Chapter 1. Introduction: Matter and Measurement

1.1. The Study of Chemistry

Chemistry is the central science:

a) study of properties of materials and changes they undergo
b) can be applied to all aspects of life (e.g. pharmaceutical)
c) has a considerable impact on society

Chemistry involves the study of the properties and behavior of matter

Matter:

a) physical material of the universe
b) has mass and occupies space
c) almost 100 elements constitute all matter

1.2. Classification of Matter

Matter is classified by state (solid, liquid or gas) or by composition (element, compound or mixture) (Figure 1, Figure 2)

Gas: No fixed volume or shape, conforms the shape of container, compressible. Molecules are far apart, move at high speed, collide often.

Liquid: volume independent of container, no fixed shape, incompressible. Molecules are closer that gas, move rapidly but can slide over each other.

Solid: volume and shape independent of container, rigid, incompressible. Molecules packed closely in definite arrangements.

Pure Substances:

a) Elements : can not be decomposed into simpler substances, i.e. only one kind of atom).

- 114 known, vary in abundance.
- Each is given a unique name and a one or two-letter symbol derived from its name
- Organized in periodic table

b) Compounds: are combination of elements. The compound H_2O is a combination of elements H and O.

Law of Constant (Definite) Proportions : A compound always consists of the same combination of elements (e.g. water is always 11% H and 89% O)

Mixture: is a combination of two or more substances:

Heterogeneous (do not have uniform composition, properties and appearance, e.g.sand)

Homogeneous (uniform throughout,e.g.air). Homogeneous mixtures are solutions.

Separation of Mixtures:

a) Filtration; remove solid from liquid (Figure 3).
b) Distillation: Boil off one or more components of the mixture (Figure 4).
c) Chromatography: Exploit solubility of components (Figure 5).

1.3. Physical and Chemical Changes (Figure 6)

a) Physical change: substances changes physical appearance without altering its identity (e.g. change of state)
b) Chemical change (or chemical reactions): substances transform into chemically different substances (e.g. decomposition of water)

1.4. Units of Measurement

In 1790 (France): Metric system, first to use the Earth as standard

In 1960 (France): **SI** units system, Systeme International d'unites. Today's version of the metric system.

The units most often used for scientific measurement are those of the metric system (Figure 7).

1.5. Uncertainty in Measurement

Two types of numbers:

a) exact numbers
b) inexact numbers

All measurements have some degree of uncertainty or error associated with them.

Precision and Accuracy (Figure 8):

a) precision: how well measured quantities agree with each other
b) accuracy: how well measured quantities agree with the 'true value'.

Significant Figures:

The exactness of a measurement is reflected in the number of significant figures.
The number of significant figures is the number of digits known with certainty plus one uncertain digit.(Example; 2.2405 g means we are sure the mass in 2.240 g but we are uncertain about the nearest 0.0001 g).

Rule:

1. Nonzero numbers and zeros between nonzero numbers are always significant
2. Zeros before the first nonzero digit are not significant (example: 0.0003 has one significant figure.
3. Zeros at the end of the number before a decimal point are significant.

Numbers are written in scientific notation. 2.50×10^2 has 3 s.f. 1.030×10^4 has 4 s.f.

Significant Figures in Calculations:

a) Multiplication and Division: report to the least number of significant figures:
For example: ($6.221 \text{ cm} \times 5.2 \text{ cm} = 32 \text{ cm}^2$)

$$\text{Area} = (6.221\,\text{cm})(5.2\,\text{cm}) = 32.3492\,\text{cm}^2 \Rightarrow \text{round off to } 32\,\text{cm}^2$$

because 5.2 has two significant figures

b) Addition and substraction: report to the least number of decimal places.

For example: (20.4 g – 1.322 g = 19.1 g)

This number limits	20.42	← two decimal places
the number of significant	1.322	← three decimal places
figures in the result →	83.1	← one decimal place
	104.842	← round off to one decimal place (104.8)

we report 104.8 because 83.1 has only one decimal place

1.6. Dimensional Analysis:

Dimensional analysis is a method of calculation utilizing a knowledge of units. It is an easy way to convert from one unit of measure to another by multiplying be an appropriate conversion factor.

Conversion factors are used to manipulate units. It a fraction in which numerator and denominator are in different units, but equal to the same quantity. The algebric value of the converstion factor is always 1.

If a = b, then a/b = 1 and b/a = 1

1 ft = 12 in so, 1 ft/12 in = 1 and 12 in/1ft = 1

Given units can be multiplied and divided to give the desired units.

given unit x (conversion factor) = Desired unit

conversion factor = (desired unit) / (given unit).

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

We often need to use more than one conversion factor in order to complete a problem. The final answer must be the correct unit.

For example, 2.54 cm and 1 in. are the same length, in. This relationship allows us to write two conversion factors: 2.54cm 1 in. and 1 in. 2.54cm.

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

What is the length in centimeters of an object that is 8.50 in. long?

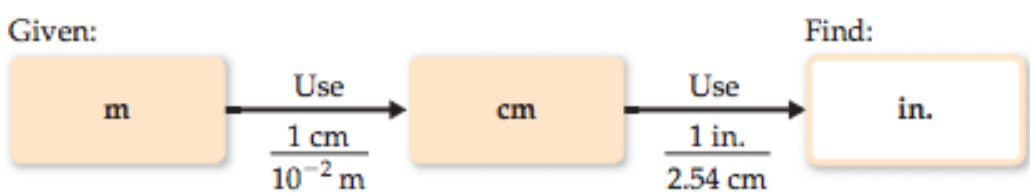
$$\text{Number of centimeters} = (8.50\,\cancel{\text{in.}})\,\frac{2.54\,\text{cm}}{1\,\cancel{\text{in.}}} = 21.6\,\text{cm}$$

Desired unit

Given unit

Using two or more conversion factors:

For example: Let us convert the length of an 8.00-m rod to inches.
We know that: 1 cm = 10^{-2} m, and 1 in = 2.54 cm.



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

Conversions Involving Volume:

Suppose that we wish to know the mass in grams of 2.00 cubic inches of gold given that the density of the gold is 19.3 g/cm^3

We know the following conversion factors:

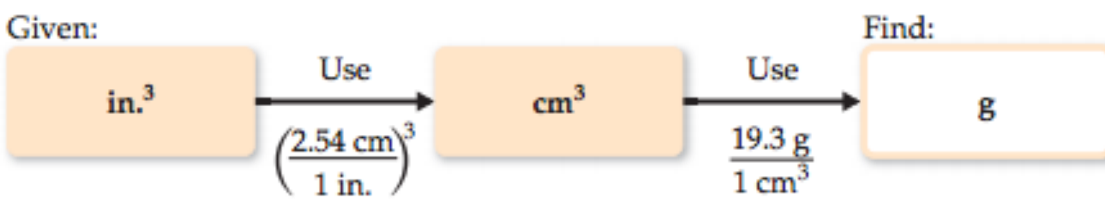
$1\ cm^3 = 19.3$ g and 2.54 cm = I inch

We can write the following conversion factors:

$$\frac{19.3\ \text{g}}{1\ \text{cm}^3} \quad \text{and} \quad \frac{1\ \text{cm}^3}{19.3\ \text{g}}$$

and

$$\frac{(2.54\ \text{cm})^3}{(1\ \text{in.})^3} = \frac{(2.54)^3\ \text{cm}^3}{(1)^3\ \text{in.}^3} = \frac{16.39\ \text{cm}^3}{1\ \text{in.}^3}$$



$(2.00\ \text{in.}^3)(2.54\ \text{cm/in.})^3\ (19.3\ \text{g gold} / 1\ \text{cm}^3) = 6.33$ g gold

$$\text{Mass in grams} = (2.00\ \cancel{\text{in.}^3})\left(\frac{16.39\ \cancel{\text{cm}^3}}{1\ \cancel{\text{in.}^3}}\right)\left(\frac{19.3\ \text{g}}{1\ \cancel{\text{cm}^3}}\right) = 633\ \text{g}$$

Summary of Dimensional Analysis:

Three questions:

1. What data are we given?
2. What quantity do we need?
3. What conversion factors are available to take us from what we are given to what we need?

1.7. Concept Map: (Figure 9)