{cm}^3}\right)$$

$$= 2.00\ \text{in.}^3 \left(\frac{2.54^3\ \text{cm}^3}{1\ \text{in.}^3}\right)\left(\frac{19.3\ \text{g}}{1\ \text{cm}^3}\right) = 633\ \text{g}$$

Sample Exercise 1.12

Earth's oceans contain approximately 1.36×10^9 km^3 of water. Calculate the volume in liters.

$$\text{Mass} = 1.36 \times 10^9\ \text{km}^3 \left(\frac{10^3\ \text{m}}{1\ \text{km}}\right)^3 \left(\frac{1\ \text{dm}}{10^{-1}\ \text{m}}\right)^3 \left(\frac{1\ \text{L}}{1\ \text{dm}^3}\right)$$

$$= 1.36 \times 10^9\ \text{km}^3 \left(\frac{10^9\ \text{m}^3}{1\ \text{km}^3}\right) \left(\frac{1\ \text{dm}^3}{10^{-3}\ \text{m}^3}\right) \left(\frac{1\ \text{L}}{1\ \text{dm}^3}\right)$$

$$= 1.36 \times 10^{21}\ \text{L}$$

Sample Exercise 1.13

What is the mass in grams of 1.00 gal of water? The density of water is 1.00 g/mL.

$$\text{Mass} = 1.00\ \text{gal}\left(\frac{4\ \text{qt}}{1\ \text{gal}}\right)\left(\frac{1\ \text{L}}{1.057\ \text{qt}}\right)\left(\frac{1\ \text{mL}}{10^{-3}\ \text{L}}\right)\left(\frac{1.00\ \text{g}}{1\ \text{mL}}\right)$$

$$= 3.78 \times 10^{3}\ \text{g}$$