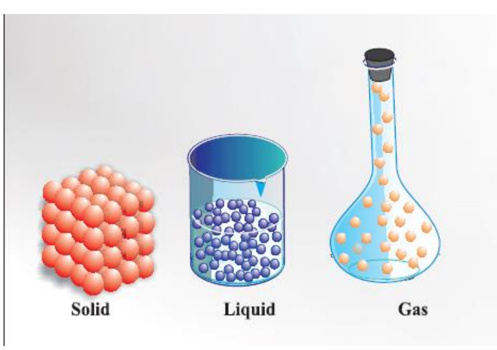
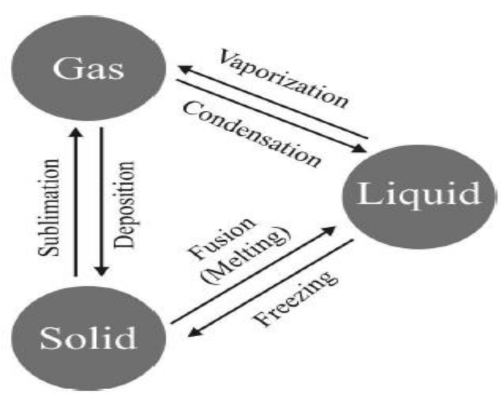
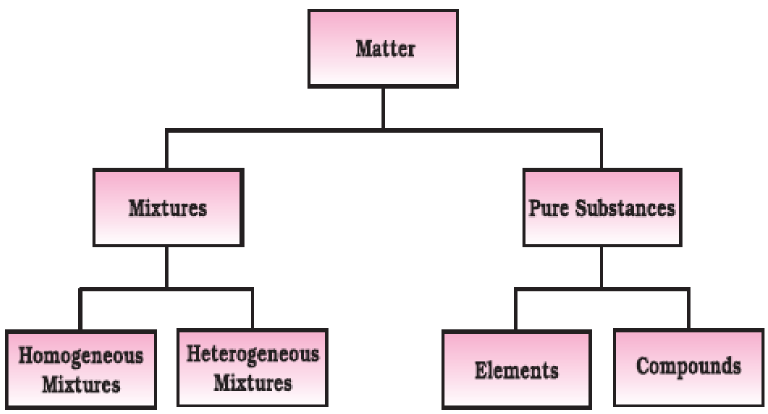
CHAPTER-1
SOME BASIC CONCEPTS OF CHEMISTRY

CHEMISTRY: Chemistry defined as the study of composition, properties and interaction of matter.
Chemistry plays a central role in science and is often intertwined with other branches of science (physics, biology, geology).
Branches of chemistry
Physical chemistry on and influence their chemical structure, properties and reactions: The branch of chemistry concerned with the way in which the physical properties of substances depend.
Inorganic chemistry: the branch of chemistry which deals with the structure, composition and behaviour of inorganic compounds.
Organic chemistry: the branch which deals with the study of structure, composition and chemical properties of organic compounds.
Analytical chemistry: the branch of chemistry dealing with separation, identification and quantitative determination of the composition of different substances.
Biochemistry : the branch which deals with the structure and behaviour of the components of cell and chemical processes in living beings.
Matter: Anything which has mass and occupies space is called matter
Example: book, pen, water, air, all living beings.
The basic constituents of matter are atoms and molecules.
States of Matter:
Matter exist in three physical states
Solid
Liquid
Gas
These three states of matter are interconvertible by changing the conditions of temperature and pressure.
solid ↔ liquid ↔ gas





Solids: Solids have definite volume and definite shape.
In solids particles are held very close to each other in an orderly fashion and there is not much freedom of movement.
Liquids: Liquids have definite volume but do not have definite shape. They take the shape of the container in which they are placed.
In liquids, the particles are close to each other but they can move around.
Gases : Gases have neither definite shape nor definite volume. They completely occupy the space in the container in which they are placed.
In gases, the particles are far apart as compare to those present in solid or liquid states and their movement is easy and fast.
Classification of matter



Mixtures: A mixture contains two or more substances present in it (in any ratio) which are called its components.
A mixture may be homogeneous or heterogeneous.
Homogeneous mixture: in this type of mixture the components completely mix with each other and its composition is uniform throughout.
Example : sugar solution
Heterogeneous mixture : in this type of mixture the composition is not uniform throughout .
Example: grains and pulses along with dirt particles
Pure substances: A material containing only one substance.
Constituent particles of pure substance have fixed composition.
Pure substances are further classified as elements and compounds.
Elements: An element is defined as a pure substance that contains only one kind of particles.
Elements are further sub-divided into
Metals
Non metals
Metalloids

Compounds: A compound is a pure substance containing two or more than two elements combined together in a fixed proportion by mass.

Physical and Chemical properties of matter

Physical properties: These are those properties which can be measured and observed without changing the identity or the composition of the substance.

Example: color, odour, melting point, boiling point, density.

Chemical properties: These are those properties in which a chemical change in the substances occurs.

Example: acidity or basicity, combustibility etc.

PHYSICAL QUANTITIES AND THEIR MEASUREMENTS

PHYSICAL QUANTITIES: all such quantities which we come across during our scientific studies are called physical quantities.

To express the measurement of any physical quantity two things are considered:

(i) Its unit,

(ii) The numerical value.

Magnitude of a physical quantity = numerical value * unit

Units: It is defined as the standard of reference chosen to measure any physical quantities.

Units are of two types:

(i) Basic units

(ii) Derived units

(i) The basic or fundamental units are those of length (m), ass (kg), time (s), electric current(A), thermodynamic temperature (K), amount of substance (mol) and luminous intensity (cd).

(ii) Derived units are basically derived from the fundamental units, e.g., unit of density is derived from units of mass and volume.

The systems used for describing measurements of various physical quantities are

(a) **CGS system** : It is based on centimetre, gram and second as the units of length, mass and

time respectively.

(b) **FPS system** : A British system which used foot(ft). pound (lb) and second (s) as the : fundamental units of length, mass and time.

(c) **MRS system** : Uses meter (m), kilogram (kg) and second (s) respectively for length, mass

and time; ampere (A) was added later on for electric current.

(d) **SI system** :(1960)International system of units and contains following seven basic and two

supplementary units:

Base Physical Quantities and their Corresponding Basic Units

Base Physical Quantity	Symbol for Quantity	Name of SI Unit	Symbol for SI Unit
Length	l	metre	m
Mass	m	kilogram	kg
Time	t	second	s
Electric current	I	ampere	A
Thermodynamic temperature	T	kelvin	K
Amount of substance	n	mole	mol
Luminous intensity	I_v	candela	cd

Unit of length	**metre**	The *metre* is the length of the path travelled by light in vacuum during a time interval of 1/299 792 458 of a second.
Unit of mass	**kilogram**	The *kilogram* is the unit of mass; it is equal to the mass of the international prototype of the kilogram.
Unit of time	**second**	The *second* is the duration of 9 192 631 770 periods of the radiation corresponding to the transition between the two hyperfine levels of the ground state of the caesium-133 atom.
Unit of electric current	**ampere**	The *ampere* is that constant current, which if maintained in two straight parallel conductors of infinite length of negligible circular cross-section and placed 1 metre apart in vacuum, would produce between these conductors a force equal to 2×10^{-7} newton per metre of length.
Unit of thermodynamic temperature	**kelvin**	The *kelvin*, unit of thermodynamic temperature, is the fraction 1/273.16 of the thermodynamic temperature of the triple* point of water.
Unit of amount of substance	**mole**	1. The *mole* is the amount of substance of a system, which contains as many elementary entities as there are atoms in 0.012 kilogram of carbon-12; its symbol is 'mol'. 2. When the mole is used, the elementary entities must be specified and these may be atoms, molecules, ions, electrons, other particles, or specified groups of such particles.
Unit of luminous intensity	**candela**	The *candela* is the luminous intensity, in a given direction, of a source that emits monochromatic radiation of frequency 540×10^{12} hertz and that has a radiant intensity in that direction of 1/683 watt per steradian.

<u>Prefixes used in SI system</u>:

Multiple	Prefix	Symbol
10^{-24}	yocto	y
10^{-21}	zepto	z
10^{-18}	atto	a
10^{-15}	femto	f
10^{-12}	pico	p
10^{-9}	nano	n
10^{-6}	micro	μ
10^{-3}	milli	m
10^{-2}	centi	c
10^{-1}	deci	d
10	deca	da
10^{2}	hecto	h
10^{3}	kilo	k
10^{6}	mega	M
10^{9}	giga	G
10^{12}	tera	T
10^{15}	peta	P
10^{18}	exa	E
10^{21}	zeta	Z
10^{24}	yotta	Y

SCIENTIFIC NOTATION:

In such notation, all measurements (however large or small) are expressed as a number between 1.000 and 9.999 multiplied or divided by 10. In general as
$N \times 10$
Here, N is called digit term (1.000-9.999) and n is known as exponent.
e.g., 138.42 cm can be written as 1.3842×10^{2} and 0.0002 can be written as 2.0×10^{-4}.

SIGNIFICANT FIGURES:

Significant figures are the meaningful digits in a measured or calculated quantity. It includes all those digits that are known with certainty plus one more which is uncertain or estimated.
Greater the number of significant figures in a measurement, smaller the uncertainty.

Rules for determining the number of significant figures are:

1. All non-zero digits are significant. For example in 285 cm, there are three significant figures and in 0.25 mL, there are two significant figures.
2. Zeros preceding to first non-zero digit are not significant. Such zero indicates the position of decimal point. Thus, 0.03 has one significant figure and 0.0052 has two significant figures. Zeros between two non-zero digits are significant. Thus, 2.005 has four significant figures.
3. Zeros at the end or right of a number are significant, provided they are on the right side of the decimal point. For example, 0.200 g has three significant

figures. The terminal zeros are not significant if there is no decimal point. For example, 100 has only one significant figure, but 100. has three significant figures and 100.0 has four significant figures.

4. Counting the numbers of object, for example, 2 balls or 20 eggs, have infinite significant figures as these are exact numbers and can be represented by writing infinite number of zeros after placing a decimal i.e.,2 = 2.000000 or 20 = 20.000000.

Calculations Involving Significant Figures

In addition or subtraction, the final result should be reported to the same number of decimal places as that of the term with the least number of decimal places, e.g.,

$$\begin{array}{l} 2.512 \\ 2.2 \\ 5.23 \\ \hline 9.942 \Rightarrow 9.9 \end{array}$$

In multiplication and division, the result is reported to the same number of significant figures as least precise term or the term with least number of significant figures, e.g.

Rounding Off the Numerical Results

When a number is rounded off, the number of significant figures is reduced. the last digit retained is increased by 1 only if the following digit is ≥ 5 and is left as such if the following digit is ≤ 4, e.g.,

n be written as 12.7

18.35 can be written as 18.4

13.93 can be written as 13.9

<u>Precision And Accuracy</u>

Precision: Precision refers the closeness of the set of values obtained from identical measurements of a quantity. Precision is simply a measure of reproducibility of an experiment.

Precision = individual value – arithmetic mean value

Accuracy: Accuracy is a measure of the difference between the experimental value or the mean value of a set of measurements and the true value.

Accuracy = mean value – true value

DIMENSIONAL ANALYSIS

Often while calculating, there is a need to convert units from one system to other. The method used to accomplish this is called factor label method or unit factor method or dimensional analysis.

Important Conversion Factor

$1 \text{dyne} = 10^{-5}$ N	$1 L = 1000$ mL
$1 \text{atm} = 101325 \text{ Nm}^{-2}$	$= 1000 \text{ cm}^3$
$= 101325$ Pa (pascal)	$= 10^{-3} \text{ m}^3$
$1 \text{bar} = 1 \times 10^5 \text{ Nm}^{-2}$	$= 1 \text{ dm}^3$
1 L atm $= 101.3$ J $= 24.21$ cal	1 gallon $= 3.7854$ L
$1 \text{cal} = 4.184 \text{ J} = 2.613 \times 10^{19}$ eV	1 eV/atom $= 96.485 \text{ kJ mol}^{-1}$
$1 \text{eV} = 1.602189 \times 10^{-19}$ J	1 amu or u $= 1.66 \times 10^{-27}$ kg
$1 \text{ J} = 10^7$ erg	$= 931.5$ MeV
$1 \text{ Å} = 10^{-10}$ m	$1 \text{esu} = 3.3356 \times 10^{-10}$ C

Laws of Chemical Combinations

The combination of elements to form compounds is governed by the following basic laws:

Law of conservation of mass (Lavoisier, 1774): This law states that in any physical or chemical change the total mass of the product is equal to the total mass of the reactant. ‘Matter can neither be created nor destroyed’.

Law of definite proportions (Proust, 1799): He stated that *a given compound always contains exactly the same proportion of elements by weigh.*

Law of multiple proportions (Dalton, 1803): According to this law, *if two elements can combine to form more than one compound, the masses of one element that combine with a fixed mass of the other element, are in the ratio of small whole numbers.* e.g., in NH_3 and N_2H_4 fixed mass of nitrogen requires hydrogen in the ratio 3 : 2.

Gay Lussac’s law of gaseous volumes: *when gases combine or are produced in a chemical reaction they do so in a simple ratio by volume, provided all gases are at the same temperature and pressure.* Example: 100 mL of hydrogen combine with 50 mL of oxygen to give 100 mL of water vapour.

Hydrogen + Oxygen → Water

100ml 50ml 100 mL

Thus, the volumes of hydrogen and oxygen which combine (i.e., 100 mL and 50 mL) bear a simple ratio of 2:1.

Avogadro’s Law: Avogadro proposed that *equal volumes of all gases at the same temperature and pressure should contain equal number of molecules.*

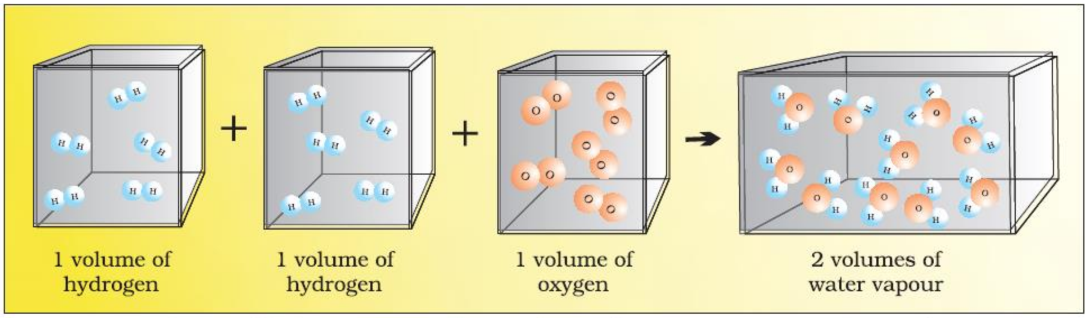


Fig. 1.9 Two volumes of hydrogen react with one volume of oxygen to give two volumes of water vapour

Dalton's Atomic Theory (1803)

This theory was based on laws of chemical combinations. It's basic postulates are
1. All substances are made up of tiny indivisible particles, called atoms.
2. In each element, the atoms are all alike and have the same mass. The atoms of different elements differ in mass.
3. Atoms can neither be created nor destroyed during any physical or chemical change.
4. Compounds or molecules result from combination of atoms in some simple numerical ratio.

Atomic mass: It is the average relative atomic mass of an atom. It indicates that how many times an atom of that element is heavier as compared with 1 / 12 of the mass of an atom of carbon.

Average atomic mass: The word average has been used in the above definition and is very significant because elements occur in nature as mixture of several isotopes.

For example, carbon has the following three isotopes with relative abundances and masses as shown against each of them.

Isotope	Relative Abundance (%)	Atomic Mass (amu)
^{12}C	98.892	12
^{13}C	1.108	13.00335
^{14}C	2×10^{-10}	14.00317

The average atomic mass of carbon will come out to be:

(0.98892) (12 u) + (0.01108) (13.00335 u) + (2 × 10–12) (14.00317 u) = 12.011 u

Molecular mass: It is the mass of a molecule, i.e., number of times a molecule is heavier than 1 / 12 th mass of C-12 atom. Molecular mass of a substance is an additive property and can be calculated by taking algebraic sum of atomic masses of all the atoms of different elements present in one molecule.

Molecular Mass = average relative mass of one molecule / 1 / 12 th * mass of C-12 atom

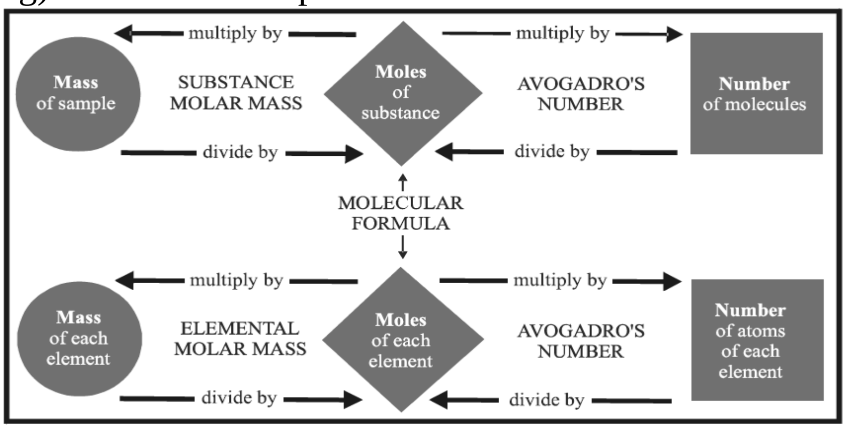
For example, molecular mass of methane, which contains one carbon atom and four hydrogen atoms, can be obtained as follows:

Molecular mass of methane,

(CH4) = (12.011 u) + 4 (1.008 u)

= 16.043 u

Mole concept: One mole is the amount of a substance that contains as many particles or entities as there are atoms in exactly 12 g (or 0.012 kg) of the 12C isotope.



Note...

- Mass of 1 atom in amu and mass of 6.022×10^{23} atoms in grams are numerically equal.
- When atomic mass is taken in grams it is also called the molar atomic mass.
- 6.022×10^{23} is also called 1 mole of atoms and this number is also called the **Avogadro's Number.**
- Mole is just a number. As 1 dozen = 12;

  1 million = 10^6; 1 mole = 6.022×10^{23}.

Percentage composition:

$$\text{Mass \% of an element} = \frac{\text{mass of that element in the compound}}{\text{molar mass of the compound}} \times 100$$

Empirical and Molecular Formulae

Empirical formula is the simplest formula of a compound giving simplest whole number ratio of atoms present in one molecule, e.g., CH is empirical formula of benzene (C_6H_6).

Molecular formula is the actual formula of a compound showing the total number of atoms of constituent elements, e.g., C_6H_6 is molecular formula of benzene.

Molecular formula = (Empirical formula)n where, n is simple whole number having values 1, 2, 3, … , etc., and can be calculated as

n = molecular formula mass / empirical formula mass

Stoichiometry

The relative proportions in which the reactants react and the products are formed, is called stoichiometry (from the Greek word meaning 'to measure an element'.)

Limiting reagent: It is the reactant which is completely consumed during the reaction.

Excess reagent: It is the reactant which is not completely consumed and remains unreacted during the reaction.

[In a irreversible chemical reaction, the extent of product can be computed on the basis of limiting reagent in the chemical reaction]

Reactions in solutions: The concentration of a solution or the amount of substance present in its given volume can be expressed in any of the

Following ways.

1. Mass per cent or weight per cent (w/w %)
2. Mole fraction
3. Molarity
4. Molality

Mass percent:

$$\text{Mass per cent} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$

Mole fraction: It is the ratio of number of moles of a particular component to the total number of moles.

If a substance 'A' dissolves in substance 'B' and their number of moles are *n*A and *n*B, respectively, then the mole fractions of A and B are given as :

$$\text{Mole fraction of A} = \frac{\text{No. of moles of A}}{\text{No. of moles of solutions}} = \frac{n_A}{n_A + n_B}$$

$$\text{Mole fraction of B} = \frac{\text{No. of moles of B}}{\text{No. of moles of solutions}} = \frac{n_B}{n_A + n_B}$$

Molarity: It is defined as the number of moles of the solute in 1 litre of the solution.

$$\text{Molarity (M)} = \frac{\text{No. of moles of solute}}{\text{Volume of solution in litres}}$$

Molality: It is defined as the number of moles of solute present in 1 kg of solvent.

$$\text{, Molality (m)} = \frac{\text{No. of moles of solute}}{\text{Mass of solvent in kg}}$$