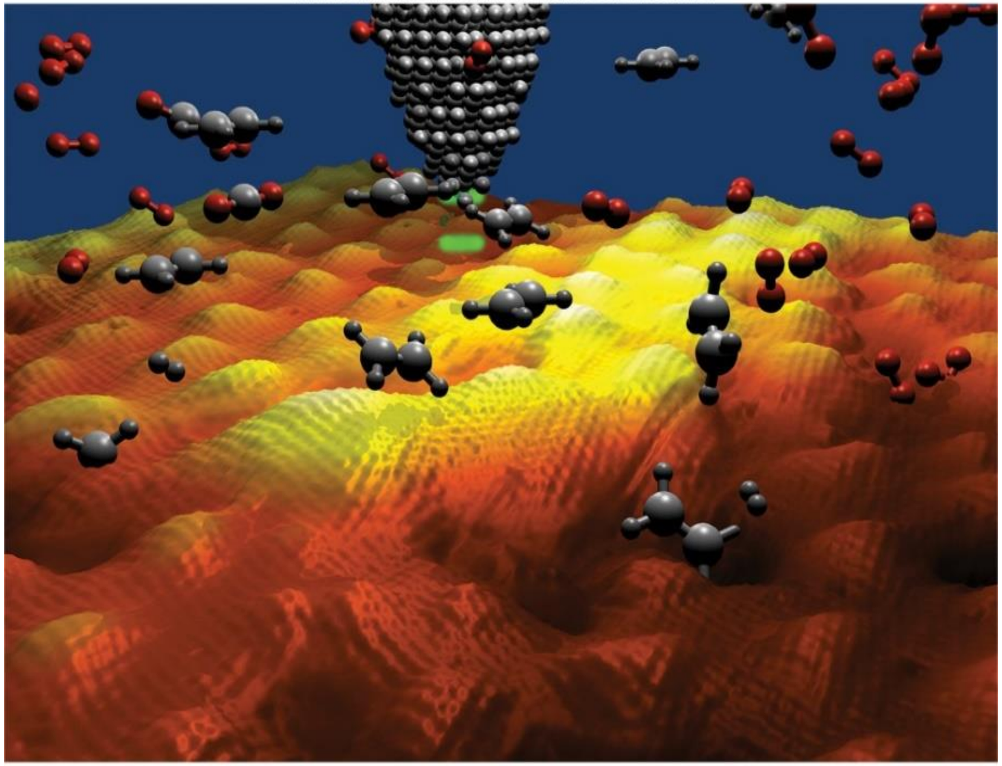
Chapter 1

Chemistry:
The Study of Change



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Defining Chemistry

Chemistry is the study of matter and the changes it undergoes.

Matter is anything that occupies space and has mass.

A ***substance*** is a form of matter that has a definite composition and distinct properties.

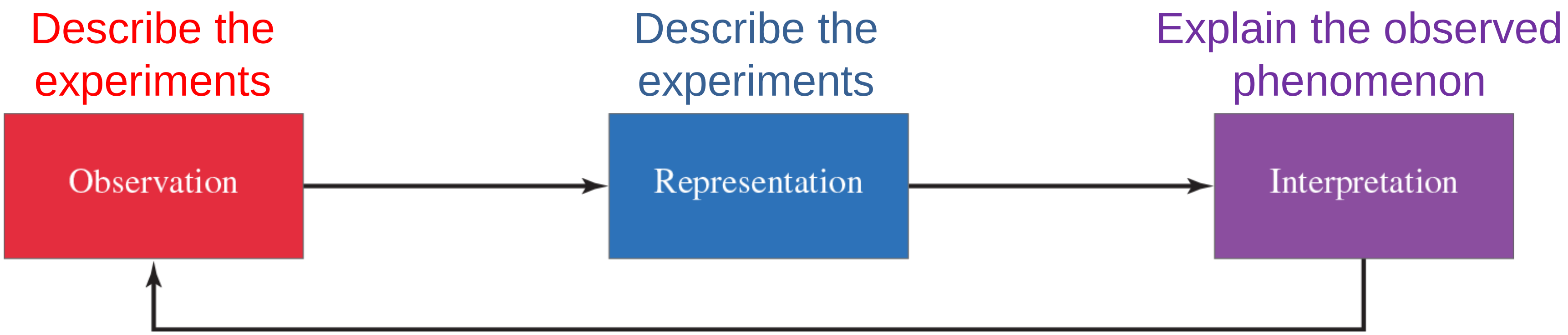
liquid nitrogen

gold ingots

silicon crystals

The Scientific Method

The ***scientific method*** is a systematic approach to research.



A ***hypothesis*** is a tentative explanation for a set of observations.

This hypothesis is then tested by developing and carrying out experiments. Based on the results of these experiments, the hypothesis is revised to account for experimental evidence. Further observations must be consistent with the hypothesis - otherwise the hypothesis must be revised.

The Scientific Method

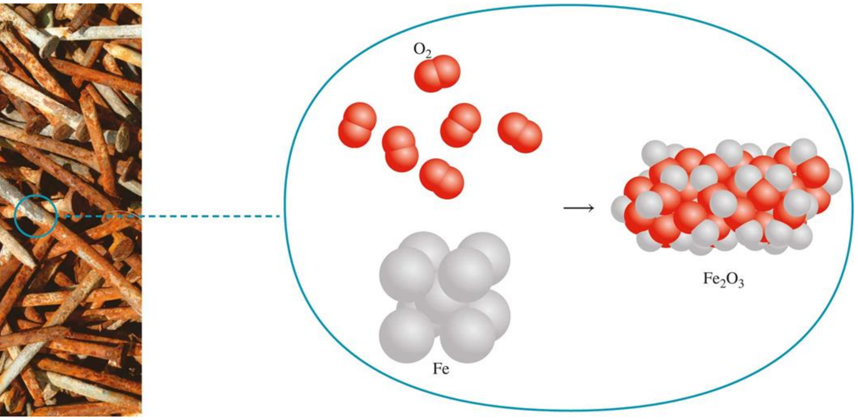
A ***hypothesis*** is a tentative explanation for a set of observations.

A ***law*** is a concise statement of a relationship between phenomena that is always the same under the same conditions.

$$\text{Force} = \text{mass} \times \text{acceleration}$$

A ***theory*** is a unifying principle that explains a body of facts and/or those laws that are based on them.

Atomic **Theory**



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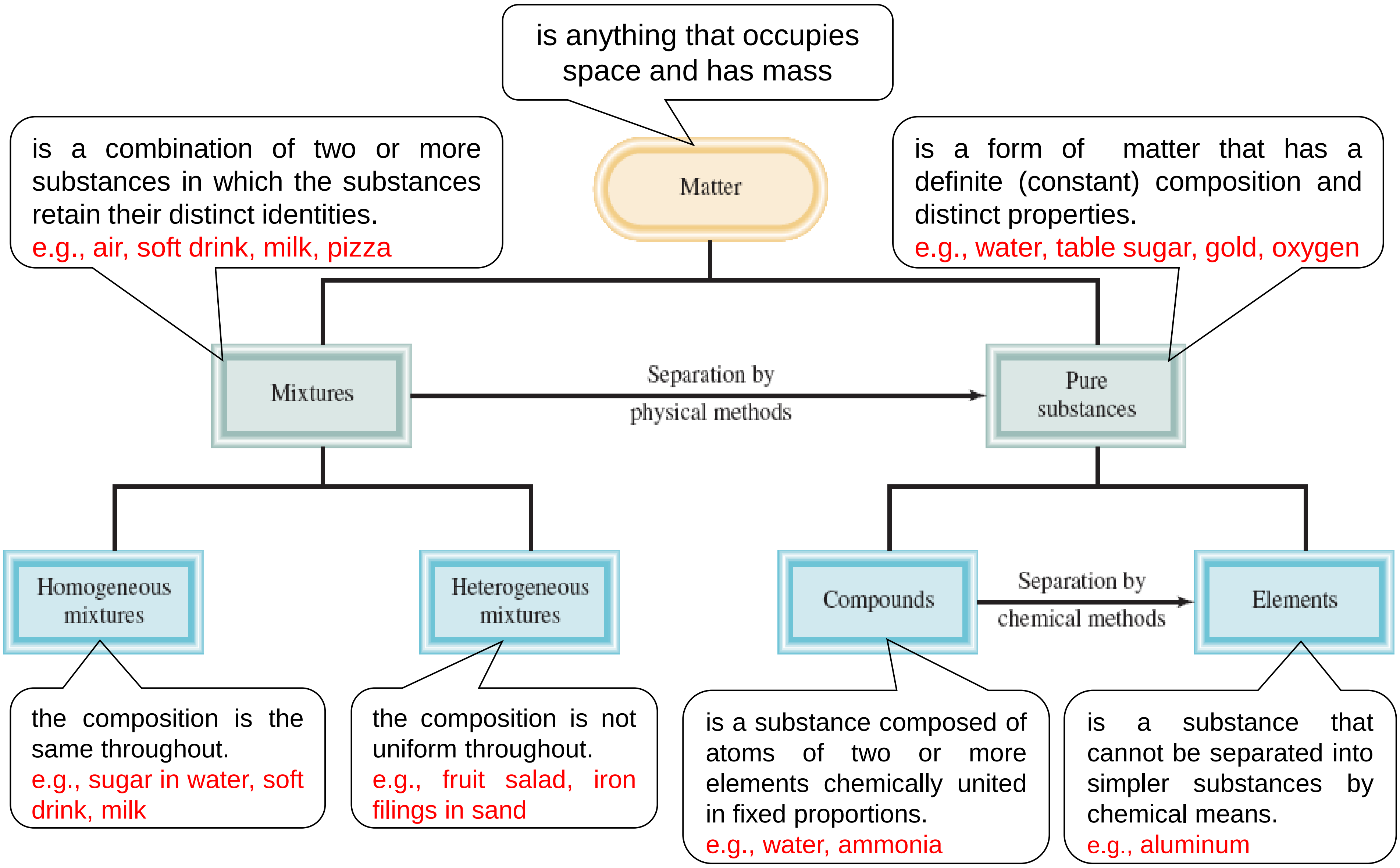
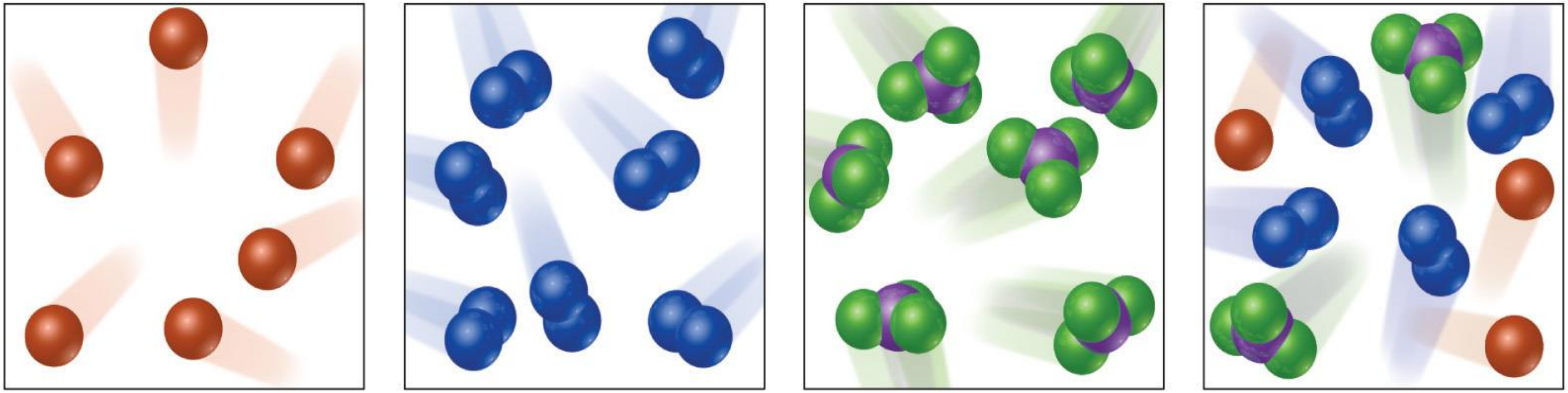
Classification of matter



Classification of matter according to its composition



(a) Atoms of an element

(b) Molecules of an element

(c) Molecules of a compound

(d) Mixture of elements and a compound

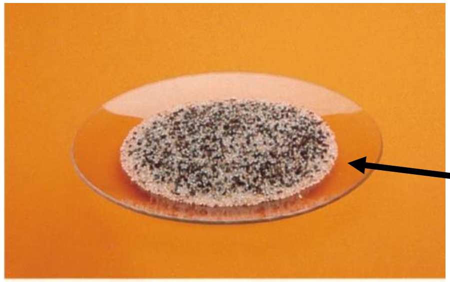
Mixtures

A ***mixture*** is a combination of two or more substances in which the substances retain their distinct identities.

1. ***Homogenous mixture*** – composition of the mixture is the same throughout.

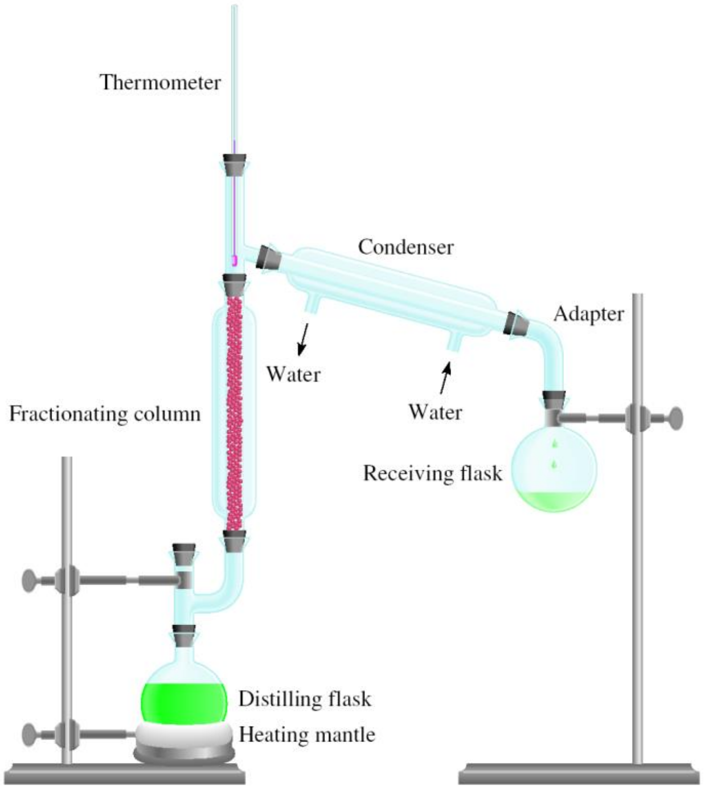
soft drink, milk, solder

2. ***Heterogeneous mixture*** – composition is not uniform throughout.



cement, iron filings in sand

Physical means can be used to separate a mixture into its pure components.



distillation

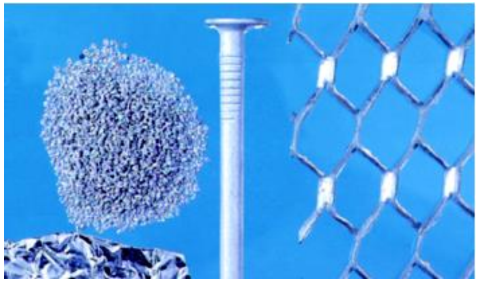
magnet

Elements

An ***element*** is a substance that <u>cannot</u> be separated into simpler substances by ***chemical means***.

- 118 elements have been identified
  - 82 elements occur naturally on Earth

    gold, aluminum, lead, oxygen, carbon, sulfur



  - 36 elements have been created by scientists

    technetium, americium, seaborgium

For convenience, chemists use symbols of one or two letters to represent the elements. The first letter of a symbol is *always* capitalized, but any following letters are not. e.g.,

- **Co** is the symbol for the element cobalt,
- **CO** is the formula for the carbon monoxide molecule.

TABLE 1.1 Some Common Elements and Their Symbols

Name	Symbol	Name	Symbol	Name	Symbol
Aluminum	Al	Fluorine	F	Oxygen	O
Arsenic	As	Gold	Au	Phosphorus	P
Barium	Ba	Hydrogen	H	Platinum	Pt
Bismuth	Bi	Iodine	I	Potassium	K
Bromine	Br	Iron	Fe	Silicon	Si
Calcium	Ca	Lead	Pb	Silver	Ag
Carbon	C	Magnesium	Mg	Sodium	Na
Chlorine	Cl	Manganese	Mn	Sulfur	S
Chromium	Cr	Mercury	Hg	Tin	Sn
Cobalt	Co	Nickel	Ni	Tungsten	W
Copper	Cu	Nitrogen	N	Zinc	Zn

The symbols of some elements are derived from their Latin names, e.g., **Au** from *aurum* (gold), **Ag** from *argentum* (silver), **Fe** from *ferrum* (iron), and **Na** from *natrium* (sodium), whereas most of them come from their English names.

Compounds

A ***compound*** is a substance composed of atoms of two or more elements chemically united in fixed proportions.

Compounds can only be separated into their pure components (elements) by ***chemical* means**.

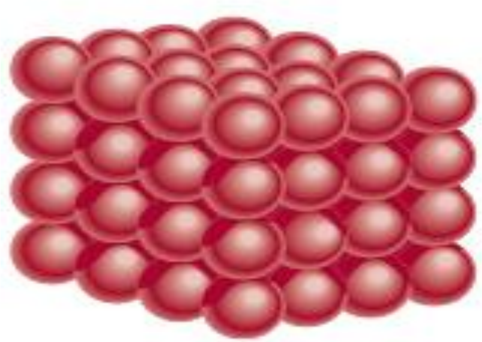
lithium fluoride

quartz

dry ice
(carbon dioxide)

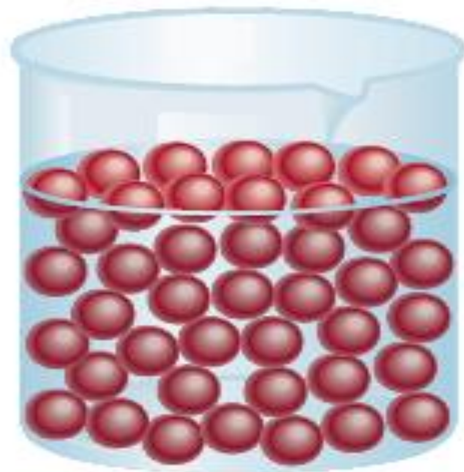
Classification of matter according to its physical state

All substances, at least in principle, can exist in three states: **solid**, **liquid** and **gas**.



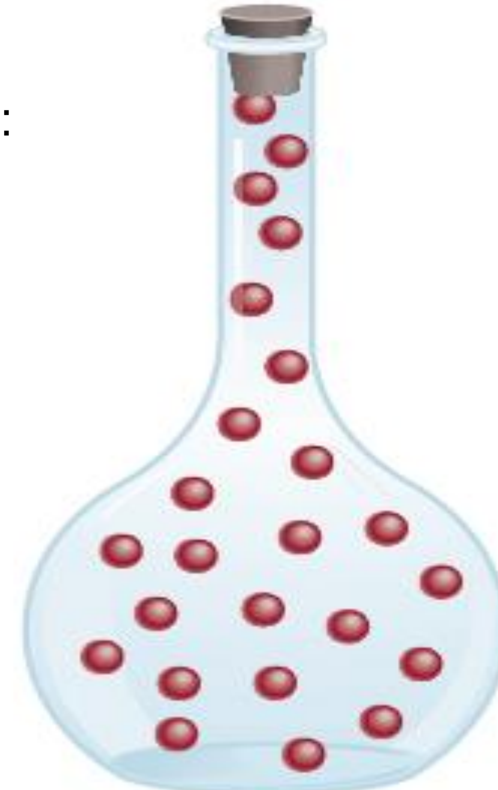
Solid

Molecules are held close together in orderly fashion with little freedom of motion. Has both definite shape and volume.



Liquid

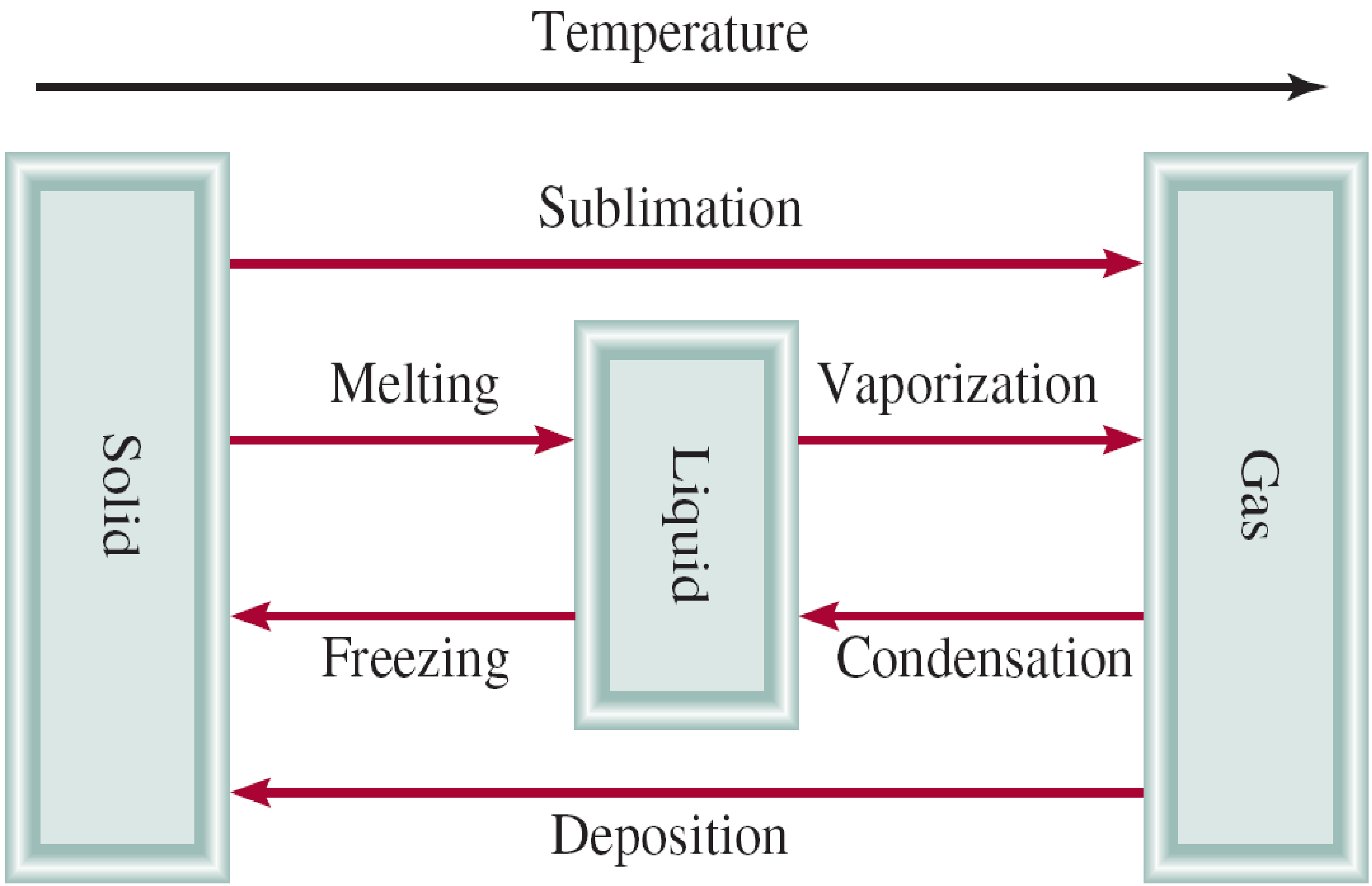
Molecules are close together but not held so rigidly in position and can move past on another. Has distinct volume but has no specific shape.



Gas

Molecules are separated by distance. Has no fixed volume or shape, they conform to the volume and shape of its containers. It can be compressed or expand.

The three states of matter can be interconverted without changing the **composition** of the substance.



Water has the same composition and chemical structure in the three states of matter (ice, water and vapor).

International System of Units (SI)

The General Conference of Weights and Measures; the international authority on units, proposed a revised metric system called the **International System of Units** (abbreviated **SI,** from the French Système Internationale d'Unites, 1960).

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

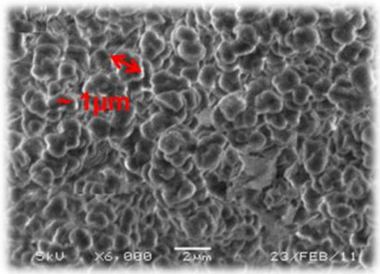
- All other units can be derived from these base units.

Prefixes Used with SI Units

1 µm

0.000001 m

1.0×10^{-6} m



150 000 000 000 m

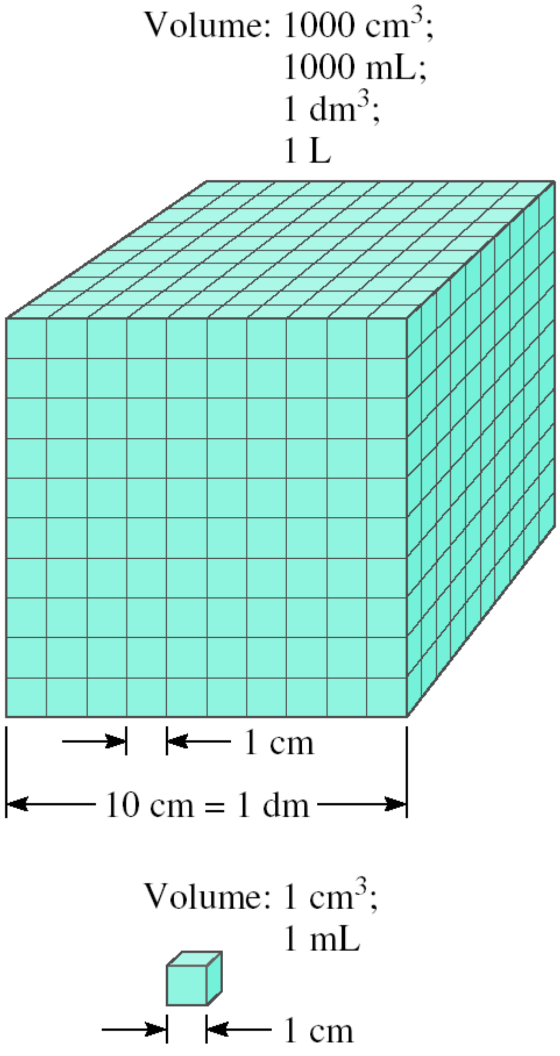
1.5×10^{11} m

Prefix	Symbol	Meaning	Example
peta-	P	1,000,000,000,000,000, or 10^{15}	1 petameter (Pm) = 1×10^{15} m
tera-	T	1,000,000,000,000, or 10^{12}	1 terameter (Tm) = 1×10^{12} m
giga	G	1,000,000,000, or 10^{9}	1 gigameter (Gm) = 1×10^{9} m
mega-	M	1,000,000, or 10^{6}	1 megameter (Mm) = 1×10^{6} m
kilo-	K	1,000, or 10^{3}	1 kilometer (km) = 1×10^{3} m
deci-	d	1/10, or 10^{-1}	1 decimeter (dm) = 0.1 m
centi-	c	1/100, or 10^{-2}	1 centimeter (cm) = 0.01 m
milli-	m	1/1,000, or 10^{-3}	1 millimeter (mm) = 0.001 m
micro-	µ	1/1,000,000, or 10^{-6}	1 micrometer (µm) = 1×10^{-6} m
nano-	n	1/1,000,000,000, or 10^{-9}	1 nanometer (nm) = 1×10^{-9} m
pico-	p	1/1,000,000,000,000, or 10^{-12}	1 picometer (pm) = 1×10^{-12} m
femto-	f	1/1,000,000,000,000,000, or 10^{-15}	1 femtometer (fm) = 1×10^{-15} m
atto	a	1/1,000,000,000,000,000,000, or 10^{-18}	1 attometer (am) = 1×10^{-18} m

Prefixes are used to indicate decimal **multiples** or **fractions** of various units. Prefixes convert the base units into units that are appropriate for the item being measured.

Volume

Volume – SI derived unit for volume is cubic meter (m^3)



$$1\ cm^3 = (1 \times 10^{-2}\ m)^3 = 1 \times 10^{-6}\ m^3$$

$$1\ dm^3 = (1 \times 10^{-1}\ m)^3 = 1 \times 10^{-3}\ m^3$$

$$1\ L = 1000\ mL = 1000\ cm^3 = 1\ dm^3$$

$$1\ mL = 1\ cm^3$$

$1\ m^3 = 1000\ dm^3$

$1\ L = 1\ dm^3$

$1\ dm^3 = 1000\ cm^3$

$1\ cm^3 = 1\ mL$

Density

The **density** of a substance is its mass per unit volume (the volumetric mass).

SI derived unit for density is **kg/m³**

1 g/cm³ = 1 g/mL = 1000 kg/m³

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

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

TABLE 1.4

Densities of Some Substances at 25°C

Substance	**Density (g/cm³)**
Air*	0.001
Ethanol	0.79
Water	1.00
Mercury	13.6
Table salt	2.2
Iron	7.9
Gold	19.3
Osmium†	22.6

*Measured at 1 atmosphere.
†Osmium (Os) is the densest element known.

A piece of platinum metal with a density of 21.5 g/cm^3 has a volume of 4.49 cm^3. What is its mass?

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

$$m = d \times V$$

$$= 21.5\ \text{g/}\cancel{\text{cm}^3} \times 4.49\ \cancel{\text{cm}^3} = \underline{\mathbf{96.5\ g}}$$

EXAMPLE 1.1

Gold is a precious metal that is chemically unreactive. It is used mainly in jewelry, dentistry, and electronic devices. A piece of gold ingot with a mass of 301 g has a volume of 15.6 cm^3. Calculate the density of gold.

Solution We are given the mass and volume and asked to calculate the density. Therefore, from Equation (1.1), we write

$$
\begin{aligned}
d &= \frac{m}{V} \\
&= \frac{301 \text{ g}}{15.6 \text{ cm}^3} \\
&= 19.3 \text{ g/cm}^3
\end{aligned}
$$

EXAMPLE 1.2

The density of mercury, the only metal that is a liquid at room temperature, is 13.6 g/mL. Calculate the mass of 5.50 mL of the liquid.

Solution We are given the density and volume of a liquid and asked to calculate the mass of the liquid. We rearrange Equation (1.1) to give

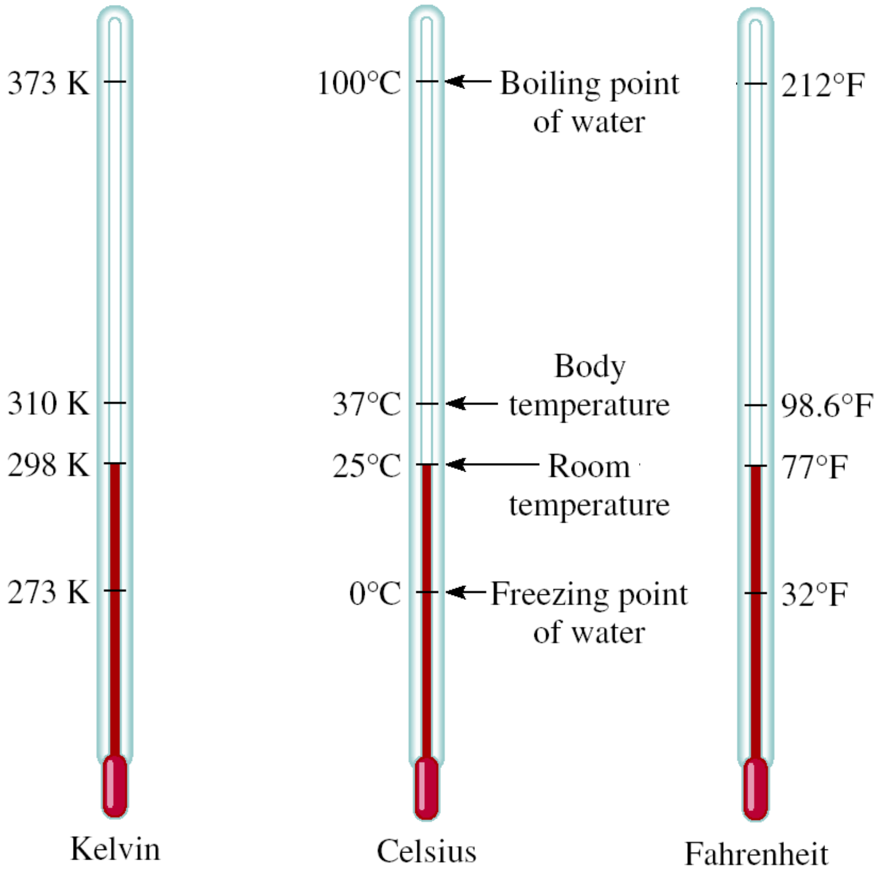
$$m = d \times V$$

$$= 13.6 \frac{\text{g}}{\cancel{\text{mL}}} \times 5.50\ \cancel{\text{mL}}$$

$$= 74.8\ \text{g}$$

Practice Exercise The density of sulfuric acid in a certain car battery is 1.41 g/mL. Calculate the mass of 242 mL of the liquid.

A Comparison of Temperature Scales



$$K = {}^{o}C + 273.15$$

$$273\ K = 0\ {}^{o}C$$
$$373\ K = 100\ {}^{o}C$$

$${}^{o}F = \frac{9}{5} \times {}^{o}C + 32$$

$$32\ {}^{o}F = 0\ {}^{o}C$$
$$212\ {}^{o}F = 100\ {}^{o}C$$

Convert 172.9 ^oF to degrees Celsius.

$$^oF = \frac{9}{5} \times {}^oC + 32$$

$$^oF - 32 = \frac{9}{5} \times {}^oC$$

$$\frac{5}{9} \times ({}^oF - 32) = {}^oC$$

$$^oC = \frac{5}{9} \times ({}^oF - 32)$$

$$^oC = \frac{5}{9} \times (172.9 - 32) = \underline{\mathbf{78.3}}$$

EXAMPLE 1.3

(a) Solder is an alloy made of tin and lead that is used in electronic circuits. A certain solder has a melting point of 224°C. What is its melting point in degrees Fahrenheit? (b) Helium has the lowest boiling point of all the elements at −452°F. Convert this temperature to degrees Celsius. (c) Mercury, the only metal that exists as a liquid at room temperature, melts at −38.9°C. Convert its melting point to kelvins.

Solution These three parts require that we carry out temperature conversions, so we need Equations (1.2), (1.3), and (1.4). Keep in mind that the lowest temperature on the Kelvin scale is zero (0 K); therefore, it can never be negative.

(a) This conversion is carried out by writing

$$\frac{9°\text{F}}{5°\text{C}} \times (224°\text{C}) + 32°\text{F} = 435°\text{F}$$

Scientific Notation

Chemists often deal with numbers that are either extremely large or extremely small.

Very Very Very Very Very Very Very Very Very Very Very Very Very Very

- The number of atoms in 12 g of carbon:

602,200,000,000,000,000,000,000

$$6.022 \times 10^{23}$$

- The mass of a single carbon atom in grams:

0.0000000000000000000000199

$$1.99 \times 10^{-23}$$

$$N \times 10^{n}$$

N is a number between 1 and 10

n is a positive or negative integer

Scientific Notation

568.762

$\leftarrow$ move decimal left

n > 0

$$568.762 = 5.68762 \times 10^{2}$$

0.00000772

$\rightarrow$ move decimal right

n < 0

$$0.00000772 = 7.72 \times 10^{-6}$$

- <u>Addition or Subtraction</u>

1. Write each quantity with the same exponent ***n***
2. Combine N_1 and N_2
3. The exponent, ***n***, remains the same

$$4.31 \times 10^{4} + 3.9 \times 10^{3} =$$

$$4.31 \times 10^{4} + 0.39 \times 10^{4} =$$

$$4.70 \times 10^{4}$$

Scientific Notation

Multiplication

1. Multiply N_1 and N_2
2. Add exponents n_1 and n_2

$$\left(4.0\times10^{-5}\right)\times\left(7.0\times10^{3}\right)=$$

$$\left(4.0\times7.0\right)\times\left(10^{-5+3}\right)=$$

$$28\times10^{-2}=$$

$$2.8\times10^{-1}$$

Division

1. Divide N_1 and N_2
2. Subtract exponents n_1 and n_2

$$8.5\times10^{4}\div5.0\times10^{9}=$$

$$\left(8.5\div5.0\right)\times10^{4-9}=$$

$$1.7\times10^{-5}$$

(1) Express 568.762 in scientific notation:

$568.762 = 5.68762 \times 10^{2}$

Note that the decimal point is moved to the left by two places and $n = 2$

(2) Express 0.00000772 in scientific notation:

$0.00000772 = 7.72 \times 10^{-6}$

Here the decimal point is moved to the right by six places and $n = -6$

Significant Figures

- Any digit that is not zero is significant
  1.234 kg **4** significant figures
- Zeros between nonzero digits are significant
  606 m **3** significant figures
- Zeros to the left of the first nonzero digit are **not** significant
  0.08 L **1** significant figure
- If a number is greater than 1, then all zeros to the right of the decimal point are significant
  2.0 mg **2** significant figures
- If a number is less than 1, then only the zeros that are at the end and in the middle of the number are significant
  0.00420 g **3** significant figures

EXAMPLE 1.4

Determine the number of significant figures in the following measurements: (a) 478 cm, (b) 6.01 g, (c) 0.825 m, (d) 0.043 kg, (e) 1.310×10^{22} atoms, (f) 7000 mL.

Solution (a) Three, because each digit is a nonzero digit. (b) Three, because zeros between nonzero digits are significant. (c) Three, because zeros to the left of the first nonzero digit do not count as significant figures. (d) Two. Same reason as in (c). (e) Four, because the number is greater than one so all the zeros written to the right of the decimal point count as significant figures. (f) This is an ambiguous case. The number of significant figures may be four (7.000×10^3), three (7.00×10^3), two (7.0×10^3), or one (7×10^3). This example illustrates why scientific notation must be used to show the proper number of significant figures.

Significant Figures

- Addition or Subtraction
- The answer cannot have more digits to the right of the decimal point than any of the original numbers.

$$
\begin{array}{r}
89.332 \\
\underline{+1.1} \\
90.432
\end{array}
$$

$+1.1$ ← one significant figure after decimal point

90.432 ← roundoff to 90.4

$$
\begin{array}{r}
3.70 \\
\underline{-2.9133} \\
0.7867
\end{array}
$$

-2.9133 ← two significant figure after decimal point

0.7867 ← roundoff to 0.79

Significant Figures

- <u>Multiplication or Division</u>
- The number of significant figures in the result is set by the original number that has the ***smallest*** number of significant figures.

$$4.51 \times 3.6666 = 16.536366 = 16.5$$

↑ 3 sig figs — ↑ round to 3 sig figs

$$6.8 \div 112.04 = 0.0606926 = 0.061$$

↑ 2 sig figs — ↑ round to 2 sig figs

Significant Figures

Exact Numbers

Numbers from definitions or numbers of objects are considered to have an infinite number of significant figures.

The average of three measured lengths: 6.64, 6.68 and 6.70?

$$\frac{6.64 + 6.68 + 6.70}{3} = 6.67333 = 6.67$$

Because 3 is an **exact number**

EXAMPLE 1.5

Carry out the following arithmetic operations to the correct number of significant figures: (a) 11,254.1 g + 0.1983 g, (b) 66.59 L − 3.113 L, (c) 8.16 m × 5.1355, (d) 0.0154 kg ÷ 88.3 mL, (e) 2.64×10^3 cm + 3.27×10^2 cm.

Solution In addition and subtraction, the number of decimal places in the answer is determined by the number having the lowest number of decimal places. In multiplication and division, the significant number of the answer is determined by the number having the smallest number of significant figures.

(a)
$$\begin{array}{r} 11{,}254.1\ \text{g} \\ +\quad 0.1983\ \text{g} \\ \hline 11{,}254.2983\ \text{g} \end{array} \longleftarrow \text{round off to } 11{,}254.3\ \text{g}$$

(b)
$$\begin{array}{r} 66.59\ \text{L} \\ -\ \ 3.113\ \text{L} \\ \hline 63.477\ \text{L} \end{array} \longleftarrow \text{round off to } 63.48\ \text{L}$$

(c) $8.16\ \text{m} \times 5.1355 = 41.90568\ \text{m} \longleftarrow$ round off to 41.9 m

(d) $\dfrac{0.0154\ \text{kg}}{88.3\ \text{mL}} = 0.000174405436\ \text{kg/mL} \longleftarrow$ round off to 0.000174 kg/mL or 1.74×10^{-4} kg/mL

(e) First we change 3.27×10^2 cm to 0.327×10^3 cm and then carry out the addition $(2.64\ \text{cm} + 0.327\ \text{cm}) \times 10^3$. Following the procedure in (a), we find the answer is 2.97×10^3 cm.

Dimensional Analysis Method of Solving Problems

1. Determine which unit conversion factor(s) are needed
2. Carry units through calculation
3. If all units cancel except for the ***desired unit(s)***, then the problem was solved correctly.

$$\text{given quantity} \times \text{conversion factor} = \text{desired quantity}$$

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

EXAMPLE 1.6

A person's average daily intake of glucose (a form of sugar) is 0.0833 pound (lb). What is this mass in milligrams (mg)? (1 lb = 453.6 g.)

Strategy The problem can be stated as

$$? \text{ mg} = 0.0833 \text{ lb}$$

The relationship between pounds and grams is given in the problem. This relationship will enable conversion from pounds to grams. A metric conversion is then needed to convert grams to milligrams ($1 \text{ mg} = 1 \times 10^{-3} \text{ g}$). Arrange the appropriate conversion factors so that pounds and grams cancel and the unit milligrams is obtained in your answer.

Solution The sequence of conversions is

$$\text{pounds} \longrightarrow \text{grams} \longrightarrow \text{milligrams}$$

Using the following conversion factors

$$\frac{453.6 \text{ g}}{1 \text{ lb}} \quad \text{and} \quad \frac{1 \text{ mg}}{1 \times 10^{-3} \text{ g}}$$

we obtain the answer in one step:

$$? \text{ mg} = 0.0833 \cancel{\text{lb}} \times \frac{453.6 \cancel{\text{g}}}{1 \cancel{\text{lb}}} \times \frac{1 \text{ mg}}{1 \times 10^{-3} \cancel{\text{g}}} = 3.78 \times 10^{4} \text{ mg}$$

EXAMPLE 1.7

An average adult has 5.2 L of blood. What is the volume of blood in m^3?

Strategy The problem can be stated as

$$? \text{ m}^3 = 5.2 \text{ L}$$

How many conversion factors are needed for this problem? Recall that $1 \text{ L} = 1000 \text{ cm}^3$ and $1 \text{ cm} = 1 \times 10^{-2} \text{ m}$.

Solution We need two conversion factors here: one to convert liters to cm^3 and one to convert centimeters to meters:

$$\frac{1000 \text{ cm}^3}{1 \text{ L}} \quad \text{and} \quad \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}}$$

Because the second conversion factor deals with length (cm and m) and we want volume here, it must therefore be cubed to give

$$\frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} \times \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} \times \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} = \left(\frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}}\right)^3$$

This means that $1 \text{ cm}^3 = 1 \times 10^{-6} \text{ m}^3$. Now we can write

$$? \text{ m}^3 = 5.2 \cancel{\text{L}} \times \frac{1000 \text{ cm}^3}{1 \cancel{\text{L}}} \times \left(\frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}}\right)^3 = 5.2 \times 10^{-3} \text{ m}^3$$