

- ✓ → Structure of Atom (4hrs)
- Periodic properties
- ✓ → Chemical Bonding
- Water Chemistry

Basic Concept of Chemistry

- Mole concept
- Empirical formulae
- Oxidation State

Eng = (Force x dist travelled)

length

$1 \text{ \AA} = 10^{-10} \text{ m}$      $1 \text{ nm} = 10^{-9} \text{ m}$      $1 \text{ \mu m} = 10^{-6} \text{ m}$

Mass

$1 \text{ mg} = 10^{-3} \text{ g}$      $1 \text{ \mu g (micro-gram)} = 10^{-6} \text{ g}$

$1 \text{ amu} = \frac{1}{12} \text{ th mass of an atom of } C^{12} \text{ (isotope)}$

Volume

✓  $1 \text{ L} = (10 \text{ cm})^3 = 10^3 \text{ m}^3 = 10^3 \text{ dm}^3$

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

Temp:-

SI unit → Kelvin

$0^\circ \text{ K} = -273.15^\circ \text{ C}$

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

Fahaenheit

$32^\circ \text{ F} \rightarrow \text{Freezing point}$

$212^\circ \text{ F} \rightarrow \text{BP}$

$180^\circ - \Delta T$

Celsius Scale

$0^\circ \text{ C} \rightarrow \text{Ice point}$

$100^\circ \text{ C} \rightarrow \text{Steam point}$

$\Delta T = 100$

$\text{C} \rightarrow \text{F}$

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

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

# Pressure .

1 atm = Pressure exerted by a column of mercury of 760 mm at 0°C.

Torr → 1 mm column of mercury.

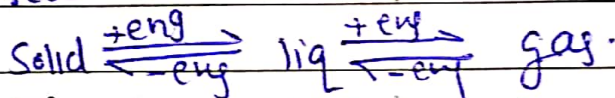
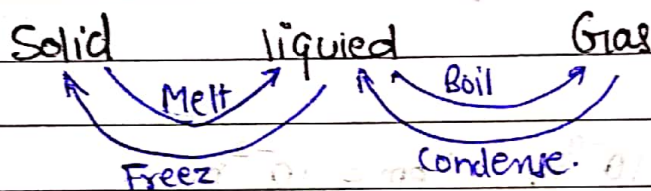
∴ 1 atm = 760 torr = 760 mm of Hg = 1.01325 × 10<sup>5</sup> Pa

# # Energy

calorie → Amt of heat req to raise temp of 1 gram of water from 14.5°C - 15.5°C.

1 C = 4.184 J.

1 J = 0.2390 Cal.



Latin names of element.

1. Na - Sodium → Natrium.
2. Au - Gold → Aurum.
3. Cu - Copper → Cuprum.
4. K - Potassium → (Kalvin).
5. Fe - Iron → (Ferrum).
6. Hg - Mercury - (Hydrargyrum).
7. Ag - Silver - (Argentum).
8. W - Tungsten - (Wolfram)
9. Pb → lead (Plumbum).

Transuranic elements → Man-made.

Non-metals - Solid → C, B, P, S, Selenium & Iodine.

liq → Bromine.

→ Metals → rarely combine with each other

- Components in mixture are present without loss of identity.

### Mixture ~~Classification~~

• Homogeneous

• Heterogeneous

→ Mixture - Uniform & in single phase.

- Not uniform.

→ Isotropic in Nature

- Can have two or more phase.

(every portion of it has same composition & properties).

- Component can be seen

ex - Air, gasoline, Alloy.

• Anisotropic properties - properties are not uniform throughout mix.

ex - Soil, Smoke

• Component can't be seen with naked eye / microscope.

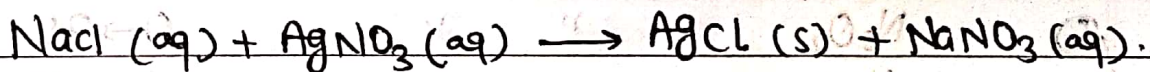
### Laws:-

1) Law of Conservation of Mass

Lavoisier - 1774 - law of Indestructibility of Matter.

- In a chemical reaction, mass is neither created nor destroyed.

- Tested by Landolt.



Total mass of reactants = Total mass of products.

Total mass of reactants = Total mass of product + Mass of unreacted reactants.

2) Law of Definite or Constant Proportion

Proust - 1799

A Chemical Compound always contains the same elements combined together in fixed proportion by mass.

or //

A CC has fixed composition & it does not depend upon method of its preparation or the source from which it is obtained.

$\begin{matrix} \rightarrow & \text{heating } \text{CaCO}_3 \\ \rightarrow & \text{heating } \text{Na}_2\text{CO}_3 \text{ (Sodium bicarbonate)} \\ \rightarrow & \text{burning } \text{C in O}_2. \end{matrix}$

$\text{CO}_2$   $\Rightarrow$   $\frac{12}{32} \Rightarrow \frac{3}{8}$   $\rightarrow$  Reaction of  $\text{CaCO}_3$  with  $\text{HCl}$

Converse is not true.

When same elements combine in same proportion, same compound will form.

$\text{C, H, O} \rightarrow$  If combine in ratio 12:3:8. Can form  $\text{C}_2\text{H}_5\text{OH}$  or  $(\text{CH}_3\text{OCH}_3) \rightarrow \text{C}_2\text{H}_5\text{OH}$  (di-methyl ether)

$\rightarrow$  limitations ex  $\rightarrow \text{CO}_2 \rightarrow \text{C}^{12} \Rightarrow \text{C:O} :: 12:32$  } isotopes - diff atoms  
 $\text{CO}_2 \rightarrow \text{C}^{14} \Rightarrow \text{C:O} :: 14:32$  } max.

ex -  $\text{H}_2\text{O}^{16}$  &  $\text{H}_2\text{O}^{18} \rightarrow \text{H:O} \rightarrow 1:8$  &  $1:9$ .

**\* Multiple Proportion**

$\rightarrow$  Dalton - 1808.

$\rightarrow$  If two elements combine to form more than 1 compound then diff masses of 1 element - which combine with fixed mass of other element, bear a simple ratio, to one another.

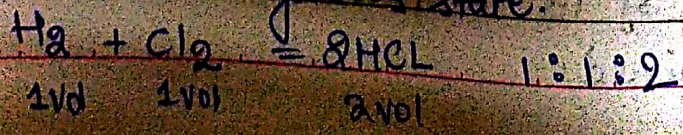
fixed mass. $\begin{matrix} \downarrow \\ \text{N} \\ \downarrow \\ \text{O} \end{matrix}$	ex	$\text{N}_2\text{O}$	Nitrogen $\rightarrow$ 28 parts	Oxygen $\rightarrow$ 16 parts.
		$\text{N}_2\text{O}_2$	" "	O $\rightarrow$ 32
		$\text{N}_2\text{O}_3$	" "	O $\rightarrow$ 48 parts.
		$\text{N}_2\text{O}_4$	" "	O $\rightarrow$ 64 parts.
		$\text{N}_2\text{O}_5$	" "	O $\rightarrow$ 80 parts.

ex	$\text{CO}$	Carbon - 12	Oxygen $\rightarrow$ 16 parts
	$\text{CO}_2$	" "	" $\rightarrow$ 32 parts

$\rightarrow$  Limitation  $\rightarrow$  for Non-stoichiometric compound.

**\* Law of gaseous Vol  $\rightarrow$  Gay-Lussac -**

Gases react with each other in simple ratio of their volume & if product is also in gaseous state.



## Non-stoichiometric Compound

→ The compound which do not follow law of constant proportion

ex-  $C_{12}H_{22}O_{11}$  (ratio of C, H → Not integral)

Date \_\_\_\_\_  
DELTA Pg No. \_\_\_\_\_

**Molecule** - Smallest particle of a compound or an element that can have a stable & independent existence.

\* Some compounds are composed of ions rather than molecules.

Ions are electrically charged particles of matter. Charge can be +ve or -ve.

Positive Charge → Cations      Negative Charge → Anions.

Simple cations and anions come into existence by loss of electrons by neutral atom.

- When ions are present in a compound, sum of positive charges on cation must balance with negative charge on anion to produce electrically neutral matter.

ex

$H^+$  - Hydrogen

$Ba^{2+}$  → Barium

$Al^{3+}$  → Aluminium.

$Ag^+$  - Silver

$Fe^{2+}$  → Ferrous

$Fe^{3+}$  → Ferric

$Na^+$  - Chloride

$Cu^{2+}$  → Cupric

$Au^{3+}$  → Auric

$F^-$  → Fluoride

$S^{2-}$  → Sulphide

$BO_3^{3-}$  → Borate

$CO_3^{2-}$  → Carbonate.

## Atomic & M. Mass:-

Atomic mass - ① Hydrogen → Oxygen → C-12

Atomic mass of an element can be defined as  $A_r$  which indicates how many times the mass of one atom of element is heavier in comparison to  $\frac{1}{12}$ th part of mass of one atom of Carbon-12.

$$A = \left( \frac{\text{Mass of one atom of element}}{(\frac{1}{12})^{\text{th}} \text{ of mass of one atom of Carbon-12}} \right)$$

Atomic Mass Unit (Amu) → the quantity  $(\frac{1}{12})^{\text{th}}$  mass of an atom of C-12.  
 Actual mass of 1 atom of C-12 =  $1.9924 \times 10^{-23}$  g.

$$1 \text{ AMU} = \left( \frac{1.9924 \times 10^{-23}}{12} \right) = 1.66 \times 10^{-24} \text{ g}$$

$$A = \left( \frac{\text{Mass of one atom of element}}{1 \text{ AMU}} \right)$$

There are avg Relative mass. } Actual mass of an atom of element = (At. mass in Amu)  $\times 1.66 \times 10^{-24}$  g  
 ex Hydrogen →  $1.008 \text{ amu} = 1.008 \times 1.66 \times 10^{-24} \text{ g} = 1.6736 \times 10^{-24} \text{ g}$   
 Oxygen →  $16.00 \text{ amu} = 16 \times 1.66 \times 10^{-24} \text{ g} = 2.656 \times 10^{-23} \text{ g}$

Isotopes → Atom of same element - having different atomic masses.

★ Gram-Atomic Mass → Atomic mass of an element - when expressed in gram.  
 → Absolute mass of 10 atom =  $16 \text{ amu} = 16 \times 1.66 \times 10^{-24} \text{ g}$ .  
 Mass of  $6.022 \times 10^{23}$  atoms =  $16 \times 1.66 \times 10^{-24} \times 6.022 \times 10^{23} = 16 \text{ g}$ .

→ GAM - of an element - can be defined as - mass of  $6.022 \times 10^{23}$  atom in gram of any element.

Molecular Mass → It is a nm which indicates - how many times molecule of substance is heavier in comparison to  $\frac{1}{16}$ th mass of an oxygen-atom

②  $(1/12)^{th}$  mass of one-atom of carbon-12.

$$H_2O = MM = (2 \times 1.008) \text{ amu} + 16.00 = 18.016 \text{ amu.}$$

$$H_2SO_4 = (2 \times 1.008) \text{ amu} + 32.06 \text{ amu} + (4 \times 16.00) = 98.016 \text{ amu.}$$

Gram Molecular Mass

$$O_2 = 32 \text{ amu. (MM). (Actual mass of 1 molecule of } O_2)$$

$$GMM = 32 \text{ g.} = 32 \times 1.66 \times 10^{-24} \text{ g}$$

$$\text{No of gm molecules} = \frac{\text{Mass of Sub (in gram)}}{\text{Molecular Mass of Sub (gram)}}$$

Mole Concept: — A mol is defined as no of atoms in 12.00g of Carbon-12.

The no of atoms in 12.00g of carbon-12, experimentally found to be  $6.022 \times 10^{23}$ .

A mole contains  $6.022 \times 10^{23}$  units. — (atoms, molecules, ion, electron)

→ How much does one mole weight?

The mass of 1 mole of any atoms/elements — is exactly equal to atomic mass in grams of that element.

$$\rightarrow \text{Mass of 1 atom of Al} = 27 \text{ amu} = 27 \times 1.66 \times 10^{-24} \text{ g}$$

$$\text{Mass of 1 mole of Al} = 27 \times 1.66 \times 10^{23} = 27 \text{ g.}$$

→ Molecular mass of water = 18 amu

$$\text{Mass of 1 mole of water} = 18 \times 1.66 \times 10^{-24} \times 6.022 \times 10^{23}$$

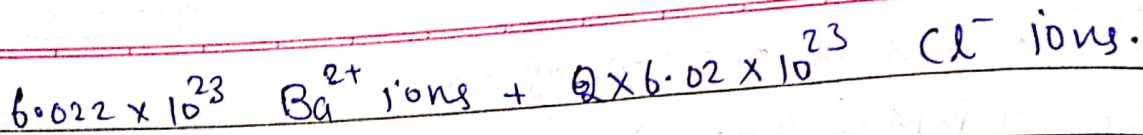
$$= 18 \text{ g.}$$

For ionic Compounds — As these compound does not contain molecules  
Formula of ionic compound represent ratio b/w its constituent ions.

The mass of  $6.022 \times 10^{23}$  formulae units — represent one mole of ionic compound.

$NaCl$ ,  $BaCl_2$

$$\text{One mole of } BaCl_2 = 6.022 \times 10^{23} \text{ units} = 208.2 \text{ g of } BaCl_2$$



\* One mole =  $6.022 \times 10^{23}$  Molecules of any gaseous substance occupies 22.4 litre of volume at NTP.

$$\text{No of moles} = \frac{\text{No of particles}}{N_A} \quad N_A = 6.022 \times 10^{23}$$

$$\text{No of moles} = \left[ \frac{\text{Given Weight}}{\text{Molar Mass (in grams)}} \right]$$

$$\text{Mass of one atom of element} = \left( \frac{\text{Gr atom of element}}{N_A} \right)$$

$$= (\text{Atomic Mass})_{\text{amu}} \times 1.66 \times 10^{-24} \text{ g.}$$

Mass of 1 molecule of substance

$$= \frac{\text{Gram-molecular mass}}{N_A}$$

$$= (\text{Molecular Mass})_{\text{in amu}} \times 1.66 \times 10^{-24} \text{ g.}$$

⊙ No of molecules present in V litre of gas at NTP.

$$= \left[ \frac{\text{Vol of gas (in litres)} \times 6.022 \times 10^{23}}{22.4 \text{ litres}} \right]$$

Q. Calculate mass of single atom of Sulphur & simple mole of  $\text{CO}_2$ .

G-Atomic mass of S = 32 g.

Mass of 1 atom of S = 32

$$\frac{32}{6.022 \times 10^{23}} = 5.31 \times 10^{-23} \text{ g.}$$

$\text{CO}_2 = 44 \text{ g. (GMM)}$

Mass of 1 mole = 44 g

$$\frac{44}{6.022 \times 10^{23}} = 7.308 \times 10^{-23} \text{ g.}$$

Q. Given weight = 0.9 g.

Molar mass = 18 g/mol.



$$n = \frac{0.1 \text{ g}}{18} = \frac{1}{20} \text{ moles.}$$

$$= 0.05 \times 6.022 \times 10^{23} = 3.01 \times 10^{22}$$

1/mole  $\rightarrow$  1 molecule of  $\text{H}_2\text{O}$   $\rightarrow$  1 oxygen atom.

$$\frac{1}{20} \text{ molecules of } \text{H}_2\text{O} = \frac{1}{20} \text{ mole of oxygen atom.}$$

a.  $N = 3.01 \times 10^{22}$  molecules of  $\text{NH}_3$ .

$$n = \left( \frac{N}{N_A} \right) =$$

$\text{NH}_3 = 17 \text{ g.}$  (17 g of  $\text{NH}_3$  contains  $6.022 \times 10^{23}$  molecules).

$$1 \text{ Mole } = \left( \frac{17}{N_A} \text{ g} \right)$$

for  $3.01 \times 10^{22}$  molecules =  $\frac{17}{6.022 \times 10^{23}} \times 3.01 \times 10^{22}$

$$\Rightarrow 0.85 \text{ g.}$$

b. from 200mg of  $\text{CO}_2$ ,  $10^{21}$  molecules are removed, how many moles of  $\text{CO}_2$  are left?

$$200 \text{ mg of } \text{CO}_2 = 200 \times 10^{-3} \text{ g of } \text{CO}_2.$$

$$44 \text{ g of } \text{CO}_2 \text{ molecules} = 6.022 \times 10^{23} \text{ molecules.}$$

$$\text{for } 1 \text{ g} = \left( \frac{6.022 \times 10^{23}}{44} \right)$$

$$\text{for } 200 \text{ mg} = \left( \frac{6.022 \times 10^{23}}{44} \times 200 \times 1000 \right)$$

$$\text{No of molecules} = \frac{6.022 \times 10^{28}}{22} - 10^{21} = 2.737 \times 10^{27}$$

no of moles =

$$\text{given weight} = 200 \text{ mg of } \text{CO}_2 = 0.2 \text{ g.}$$

$$\text{Gram molecular mass of } \text{CO}_2 = 44 \text{ g.}$$

$$\text{Mass of 1 molecule of } \text{CO}_2 = \frac{44}{6.022 \times 10^{23}}$$

$$\text{Mass of } 10^{21} \text{ mole} = \frac{44}{6.022 \times 10^{23}} \times 10^{21} = 0.07$$

$$\text{Mass of CO}_2 \text{ left} = (0.2 \text{ g} - 0.073 \text{ g}) \\ = 0.127 \text{ g.}$$

$$\text{No of moles of CO}_2 \text{ left} \\ = \left( \frac{0.127}{44} \right) =$$

1) How many molecules & atoms of oxygen present in 5.6 litres of  $\text{O}_2$  at NTP?

$$22.4 \text{ L} \rightarrow 6.022 \times 10^{23} \text{ Molecules}$$

$$5.6 \text{ L} = \left( \frac{6.022 \times 10^{23}}{22.4} \times 5.6 \right) = 1.505 \times 10^{23} \text{ molecules}$$

No of atoms in 5.6 Litres

$$1 \text{ molecule} = \text{O}_2 = 2 \text{ atom.}$$

$$1.505 \times 10^{23} \text{ molecules} = 2 \times 1.505 \times 10^{23} \\ = 3.01 \times 10^{23} \text{ atoms.}$$

~~Solve~~

How many electrons are present in 1.6 g of methane.

$$W = 1.6 \text{ g.}$$

$$\text{Gr-MM of CH}_4 = 12 + 4 \times 1 = 16 \text{ g.}$$

$$\text{No of moles} = \frac{1.6}{16}$$

$$16 \text{ g} \rightarrow 1 \text{ mole}$$

$$1.6 \text{ g} \rightarrow \frac{1}{16} \times 1.6 = 0.1 \text{ mole.}$$

No of methane molecules

$$= 0.1 \times 6.022 \times 10^{23} = 6.022 \times 10^{22}$$

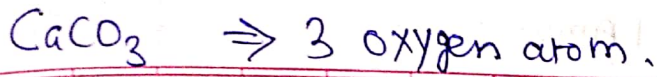
$$1 \text{ molecule of CH}_4 = 4 + 6 = 10 \text{ electrons.}$$

$$6.022 \times 10^{22} \text{ Molecules} = 6.022 \times 10^{23} = N_A \text{ electrons.}$$

2. Calculate no of moles in 25g of  $\text{CaCO}_3$  of oxygen atoms.

$$\text{Formulae mass of CaCO}_3 = 100.$$

$$\text{No of moles of CaCO}_3 = \frac{\text{Mass (in gram)}}{\text{Formulae Mass}}$$

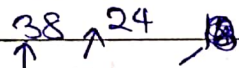


$= \frac{25}{100} = 0.25 \text{ mole.}$

No of oxygen atom in 1 mole =  $3 \text{ NA}$ .

for 0.25 mole =  $0.25 \times 3 \times \text{NA}$

Atomic Nm (z



32 Calculate total no of atoms in 0.5 mol of  $\text{K}_2\text{Cr}_2\text{O}_7$ .

Molecular mass =

15 Total nm of electrons in 1.4 g Nitrogen gas:-

$\text{N}_2 \text{ gas} \Rightarrow 14 \text{ electron.}$

$(14 \text{ g} = 1 \text{ mole})$   
 $1 \text{ g} = \frac{1}{14}$

~~1 mole of  $\text{N}_2 \text{ gas}$~~  No of moles =  $\frac{1.4}{14} = 0.1 \text{ mole.}$

No of molecules in 0.1 mole

$= 0.1 \times 6.022 \times 10^{23} = 6.022 \times 10^{22}$

~~1 mole of molecules =  $6.022 \times 10^{23} \text{ NA}$~~

$M_{\text{proton}} = 1.67 \times 10^{-24} \text{ g} \rightarrow 1 \text{ amu}$   
 $M_{\text{neutron}} = 1.6 \times 10^{-24} \text{ g}$   
 $M_{\text{elec}} = 9.1 \times 10^{-27} \text{ g}$

$m_e = \frac{1}{1837} M_{\text{proton}}$

for 1 C atom  
 $= 6p + 6n + 6e = 6p + 6n$   
 $= 12p = 12 \times 1.67 \times 10^{-24}$

1 atom of C = 12 amu.

1 amu =  $\left(\frac{1}{12}\right)^{\text{th}}$  of mass of 1 atom of C.

$6.022 \times 10^{23}$  atoms of C = 1 mole of C atoms.

mass " " " =  $12 \times 1.67 \times 10^{-24} \times 6.022 \times 10^{23}$   
 $= 12 \text{ g}$  (Molar Mass)  
 $\rightarrow$  Gram Atomic Mass.

Molar Mass  $\rightarrow$  mass of 1 mole of atoms.

$O_2$

$8p + 8n = 16 \times 1.67 \times 10^{-24} \text{ g} = 16 \text{ amu}$

1 mole of  $O_2$  atom =  $16 \times 1.67 \times 10^{-24} \times 6.022 \times 10^{23}$   
 $= 16 \text{ g}$ .

For Molecules

$N_2 = 2 \text{ atoms of N}$  ( $N = 7p + 7n$ )  
 $= 2 \times (7p + 7n)$   $p = n$   
 $\Rightarrow 28p$   
 $\Rightarrow 28 \times 1.67 \times 10^{-24} \text{ g}$   
 $= 28 \text{ amu}$  (Molecular mass)

Q. Find in 98 g of  $H_2SO_4$ .

① mole of  $H_2SO_4$  molecules

G-Molecular mass =  $2 + 32 + 64 = 98 \text{ g}$ .

$n = \frac{\text{Given weight}}{\text{Molar mass}} = 1$

Gram atom  $\rightarrow$  Mole of atom.

Gram molecules  $\rightarrow$  Molecules.

b) mole of S atom

$H_2SO_4$  - 1 S atom ( $\because$  1 Mole of  $H_2SO_4$  = 1 mole of S atom)

$\Rightarrow$  G.M of S = 32 g.

$$n = \frac{W}{G.M} = \frac{98g}{32g} = 1.$$

c) no of S atom =  $N_A$  atom.

d) mole of H atom = 2 mole of H atoms

e) Gram atom/Mole of H = 2 moles.

f) Mass of S atom

$$\Rightarrow 1 \text{ Mole} = 32 \text{ g.}$$

Q. Find the no of atoms of each type - present in 3.42g of Sugar-Cane  $\rightarrow C_{12}H_{22}O_{11}$ .

$$\text{Gram-Molecular Mass} = 144 + 22 + 176 = 342 \text{ g.}$$

$$\text{no of moles} = \frac{3.42}{342} = 0.01 \text{ mole.}$$

1 mole of  $C_{12}H_{22}O_{11}$  contains - containing 12 C atom, 22 H atom & 11 O atom.

In 0.01 mole

$$\text{Carbon atom} = 0.01 \times 12 \times 6.022 \times 10^{23} = 7.22 \times 10^{22}$$

$$\text{H atom} = 22 \times 6.022 \times 10^{23} \times 0.01$$

$$\text{O atom} = 11 \times 0.01 \times 6.022 \times 10^{23}$$

Q. How many atoms and molecules are present in 64g of Sulphur  $S_8$ .

$$n = \frac{64}{32 \times 8} = 0.25 \text{ mole.}$$

$$\text{No of atom} = 8 \times 0.25 = 2 N_A$$

$$\text{No of molecules of } S_8 = 0.25 \times N_A$$

# Methods of Expressing Conc of Sol<sup>n</sup>.

① Mass percentage by Mass

$$\Rightarrow \left( \frac{\text{Mass of solute}}{\text{Mass of Solution}} \times 100 \right) = \left( \frac{\text{Mass of solute}}{\text{Mass of solvent} + \text{Mass of solute}} \right) \times 100$$

↓  
 $\frac{\text{Mass of solute}}{(V \times \rho) \text{ sol}^n} \times 100$       • Solid-Solid  
→ Alloys.

Mass of solute = Mass fraction. a  
Mass fraction

10% of sugar → 10 g of sugar present in 100 gram of sol<sup>n</sup>.

② Percent by Vol = Vol of solute present in 100 mL of sol<sup>n</sup>

$$= \left( \frac{\text{Volume of Solute}}{\text{Vol}^n \text{ of Sol}^n} \times 100 \right). \quad (\text{liq} - \text{liq}).$$

✓ ③ Percent Mass by Vol (Solid - liq)

$$\Rightarrow \left( \frac{\text{Mass of solute}}{\text{Vol of sol}^n} \right) \times 100$$

④ Strength of conc (Gram per litre) → Amount of solute in gram present in 1 L of sol<sup>n</sup>.

$$= \frac{\text{Mass of solute (in gram)}}{\text{Vol of sol}^n \text{ (in litre)}} \quad \frac{W}{V} \text{ g l}^{-1}.$$

⑤ (PPM) =  $\frac{\text{Mass of solute}}{\text{Mass of sol}^n} \times 10^6$

Volume of pollutant in 10<sup>6</sup> units of volume.  
Ex 10 ppm of SO<sub>2</sub> in 10<sup>6</sup> mL of Air.

⑦ Mole-fraction - It is used when sol<sup>n</sup> is constituted by mixing two or more components.

$$X = \left( \frac{\text{No of moles of one component}}{\text{Total nm of moles of sol}^n} \right)$$

Components	A	B	C
Mass (in gram)	w <sub>1</sub>	w <sub>2</sub>	w <sub>3</sub>
Molecular Mass	m <sub>1</sub>	m <sub>2</sub>	m <sub>3</sub>

No of g moles =  $\frac{w_1}{m_1}$      $\frac{w_2}{m_2}$      $\frac{w_3}{m_3}$

Total no of g moles =  $\frac{w_1}{m_1} + \frac{w_2}{m_2} + \frac{w_3}{m_3}$

~~$\frac{w_1/m_1}{(w_1/m_1 + w_2/m_2 + w_3/m_3)} = f_A$~~      ~~$\frac{w_2/m_2}{(w_1/m_1 + w_2/m_2 + w_3/m_3)}$~~      ~~$\frac{w_3/m_3}{(w_1/m_1 + w_2/m_2 + w_3/m_3)}$~~

~~$x_A + x_B + x_C = 1$~~      $f_A + f_B + f_C = 1$

for binary soln

Mole-fraction of solute + Mole-fraction of solvent = 1.

Molality (m) =  $\left( \frac{\text{No of mole of solute}}{\text{No of kg of solvent}} \right)$

$x_{\text{solute}} = \frac{n_B}{n_A + n_B} = x_B$

$m = \left( \frac{w_B}{m_B \times W_A} \times 1000 \right)$

$x_{\text{solvent}} = \left( \frac{n_A}{n_A + n_B} \right) = x_A$

let  $w_B$  gram of solute of molecular mass  $m_B$  present in  $W_A$  grams of solvent.     $x_A + x_B = 1$

$m = \left( \frac{w_B}{m_B \times W_B} \right) \times 1000$      $n_A = \left( \frac{w_B}{m_B} \right)$

→ Independent of temp variation.

$m = \left( \frac{n_{\text{solute}}}{W_{\text{solvent}}} \right)$

Molarity (Molar concentration) = No of mole of solute <sup>per</sup> <sub>per</sub> litre of soln.

$w_B \rightarrow$  g of solute of mass  $m_B$  - dissolved in  $V$  litre of soln.

Molarity =  $\left( \frac{w_B/m_B}{V} \right)$

Molarity  $\times m_B = \frac{w_B}{V}$

Molarity  $\times m_B =$  Strength of soln.

→ It depend upon temp.

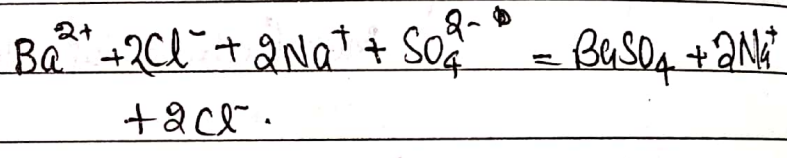
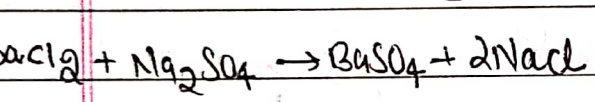
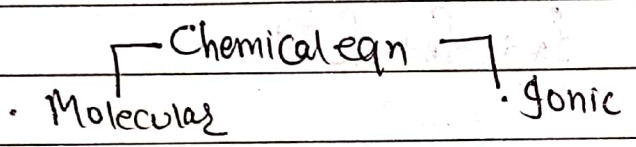
Avg / relative <sup>Mass</sup> Atomic of an element

$$\% \left( \frac{\text{Abundance} \times \text{Atomic Mass}}{100} \right)$$

↓  
Isotopic mass.

Chemical eqn.

Rules - always written in molecular form.  
 ↑ → gaseous product - evolve



Balancing

— Law of conservation - of mass → Dalton theory.

- Hit & trial
- Oxidation nm method.
- Ion-electron method.