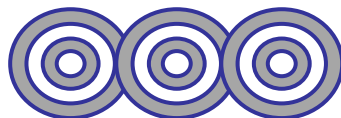




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Chapter 1

Chemistry: The Study of Change

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1.4

Classifications of Matter

Classification of matter

is anything that occupies space and has mass

Matter

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

e.g., air, soft drink, milk, pizza

is a form of matter that has a definite (constant) composition and distinct properties.

e.g., water, table sugar, gold, oxygen

Mixtures

Separation by
physical methods

Pure
substances

Homogeneous
mixtures

Heterogeneous
mixtures

Compounds

Separation by
chemical methods

Elements

the composition is the same throughout.

e.g., sugar in water, soft drink, milk

the composition is not uniform throughout.

e.g., fruit salad, iron filings in sand

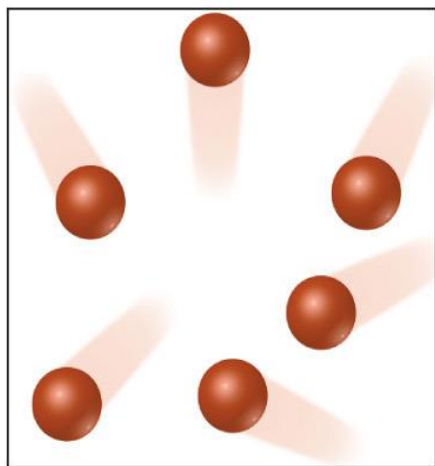
is a substance composed of atoms of two or more elements chemically united in fixed proportions.

e.g., water, ammonia

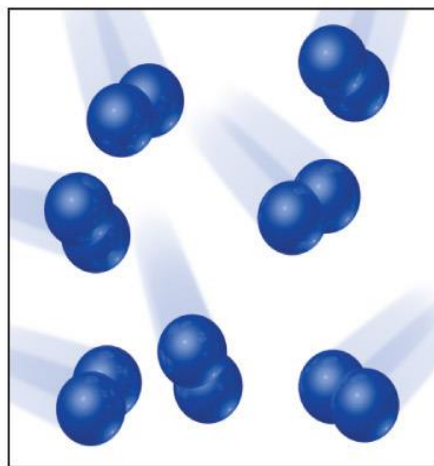
is a substance that cannot be separated into simpler substances by chemical means.

e.g., aluminum

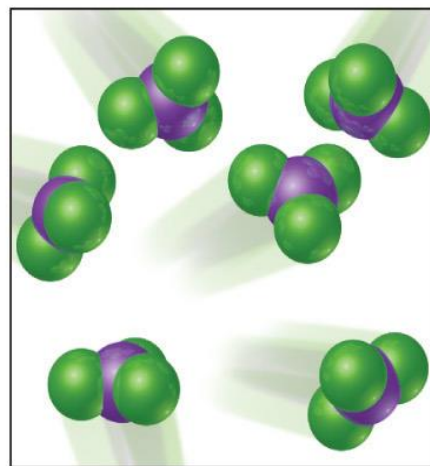
Classification of matter according to its composition



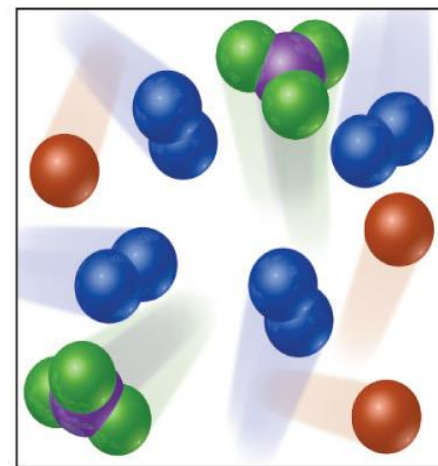
(a) Atoms of an element



(b) Molecules of an element



(c) Molecules of a compound



(d) Mixture of elements and a compound

How to write chemical symbols for elements

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.

Some common elements and their symbols

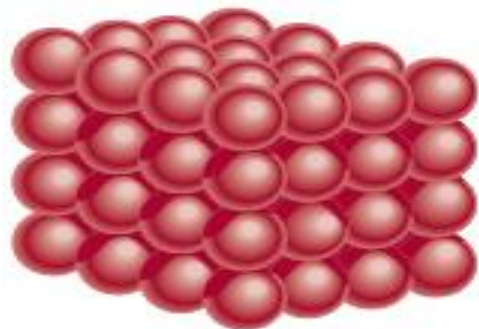
Name	Symbol	Name	Symbol	Name	Symbol
Aluminum	Al	Fluorine	F	Oxygen	O
Arsenic	As	Gold	Au	Phosphorus	P
Barium	Ba	Hydrogen	H	Platinum	Pt
Bismuth	Bi	Iodine	I	Potassium	K
Bromine	Br	Iron	Fe	Silicon	Si
Calcium	Ca	Lead	Pb	Silver	Ag
Carbon	C	Magnesium	Mg	Sodium	Na
Chlorine	Cl	Manganese	Mn	Sulfur	S
Chromium	Cr	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.

1.5

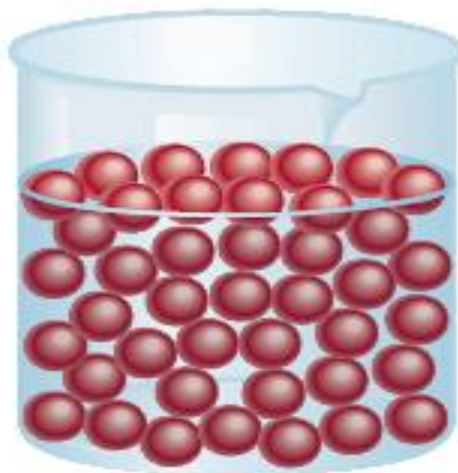
The Three States of Matter

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



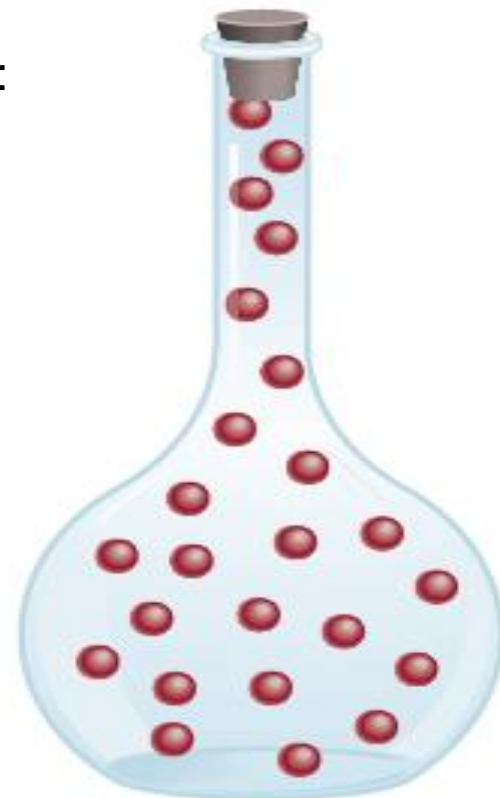
Solid

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



Liquid

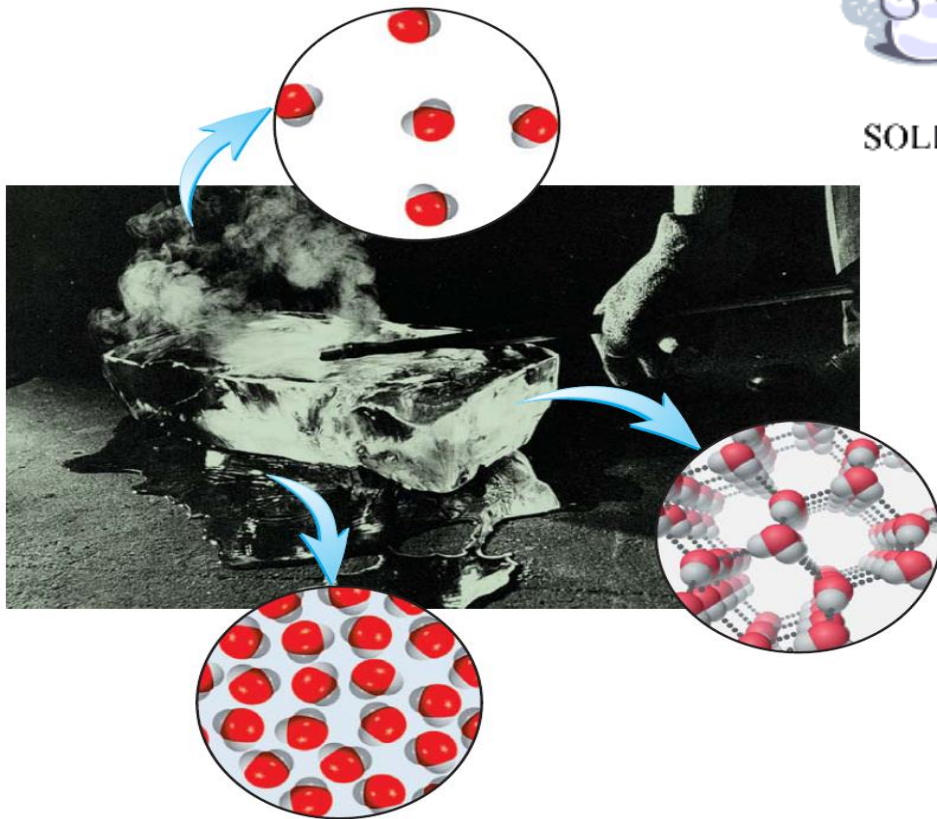
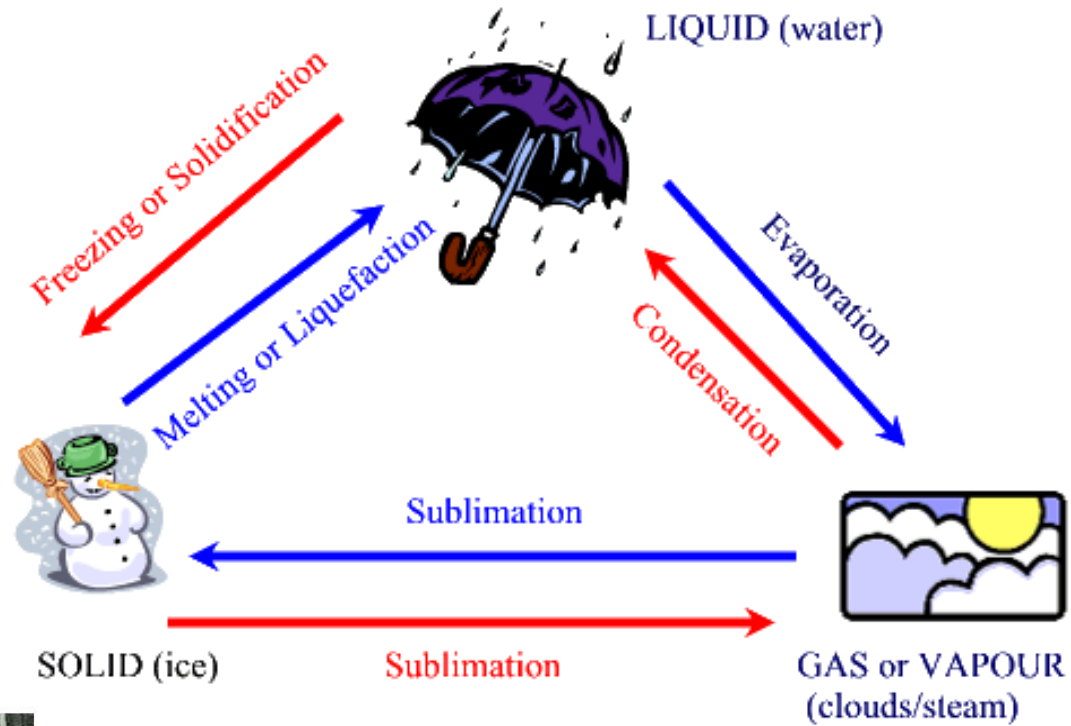
Molecules are close together but not held so rigidly in position and can move past on another. Has distinct volume but has no specific shape.



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).

1.6

Physical and Chemical Properties of Matter

Physical & Chemical properties

Substances are identified by their **properties** as well as by their **composition**.

A chemical property and a physical property are related to chemical and physical changes of matter.

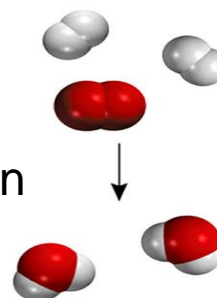
A **physical property** is an aspect of matter that can be observed or measured without changing the composition or identity of a substance.

e.g., color, volume, density, molecular weight, melting point & boiling point.

A **chemical property** may only be observed by changing the chemical composition or identity of a substance.

e.g., reactivity, flammability, pH & oxidation states.

Hydrogen burns in
air to form water



Extensive & Intensive properties

All measurable properties of matter fall into one of two additional categories;

extensive properties and **intensive** properties.

An **extensive property** of a material depends upon how much matter is being considered. The value of an extensive quantity depends on the amount of matter. Extensive properties are additive.

e.g., mass, length, volume, surface area, force.

An **intensive property** of a material does not depend upon how much matter is being considered. Intensive properties are not additive.

e.g., density, temperature, color, pressure.



Sulfur crystals

EXTENSIVE
(dependent on amount of substance)

39 g	mass	0.84 g
18.8 cm ³	volume	4.1 cm ³

INTENSIVE
(independent of amount of substance)

yellow	color	yellow
115.2°C	melting point	115.2°C



Sulfur powder

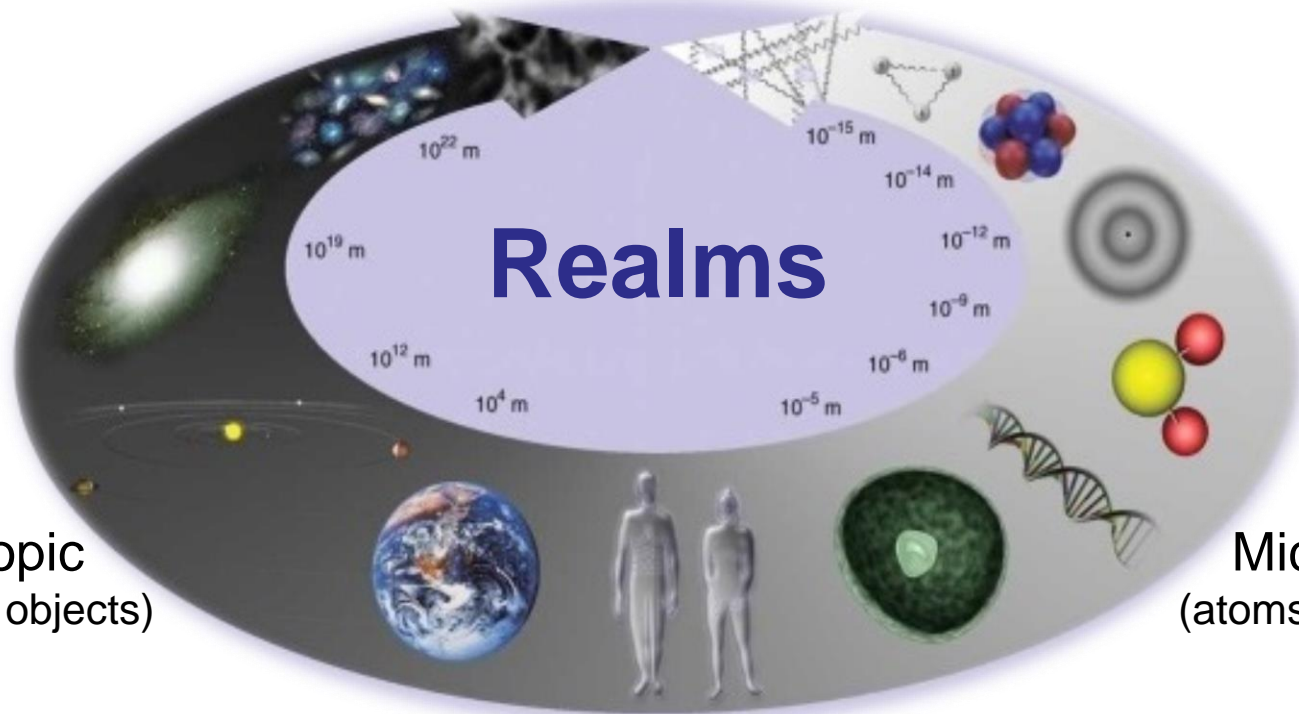


100 g	Mass of water	10 g
100 mL	Volume of water	10 mL
25 °C	Temperature of water	25 °C
1.0 g/mL	Density of water	1.0 g/mL

1.7

Measurement

Units of Measurement



Macroscopic
(ordinary sized objects)

Microscopic
(atoms & molecules)

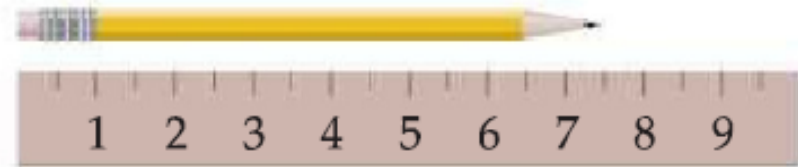
Macroscopic properties, on the ordinary scale, which can be determined directly. e.g., length, mass, volume, temperature.

Microscopic properties, on the atomic or molecular scale, must be determined by an indirect method.

Units

A measured quantity is usually written as a number with an appropriate unit.

7.5 meaningless
7.5 cm specifies length



Units are essential to stating measurements correctly.

The units used for scientific measurements are those of the **metric units**.

The metric system is an internationally agreed decimal system of measurement that was originally based on the mètre des Archives and the kilogramme des Archives introduced by France in 1799.

SI Units

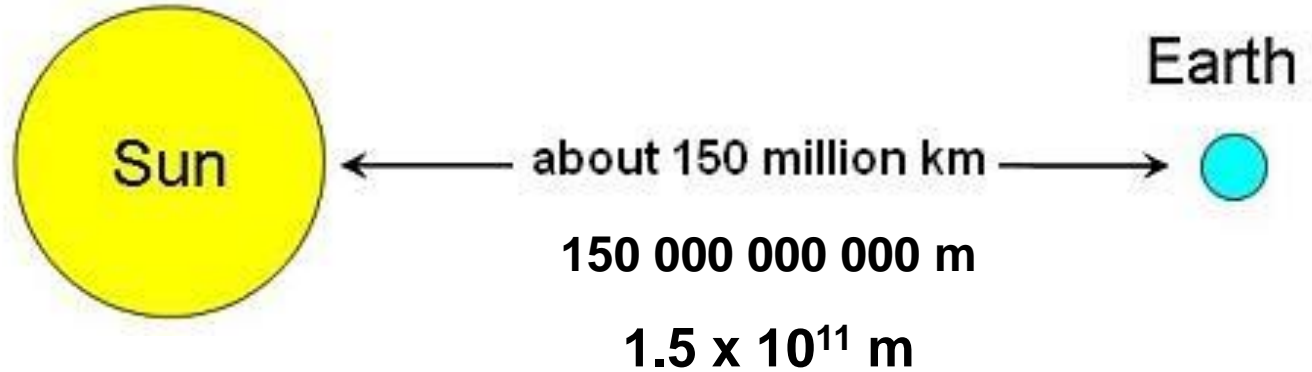
The General Conference of Weights and Measures; the international authority on units, proposed a revised metric system called the **International System of Units** (abbreviated **SI**, from the French *Système Internationale d'Unites*, 1960).

SI Base Units (seven base units)

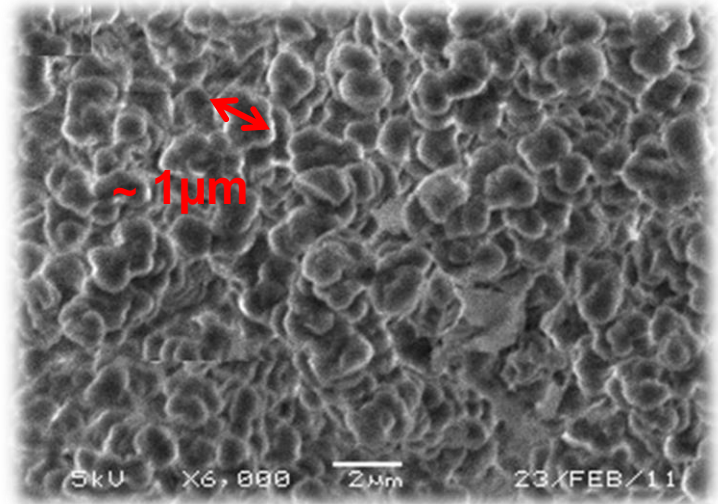
Base Quantity	Name of Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electrical current	ampere	A
Temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

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

Prefixes



1 μm
0.000001 m
1.0 x 10⁻⁶ m



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

Prefix	Symbol	Meaning	Example
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

Note that a metric prefix simply represents a number:

$$1 \text{ mm} = 1 \times 10^{-3} \text{ m}$$

Mass & Weight

Mass: is a measure of the amount of matter in an object
(SI unit of mass is the kilogram, kg).

Weight: is the force that gravity exerts on an object
(SI unit is Newton, N).

In chemistry, the smaller gram (g) is more convenient:

$$1 \text{ kg} = 1000 \text{ g} = 1 \times 10^3 \text{ g}$$

Temperature

Temperature: is a measure of the hotness or coldness of an object (SI unit K).

The Celsius and Kelvin scales are most often used in scientific measurements.

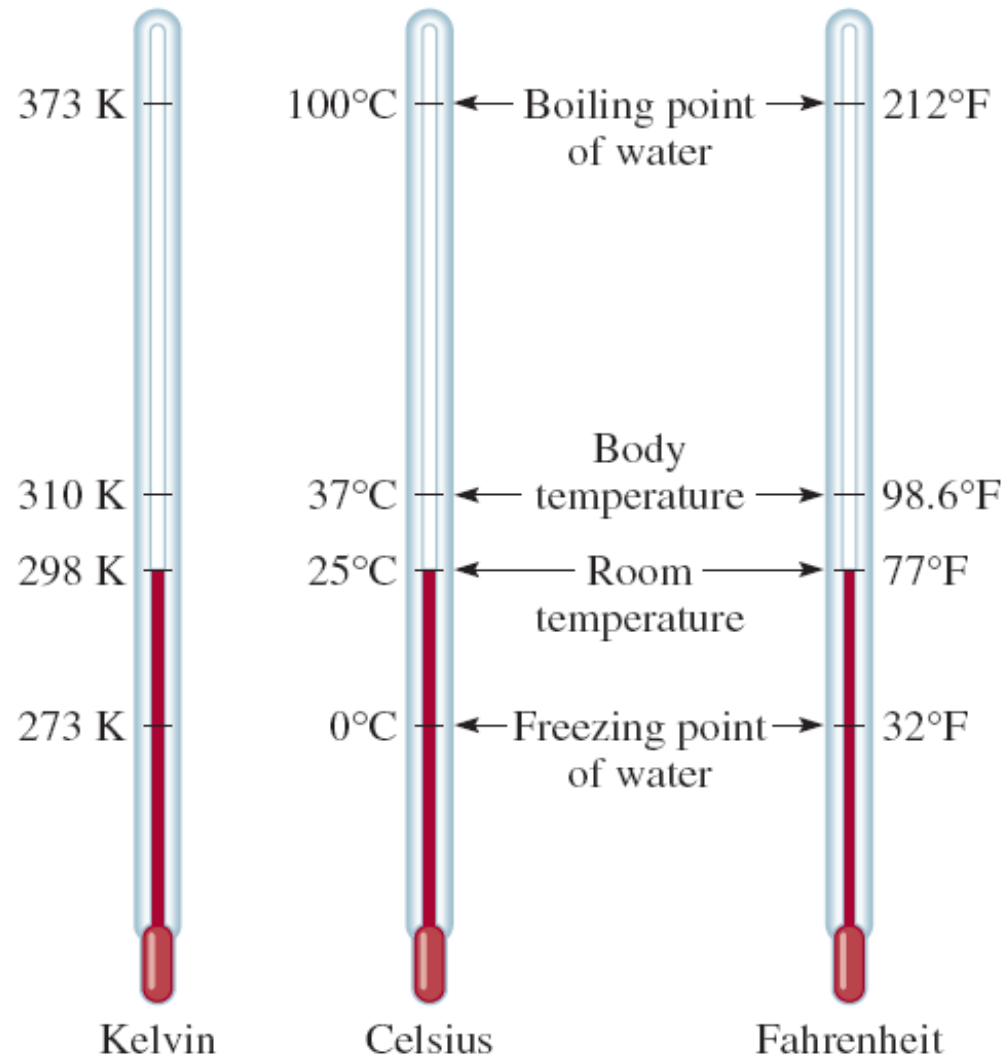
The Kelvin is based on the properties of gases.

There are no -ve Kelvin temperatures.

At $-273.15\text{ }^{\circ}\text{C}$ (0 K) called absolute zero.

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

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



EXAMPLE

(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?

Solution:

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

(b) Helium has the lowest boiling point of all the elements at -452°F . Convert this temperature to degrees Celsius.

Solution:

$$(-452^{\circ}\text{F} - 32^{\circ}\text{F}) \times \frac{5^{\circ}\text{C}}{9^{\circ}\text{F}} = -269^{\circ}\text{C}$$

(c) Mercury, the only metal that exists as a liquid at room temperature. Convert its melting point to kelvins, if it melts at -38.9°C .

Solution:

$$(-38.9^{\circ}\text{C} + 273.15^{\circ}\text{C}) \times \frac{1\text{ K}}{1^{\circ}\text{C}} = 234.3\text{ K}$$

Practice Exercise

Convert (a) 327.5°C (the melting point of lead) to degrees Fahrenheit; (b) 172.9°F (the boiling point of ethanol) to degrees Celsius; and (c) 77 K , the boiling point of liquid nitrogen, to degrees Celsius.

Derived SI Units

The SI are used to derive the units of other quantities.

For example:

Speed is defined as the ratio of distance traveled to elapsed time.

$$\text{Speed} = \frac{\textit{length}}{\textit{time}} = \frac{m}{s}$$

Thus, the SI unit for speed is meters per second (m/s).

Volume

SI-derived unit for volume is the *cubic meter* (m^3).

Generally, however, chemists work with much smaller volumes, such as the cubic centimeter (cm^3) and the cubic decimeter (dm^3):

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

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

$$1 \text{ m}^3 = 1000 \text{ dm}^3$$

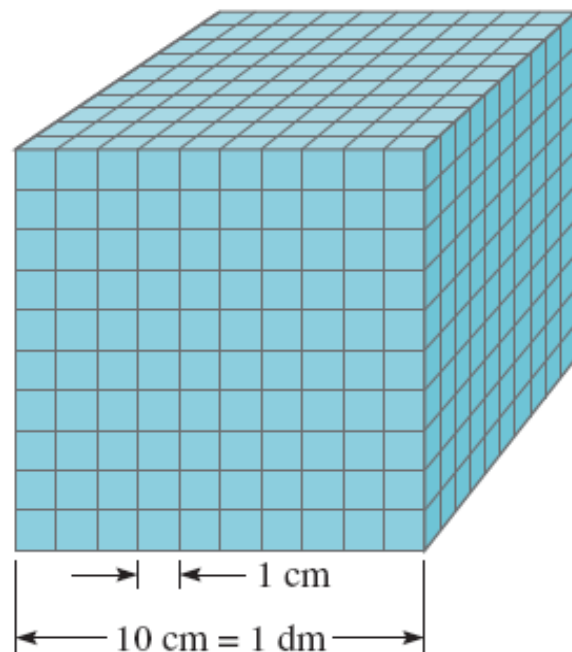
$$1 \text{ L} = 1 \text{ dm}^3$$

$$1 \text{ dm}^3 = 1000 \text{ cm}^3$$

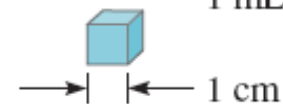
$$1 \text{ cm}^3 = 1 \text{ mL}$$

Even though the liter is not an SI unit, volumes are usually expressed in liters (L) and milliliters (mL).

Volume: 1000 cm^3 ;
 1000 mL ;
 1 dm^3 ;
 1 L



Volume: 1 cm^3 ;
 1 mL



Density

Density is the amount of mass in a unit volume of the substance.

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

The SI-derived unit for density is the kilogram per cubic meter (kg/m^3).

This unit is awkwardly large for most chemical applications. Therefore, grams per cubic centimeter (g/cm^3) and its equivalent, grams per milliliter (g/mL), are more commonly used for solid and liquid densities.

$$1 \text{ g/cm}^3 = 1 \text{ g/mL} = 1000 \text{ kg/m}^3$$

Because gas densities are often very low, we express them in units of grams per liter (g/L):

$$1 \text{ g/L} = 0.001 \text{ g/mL}$$

-The density of water is 1.00 g/mL (mass equal volume).

-Densities are temperature dependent (because most substances change volume when they are heated or cooled). So, the temperature should be specified.

-Density usually decreases with temperature.

-The density and weight are sometimes confused.

For example; iron has density more than air, but 1kg of air has the same mass as 1kg of iron, but iron occupies a smaller volume, which giving it a higher density.

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

EXAMPLE

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 cm³. Calculate the density of gold.

Solution:

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

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

Practice Exercise

A piece of platinum metal with a density of 21.5 g/cm³ has a volume of 4.49 cm³. What is its mass?

EXAMPLE

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:

$$m = d \times V$$

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

$$= 74.8 \text{ g}$$

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.

Other Derived SI Units

Distance = L = m (SI unit)

Area = L x L = m x m = m²

Volume = L x L x L = m x m x m = m³

Force = m a = kg (m/s²) = kg m s⁻² = N

Energy = $\frac{1}{2} m v^2 = \text{kg (m/s)}^2 = \text{kg m}^2 \text{ s}^{-2} = \text{J}$

Pressure = F / A = kg m s⁻² / m² = kg m⁻¹ s⁻² = Pa

1.8

Handling Numbers

Scientific Notation

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

For example, in **1 g** of the element hydrogen there are roughly **602,200,000,000,000,000,000 (6.022x10²³) hydrogen atoms.**

Each hydrogen atom has a mass of only **0.000000000000000000000000166 g (1.66x10⁻²⁴ g)**

Consequently, when working with very large and very small numbers, we use a system called **scientific notation.**

Regardless of their magnitude, all numbers can be expressed in the form

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

where **N** is a number between 1 and 10 and n, the exponent, is a positive or negative integer (whole number).

EXAMPLES

(1) Express 568.762 in scientific notation:

$$568.762 = 5.68762 \times 10^2 \quad \text{move decimal to left (n>0, +ve)}$$

(2) Express 0.00000772 in scientific notation:

$$0.00000772 = 7.72 \times 10^{-6} \quad \text{move decimal to right (n<0, -ve)}$$

Addition & Subtraction

$$(7.4 \times 10^3) + (2.1 \times 10^3) = 9.5 \times 10^3$$

$$(4.31 \times 10^4) + (3.9 \times 10^3) = (4.31 \times 10^4) + (0.39 \times 10^4) = 4.70 \times 10^4$$

$$(2.22 \times 10^{-2}) - (4.10 \times 10^{-3}) = (2.22 \times 10^{-2}) - (0.41 \times 10^{-2}) = 1.81 \times 10^{-2}$$

Multiplication & Division

$$(8.0 \times 10^4) \times (5.0 \times 10^2) = (8.0 \times 5.0) (10^{4+2}) = 40 \times 10^6 = 4.0 \times 10^7$$

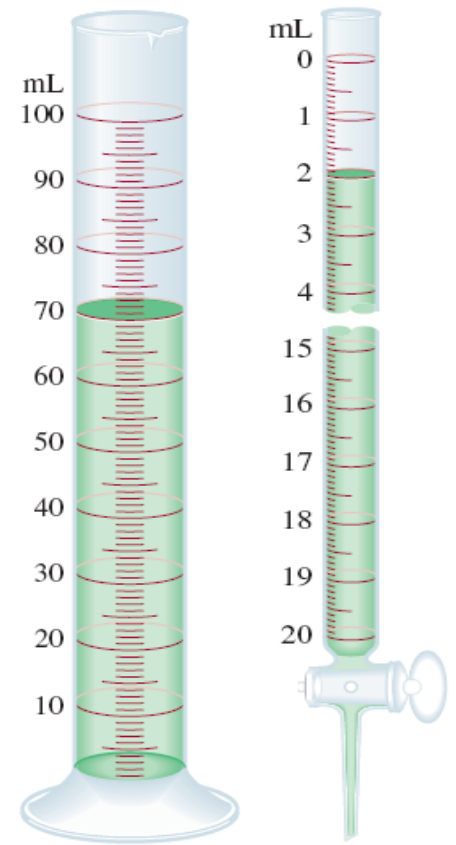
$$\frac{6.9 \times 10^7}{3.0 \times 10^{-5}} = \frac{6.9}{3.0} \times 10^{7-(-5)} = 2.3 \times 10^{12}$$

Significant Figures

In scientific measurements, it is important to indicate the margin of error in a measurement by clearly indicating the number of significant figures, which are the meaningful digits in a measured or calculated quantity.

Measuring the volume using a graduated cylinder with a scale that gives an uncertainty of **1 mL**. If the volume is found to be 6 mL, then the actual volume is in the range of 5 to 7 mL. Represent as (6 ± 1) mL.

For greater accuracy, we might use a graduated cylinder or a buret with **0.1 mL** uncertainty. If the volume of the liquid is 6.0 mL, we may express the quantity as (6.0 ± 0.1) mL, and the actual value is somewhere between 5.9 and 6.1 mL.



When significant figures are used, the last digit is understood to be uncertain.

Guidelines for using significant figures

1. Any digit that is not zero is significant.

e.g., 845 has 3 sig. fig., 1.234 has 4 sig. fig.

2. Zeros between nonzero digits are significant.

e.g., 606 contains 3 sig. fig., 40,501 contains 5 sig. fig.

3. Zeros to the left of the first nonzero digit are not significant.

e.g., 0.08 contains 1 sig. fig., 0.0000349 contains 3 sig. fig.

4. If a number is greater than 1, then all the zeros written to the right of the decimal point count as significant figures.

e.g., 2.0 has 2 sig. fig., 40.062 has 5 sig. fig., 3.040 has 4 sig. fig.

If a number is less than 1, then only the zeros that are at the end of the number and the zeros that are between nonzero digits are significant.

e.g., 0.090 has 2 sig. fig., 0.3005 has 4 sig. fig., 0.00420 has 3 sig. fig.

5. For numbers that do not contain decimal points, the trailing zeros (that is, zeros after the last nonzero digit) may or may not be significant.

e.g., 400 may have 1 sig. fig., 2 sig. fig., or 3 sig. fig. (**ambiguous case**)

We cannot know without using scientific notation. We can express the number 400 as 4×10^2 for 1 sig. fig., 4.0×10^2 for 2 sig. fig., or 4.00×10^2 for 3 sig. fig.

EXAMPLE

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.

(a) 478 cm, 3 sig. fig.

(b) 6.01 g, 3 sig. fig.

(c) 0.825 m, 3 sig. fig.

(d) 0.043 kg, 2 sig. fig.

(e) 1.310×10^{22} atoms, 4 sig. fig.

(f) 7000 mL, This is an ambiguous case; the number of significant figures may be: 4 for (7.000×10^3), 3 for (7.00×10^3), 2 for (7.0×10^3), 1 for (7×10^3).

Practice Exercise

Determine the number of significant figures in each of the following measurements:

(a) 24 mL, (b) 3001 g, (c) 0.0320 m³, (d) 6.4×10^4 molecules, (e) 560 kg.

Significant figures in calculations

1. In addition and subtraction, the answer cannot have more digits to the right of the decimal point than either of the original numbers.

Examples:

$$\begin{array}{r} 89.332 \\ + 1.1 \\ \hline 90.432 \end{array} \leftarrow \text{one digit after the decimal point} \quad \left/ \quad \begin{array}{r} 2.097 \\ - 0.12 \\ \hline 1.977 \end{array} \leftarrow \text{two digits after the decimal point}$$

$90.432 \leftarrow \text{round off to } 90.4$ $1.977 \leftarrow \text{round off to } 1.98$

If the first digit following the point of rounding off is equal to or greater than 5, we add 1 to the preceding digit.

e.g., 8.727 rounds off to 8.73, and 0.425 rounds off to 0.43.

2. In multiplication and division, the number of significant figures in the final product or quotient is determined by the original number that has the *smallest* number of significant figures.

Examples:

$$2.8 \times 4.5039 = 12.61092 \leftarrow \text{round off to } 13 \quad \left/ \quad \frac{6.85}{112.04} = 0.0611388789 \leftarrow \text{round off to } 0.0611$$

3. The *exact numbers* obtained from definitions, conversion factors or by counting numbers of objects can be considered to have an infinite number of significant figures.

Example (1)

The inch is defined to be exactly 2.54 centimeters; that is,

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

Thus, the “2.54” in the equation should not be interpreted as a measured number with 3 significant figures. In calculations involving conversion between “in” and “cm”, we treat both “1” & “2.54” as having an infinite number of significant figures.

Example (2)

If an object has a mass of 5.0 g, then the mass of nine such objects is,

$$5.0 \text{ g} \times 9 = 45 \text{ g}$$

The answer has two significant figures because 5.0 g has two significant figures.

The number 9 is exact and does not determine the number of significant figures.

Example (3)

The average of three measured lengths; 6.64, 6.68 and 6.70 is,

$$\frac{(6.64 + 6.68 + 6.70)}{3} = 6.67333 = 6.67 \text{ (two decimals because 3 is an exact number and not measured value).}$$

EXAMPLE

Carry out the following arithmetic operations to the correct number of significant figures:

(a) $11,254.1 \text{ g} + 0.1983 \text{ g}$,

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

(b) $66.59 \text{ L} - 3.113 \text{ L}$,

$$\begin{array}{r} 66.59 \text{ L} \\ - \quad 3.113 \text{ L} \\ \hline 63.477 \text{ L} \leftarrow \text{round off to } 63.48 \text{ L} \end{array}$$

(c) $8.16 \text{ m} \times 5.1355$,

$$8.16 \text{ m} \times 5.1355 = 41.90568 \text{ m} \leftarrow \text{round off to } 41.9 \text{ m}$$

(d) $0.0154 \text{ kg} \div 88.3 \text{ mL}$,

$$\frac{0.0154 \text{ kg}}{88.3 \text{ mL}} = 0.000174405436 \text{ kg/mL} \leftarrow \text{round off to } 0.000174 \text{ kg/mL} \text{ or } 1.74 \times 10^{-4} \text{ kg/mL}$$

(e) $2.64 \times 10^3 \text{ cm} + 3.27 \times 10^2 \text{ cm}$.

First we change $3.27 \times 10^2 \text{ cm}$ to $0.327 \times 10^3 \text{ cm}$

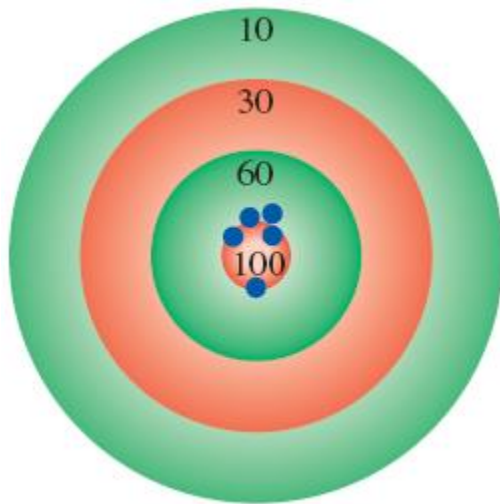
$$(2.64 \text{ cm} + 0.327 \text{ cm}) \times 10^3$$

$$2.97 \times 10^3 \text{ cm}$$

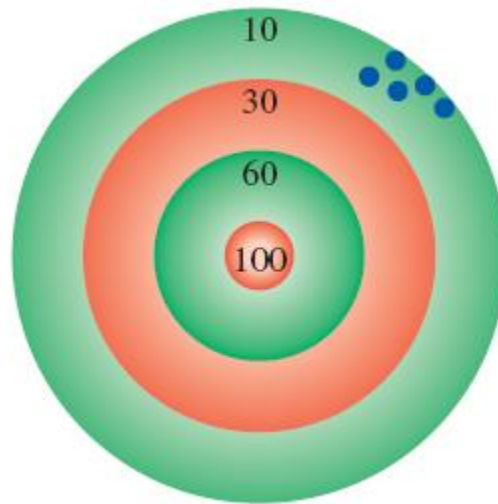
Accuracy & Precision

Accuracy tells us how close a measurement is to the true value of the quantity that was measured.

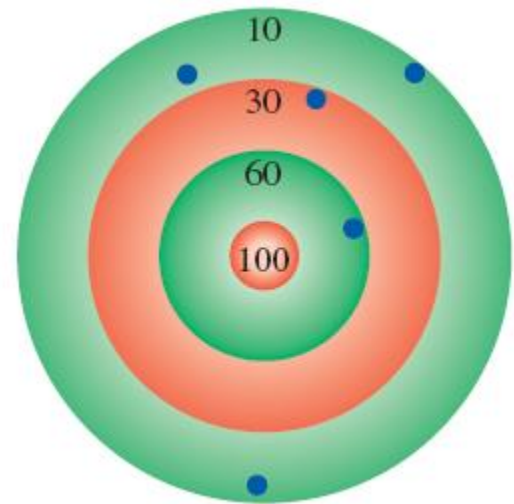
Precision refers to how closely two or more measurements of the same quantity agree with one another.



Good accuracy
&
Good precision



Poor accuracy
&
Good precision



Poor accuracy
&
Poor precision

Example,

Three students are asked to determine the mass of a piece of copper wire. The true mass of the wire is 2.000 g.

The results of two successive weighings by each student are:

	Student A	Student B	Student C
	1.964 g	1.972 g	2.000 g
	1.978 g	1.968 g	2.002 g
Average value	$\frac{1.971 \text{ g}}$	$\frac{1.970 \text{ g}}$	$\frac{2.001 \text{ g}}$

Therefore, **Student B's** results are more precise than those of **Student A**, but neither set of results is very accurate. **Student C's** results are not only the most precise, but also the most accurate, because the average value is closest to the true value.

Highly accurate measurements are usually precise too. On the other hand, **highly precise measurements do not necessarily guarantee accurate results.** For example, an improperly calibrated meterstick or a faulty balance may give precise readings that are in error.

1.9

Dimensional Analysis in Solving Problems

The procedure we use to convert between units in solving chemistry problems is called *dimensional analysis* (also called the *factor-label method*).

Dimensional analysis is based on the relationship between different units that express the same physical quantity.

For example, by definition 1 in = 2.54 cm

we can write the conversion factor as $\frac{1 \text{ in}}{2.54 \text{ cm}}$ or as $\frac{2.54 \text{ cm}}{1 \text{ in}}$

Convert 12.00 in to cm ?

$$12.00 \text{ in} \times \frac{2.54 \text{ cm}}{1 \text{ in}} = 30.48 \text{ cm}$$

In general, to apply dimensional analysis we use the relationship

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

and the units cancel as follows:

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

EXAMPLE

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

$$1 \text{ lb} = 453.6 \text{ g}$$

$$1 \text{ mg} = 10^{-3} \text{ g}$$

pounds \longrightarrow grams \longrightarrow milligrams

Conversion factors

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

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

Practice Exercise

A roll of aluminum foil has a mass of 1.07 kg. What is its mass in pounds?

EXAMPLE

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

$$1 \text{ L} = 1000 \text{ cm}^3$$

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

Conversion factors

$$\frac{1000 \text{ cm}^3}{1 \text{ L}} \quad \text{and} \quad \frac{1 \times 10^{-2} \text{ m}}{1 \text{ cm}} \quad \frac{1 \times 10^{-6} \text{ m}^3}{1 \text{ cm}^3}$$

$$? \text{ m}^3 = 5.2 \cancel{\text{ L}} \times \frac{1000 \cancel{\text{ cm}}^3}{1 \cancel{\text{ L}}} \times \frac{1 \times 10^{-6} \text{ m}^3}{1 \cancel{\text{ cm}}^3} = 5.2 \times 10^{-3} \text{ m}^3$$

Practice Exercise

The volume of a room is $1.08 \times 10^8 \text{ dm}^3$. What is the volume in m^3 ?

EXAMPLE

Liquid nitrogen is obtained from liquefied air and is used to prepare frozen goods and in low-temperature research. The density of the liquid at its boiling point (-196°C or 77 K) is 0.808 g/cm^3 . Convert the density to units of kg/m^3 .

$$1\text{ kg} = 1000\text{ g}\text{ and } 1\text{ cm} = 1 \times 10^{-2}\text{ m}\text{ or } 1\text{ cm}^3 = 1 \times 10^{-6}\text{ m}^3$$

Conversion factors

$$\frac{1\text{ kg}}{1000\text{ g}}\text{ and } \frac{1\text{ cm}^3}{1 \times 10^{-6}\text{ m}^3}$$

$$?\text{ kg/m}^3 = \frac{0.808\cancel{\text{ g}}}{1\cancel{\text{ cm}^3}} \times \frac{1\text{ kg}}{1000\cancel{\text{ g}}} \times \frac{1\cancel{\text{ cm}^3}}{1 \times 10^{-6}\text{ m}^3} = 808\text{ kg/m}^3$$

Practice Exercise

The density of the lightest metal, lithium (Li), is $5.34 \times 10^2\text{ kg/m}^3$. Convert the density to g/cm^3 .

EXAMPLE

The speed of sound in air is about 343 m/s. What is this speed in miles per hour?

Conversion factors:

1 mi = 1609 m, 1 min = 60 s, 1 hour = 60 min

$$343 \frac{\cancel{m}}{\cancel{s}} \times \frac{1 \text{ mi}}{1609 \cancel{m}} \times \frac{60 \cancel{s}}{1 \cancel{\text{min}}} \times \frac{60 \cancel{\text{min}}}{1 \text{ hour}} = 767 \frac{\text{mi}}{\text{hour}}$$

