CLASS-11

CHEMISTRY

CHAPTER-

“SOME BASIC CONCEPTS OF CHEMISTRY”

*. Matter → Anything that occupies mass & space.

States of Matter

→ Solid (Def. volume & shape)
→ Liquid (Def. volume but no shape)
→ Gas (Neither def. volume nor shape).

(Interconvertible by changing Temp. & Pressure.

$$s \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} l \underset{\text{cool}}{\overset{\text{heat}}{\rightleftharpoons}} g$$

Classification of matter based at bulk level on the basis of purity

- Mixture (Contains 2 / more substance)
  - Homogeneous (Uniform) composition
    Eg: Sugar solution
  - Heterogeneous (Variable composition)
    Eg:- Mixture of grains & pulses.
- Pure substances (Single component)
  - Elements (Same type of atoms)
    Eg :- O_2, He
  - Compounds (Diff. elements)
    Eg: H_2O

Properties of matter & measurement:-

→ Physical properties: Can be measured without changing the identity / composition of substance

→ Chemical properties: Measurement require chemical change to occur

Base Physical Quantities and their 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

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

Laws of chemical combination :-

* Law of conservation of mass.
(Mass can neither be created nor destroyed. (except nuclear rxn.))

* Law of definite / constant proportion.
Compounds contains fixed ratio of elements (by wt), irrespective of the source. (exception: isotopes)

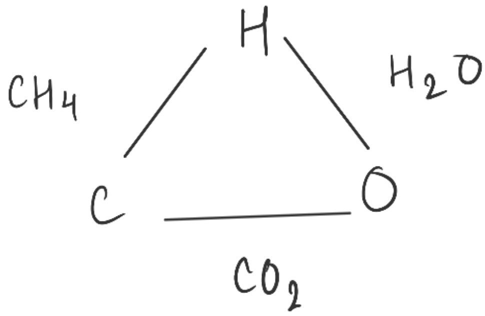
* Law of Multiple proportion.
When two elements combine to form more than one compound, the mass of one element that combines with fixed mass of other element are in the ratio of fixed whole numbers.

$$\underset{2g}{H} + \underset{16g}{O} \longrightarrow \underset{18g}{H_2O}$$

$$\underset{2g}{H} + \underset{32g}{O} \longrightarrow \underset{34g}{H_2O_2}$$

Mass of H (fixed) = 2gm, mass of O ratio $\frac{16}{32} = 2:1$

When two elements combine separately with a fixed mass of third element, the ratio of their masses is either same or simple whole number multiple of the ratio in which they combine with each other.



CH_4 $\frac{12}{4}$

CO_2 $\frac{12}{32}$

$\rightarrow \frac{4}{32} = \frac{1}{8}$

H_2O $\frac{2}{16} = \frac{1}{8}$

* Gay lussac's law of constant volumes

(When gases react, they do so in simple whole no. ratio of volume, when temp. & pressure are const.)

* Avogadro's law

(equal volume of gases, at the same temp. & pressure have the same no. of molecules)

* Dalton's Atomic Theory

→ Matter consist of indivisible atoms.

→ Compounds are formed when atoms of diff. elements combine in a fixed ratio.

→ All atoms of a given element are identical.

→ Atoms of diff. element have diff. masses & other properties.

→ Atoms are neither created nor destroyed.

$1u = 1.667 \times 10^{-24} g$ = Mass of $\frac{1}{12}$ of C-12 isotope.

$H = 1 \times 1.667 \times 10^{-24} g = 1u$

$C = 12 \times 1.667 \times 10^{-24} g = 12u$

Relative atomic mass unit : u or amu.

Atomic mass : gms.

* Atomic mass $= \dfrac{\text{Mass of an element}}{\frac{1}{12} \times \text{Mass of C-12 element}} = \dfrac{\text{Mass of atom (amu)}}{1 \text{ amu}}$

* Average atomic mass

$$= \frac{(\% \text{ abundance of } A_1 \times A_1) + (\% \text{ abundance of } A_2 \times A_2)}{100}$$

* Molecular mass : Sum of atomic mass of element present in molecule.

Relative molecular mass

[$CH_4 = 12u + 4 \times 1u = 16u$]

[Actual molecular mass $= 16 \times 1.66 \times 10^{-24} gm$]

Formula mass

For ionic compounds, the sum of the atomic masses of all atoms represents formula mass.

$= 23\,u + 35.5\,u$
$= 58.5\,u$

*. Mole concept & Molar mass (Mass of 1 mol of substance)

* 1 mole = 6.022×10^{23} number $= N_A$
  ↳ represents avogadro's number of any entity.

* No. of moles = $\frac{\text{given entities}}{6.022 \times 10^{23}}$

* 1 g atom = 1 mole of atom.
* 1 g molecule = 1 mole of molecule.

* Numerically, R.A.M = M.M

Eg:- R.A.M of $H_2 = 2\,u$; Molar mass of $H_2 = 2\,gm$

* No. of moles = $\frac{\text{given mass}}{M.M} = \frac{\text{given volume (L)}}{\text{Molar volume (L/mol)}}$

* Avg. Molecular mass = $\frac{\text{Total mass}}{\text{total moles}} = \frac{n_1M_1 + n_2M_2 + \ldots}{n_1 + n_2 + \ldots}$

* STP	* NTP
P = 1 bar	P = 1 atm
T = 273 K	T = 273 K
$V_m = 22.7\,L$	$V_m = 22.4\,L$

* V_m = Molar volume (Vol of 1 mol of gas)

* No. of entities = No. of mol × N_A

Mole percentage

$$\% \text{ of } n_A = \frac{n_A}{n_T} \times 100$$

* Vapour density = $\frac{\text{Molar Mass}}{2}$

* Percentage composition of element

$$= \frac{\text{Mass of that element}}{\text{Molar mass of compound}} \times 100$$

* Empirical formula → Simplest whole no. ratio of elements present in the compound.

Eg: $C_6H_{12}O_6 \rightarrow$ E.F = $\cancel{6} : \cancel{12}^{2} : \cancel{6}^{1}$

= 1 : 2 : 1

∴ E.F = $C_1H_2O_1$

* Molecular formula → exact no. of atoms present in compound.

Eg:- $C_6H_{12}O_6$

* Molecular formula = n × Empirical formula.

$$n = \frac{\text{Molar mass}}{\text{Empirical formula mass}}.$$

mass percent :-

1. Mass percent → grams
2. grams → moles
3. Moles ÷ Smallest no. of moles = simple whole no. ratio.
4. Write empirical & molecular formula.

		Mass (g)	No. of mol	mol
Eg:- H	4.07%	= 4.07 g	= 4.07/1	= 4.07
C	24.27%	= 24.27 g	= 24.27/12	= 2.02
Cl	71.65%	= 71.65 g	= 71.65/35.5	= 2.02

Simplest whole no. ratio

$\frac{4.07}{2.02} = 2$ $\quad \frac{2.02}{2.02} = 1$ $\quad \frac{2.02}{2.02} = 1$

E.F = H_2CCl = 2 + 12 + 35.5 = 49.5

Molar mass = 98.96 (given)

then, $n = \frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{98.96}{49.5}$

$n = 2$

*. STOCHIOMETRY :-

It deals with calculation of moles, molar masses & volume of reactants & products in balanced chemical reaction.

* Limiting reagent - Reactant which gets completely consumed during the reaction is called limiting reagent

react completely is called excess reagent.

$$1A + 2B + 3C \rightarrow AB_2C_3$$

Moles given : 2   3   3

Then $L\cdot R = \frac{2}{1} = 2$; $\frac{3}{2} = 1.5$; $\frac{3}{3} = 1$ (least value of C)

$\therefore C = L\cdot R$

* $\% \text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$

* $\% \text{ purity} = \frac{\text{Mass of pure sample}}{\text{Mass of impure sample}} \times 100$

⇒ Concentration terms :-

1. $\% \text{ Mass } (W/W) = \frac{\text{Mass of the component}}{\text{Total mass of solution}} \times 100$

2. $\% \text{ Volume } (V/V) = \frac{\text{Volume of the component}}{\text{Total volume of solution}} \times 100$

3. $\text{Normality} = \frac{\text{No. of gm equivalent of solute}}{\text{Vol. of solution (in L)}}$

4. $\% \text{ Mass / volume } (w/v) = \frac{\text{Mass of the component}}{\text{Total volume of solution}} \times 100$

5. $\text{ppm} = \frac{\text{No. of parts of the component}}{\text{Total no. of parts of all components of solution}} \times 10^6$.

6. $\text{Mole fraction} = \frac{\text{No. of moles of the component}}{\text{Total no. of moles of all components}}$

7. $\text{Molarity} = \frac{\text{No. of moles of solute}}{\text{Vol. of solution (in L)}}$

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