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Chapter 1

Chemistry: The Study of Change

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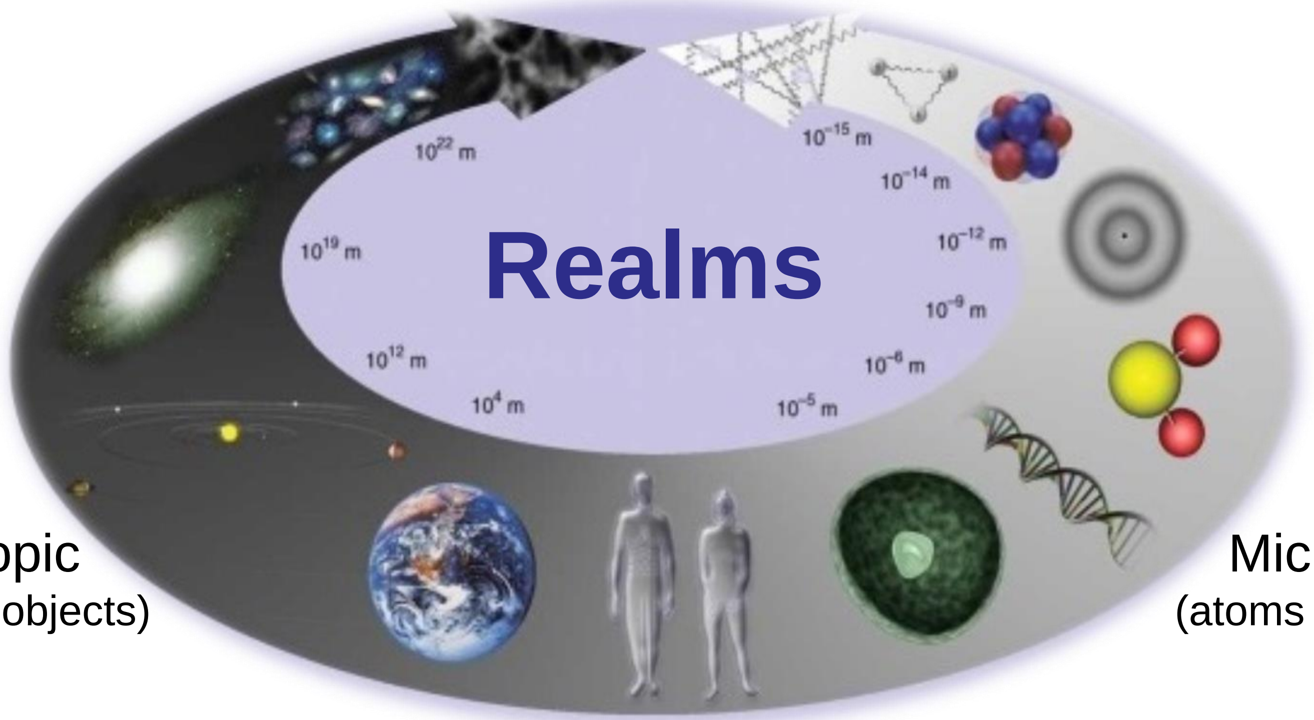
1.7
Measurement

Units of Measurement



Macroscopic
(ordinary sized objects)

Microscopic
(atoms & molecules)

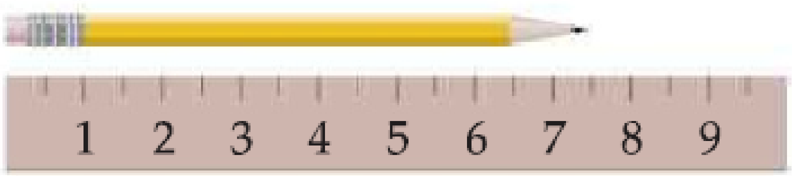
Macroscopic properties, on the ordinary scale, which can be determined directly. e.g., length, mass, volume, temperature.

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

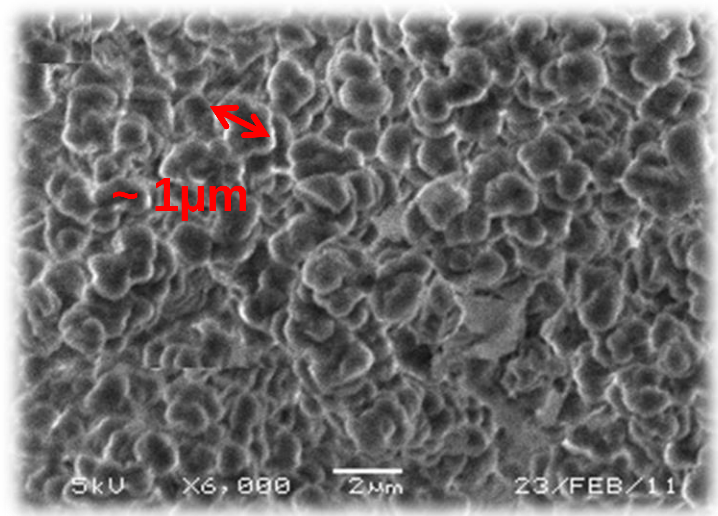
150 000 000 000 m

1.5×10^{11} m

1 μm

0.000001 m

1.0×10^{-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) $= 1 \times 10^{12}$ m
giga-	G	1,000,000,000, or 10^{9}	1 gigameter (Gm) $= 1 \times 10^{9}$ m
mega-	M	1,000,000, or 10^{6}	1 megameter (Mm) $= 1 \times 10^{6}$ m
kilo-	k	1,000, or 10^{3}	1 kilometer (km) $= 1 \times 10^{3}$ 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) $= 1 \times 10^{-6}$ m
nano-	n	1/1,000,000,000, or 10^{-9}	1 nanometer (nm) $= 1 \times 10^{-9}$ m
pico-	p	1/1,000,000,000,000, or 10^{-12}	1 picometer (pm) $= 1 \times 10^{-12}$ 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.

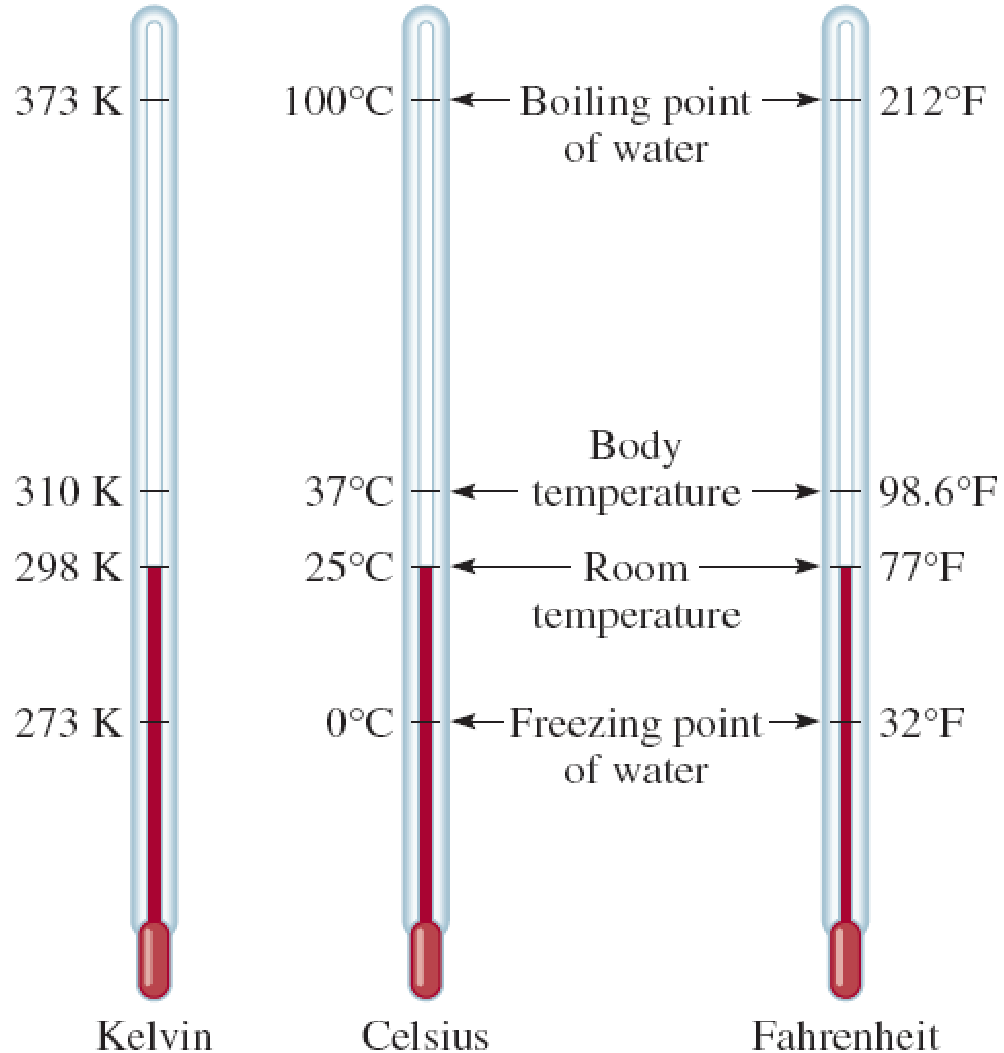
The Kelvin is based on the properties of gases.

There are no -ve Kelvin temperatures.

At -273.15 °C (0 K) called absolute zero.

$$\boxed{K = °C + 273.15}$$

$$\boxed{°C = \frac{5}{9}(°F - 32)}$$



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:

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

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

Solution:

$$(-452°\mathrm{F} - 32°\mathrm{F}) \times \frac{5°\mathrm{C}}{9°\mathrm{F}} = \boxed{-269°\mathrm{C}}$$

(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:

$$(-38.9°\mathrm{C} + 273.15°\mathrm{C}) \times \frac{1\ \mathrm{K}}{1°\mathrm{C}} = \boxed{234.3\ \mathrm{K}}$$

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:

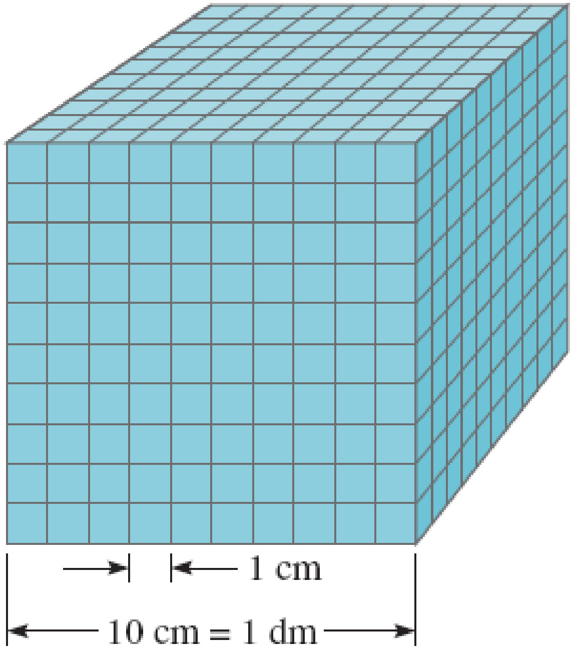
Speed is defined as the ratio of distance traveled to elapsed time.

$$\textbf{Speed} = \frac{\boldsymbol{length}}{\boldsymbol{time}} = \frac{\boldsymbol{m}}{\boldsymbol{s}}$$

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):

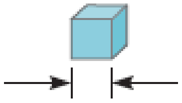
$$1\ cm^3 = (1 \times 10^{-2}\ m)^3 = 1 \times 10^{-6}\ m^3$$
$$1\ dm^3 = (1 \times 10^{-1}\ m)^3 = 1 \times 10^{-3}\ m^3$$

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

$1\ L = 1\ dm^3$

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

$1\ cm^3 = 1\ mL$



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

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

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

$$1\ g/L = 0.001\ g/mL$$

-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/cm^3)
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 cm^3. Calculate the density of gold.

Solution:

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

$$= \frac{301\ \text{g}}{15.6\ \text{cm}^3}$$

$$= \boxed{19.3\ \text{g/cm}^3}$$

Practice Exercise

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?

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:

$$m = d \times V$$

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

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

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 = m^2
Volume = L x L x L = m x m x m = m^3

Force = m a = kg (m/s^2) = kg m s^{-2} = N

Energy = ½ m v^2 = kg $(m/s)^2$ = kg m^2 s^{-2} = J

Pressure = F / A = kg m s^{-2} / m^2 = 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
602,200,000,000,000,000,000,000 (6.022x10^{23}) hydrogen atoms.
Each hydrogen atom has a mass of only
0.00000000000000000000000166 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:
$568.762 = 5.68762 \times 10^2$ **move decimal to left (n>0, +ve)**
(2) Express 0.00000772 in scientific notation:
$0.00000772 = 7.72 \times 10^{-6}$ **move decimal to right (n<0, -ve)**

Addition & Subtraction

$(7.4 \text{ x } 10^3) + (2.1 \text{ x } 10^3) = 9.5 \text{ x } 10^3$
$(4.31 \text{ x } 10^4) + (3.9 \text{ x } 10^3) = (4.31 \text{ x } 10^4) + (0.39 \text{ x } 10^4) = 4.70 \text{ x } 10^4$
$(2.22 \text{ x } 10^{-2}) – (4.10 \text{ x } 10^{-3}) = (2.22 \text{ x } 10^{-2}) – (0.41 \text{ x } 10^{-2}) = 1.81 \text{ x } 10^{-2}$

Multiplication & Division

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

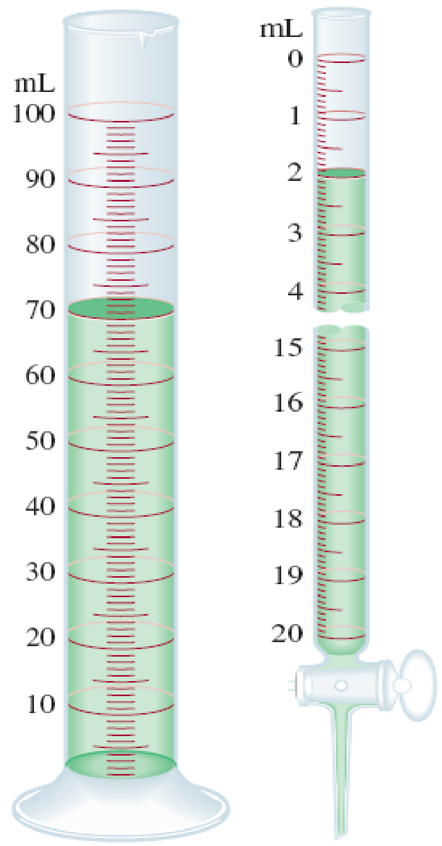
$$\frac{6.9\ x\ 107}{3.0\ x\ 10^{-5}} = \frac{6.9}{3.0}\ x10^{7-(-5)} = 2.3 \text{ x } 1012$$

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 m^3, (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:

$$\begin{array}{r l} 89.332 & \\ +\ 1.1 & \longleftarrow \text{one digit after the decimal point} \\ \hline 90.432 & \longleftarrow \text{round off to } 90.4 \end{array}$$

$$\begin{array}{r l} 2.097 & \\ -\ 0.12 & \longleftarrow \text{two digits after the decimal point} \\ \hline 1.977 & \longleftarrow \text{round off to } 1.98 \end{array}$$

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:

$$2.8 \times 4.5039 = 12.61092 \longleftarrow \text{round off to } 13$$

$$\frac{6.85}{112.04} = 0.0611388789 \longleftarrow \text{round off to } 0.0611$$

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 in = 2.54 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 g x 9 = 45 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,

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

(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 g + 0.1983 g,

$$\begin{array}{rl} & 11{,}254.1 \text{ g} \\ + & 0.1983 \text{ g} \\ \hline & 11{,}254.2983 \text{ g} \longleftarrow \text{round off to } 11{,}254.3 \text{ g} \end{array}$$

(b) 66.59 L – 3.113 L,

$$\begin{array}{rl} & 66.59 \text{ L} \\ - & 3.113 \text{ L} \\ \hline & 63.477 \text{ L} \longleftarrow \text{round off to } 63.48 \text{ L} \end{array}$$

(c) 8.16 m x 5.1355,

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

(d) 0.0154 kg ÷ 88.3 mL,

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

(e) 2.64×10^3 cm + 3.27×10^2 cm.

First we change 3.27×10^2 cm to 0.327×10^3 cm

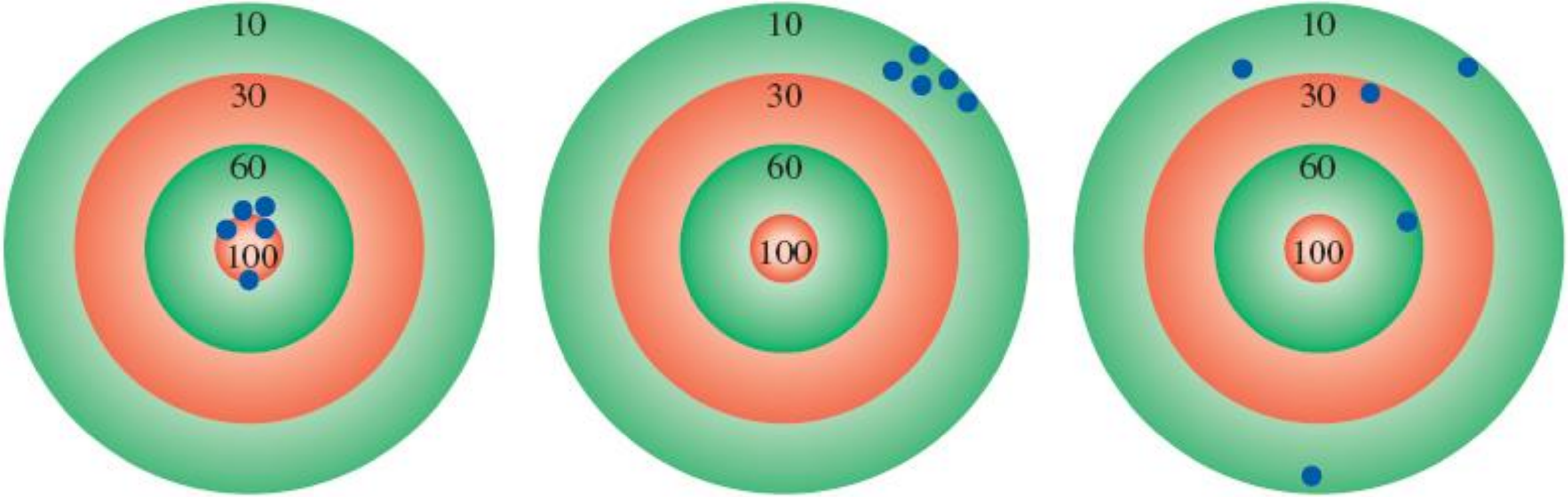
$(2.64 \text{ cm} + 0.327 \text{ cm}) \times 10^3$

2.97×10^3 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	1.971 g	1.970 g	2.001 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 $\dfrac{\mathbf{1\ in}}{\mathbf{2.54\ cm}}$ or as $\dfrac{\mathbf{2.54\ cm}}{\mathbf{1\ in}}$

Convert 12.00 in to cm ?

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

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 lb = 453.6 g
1 mg = 10^{-3} g

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

Conversion factors

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

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

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 L = 1000 cm^3
1 cm = 1 x 10^{-2} m, 1 cm^3 = $(1 \times 10^{-2})^3$ m^3, 1 cm^3 = 1 x 10^{-6} m^3.

Conversion factors

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

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

Practice Exercise

The volume of a room is 1.08 x 10^8 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 kg = 1000 g and 1 cm = 1 x 10^{-2} m or 1 cm^3 = 1 x 10^{-6} m^3

Conversion factors

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

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

Practice Exercise

The density of the lightest metal, lithium (Li), is 5.34 x 10^2 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

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


Thank You!