Some Basic Concept of Chemistry

Super Short Tricky Chemistry By
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Conceptual Notes for NEET/JEE/Boards

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Introduction: **Chemistry** is the branch of science that deals with the composition, structure & properties of the matter under different condition of temperature & pressure.

➢ **Importance of Chemistry:** Chemistry has a direct impact on our life & has wide range of applications in diff. fields.

1. In Agriculture and Food:

(i) It has provided chemical fertilizers such as urea, calcium phosphate, sodium nitrate, ammonium phosphate etc.

(ii) It has helped to protect the crops from insects & harmful bacteria, by the use of certain effective insecticides & pesticides.

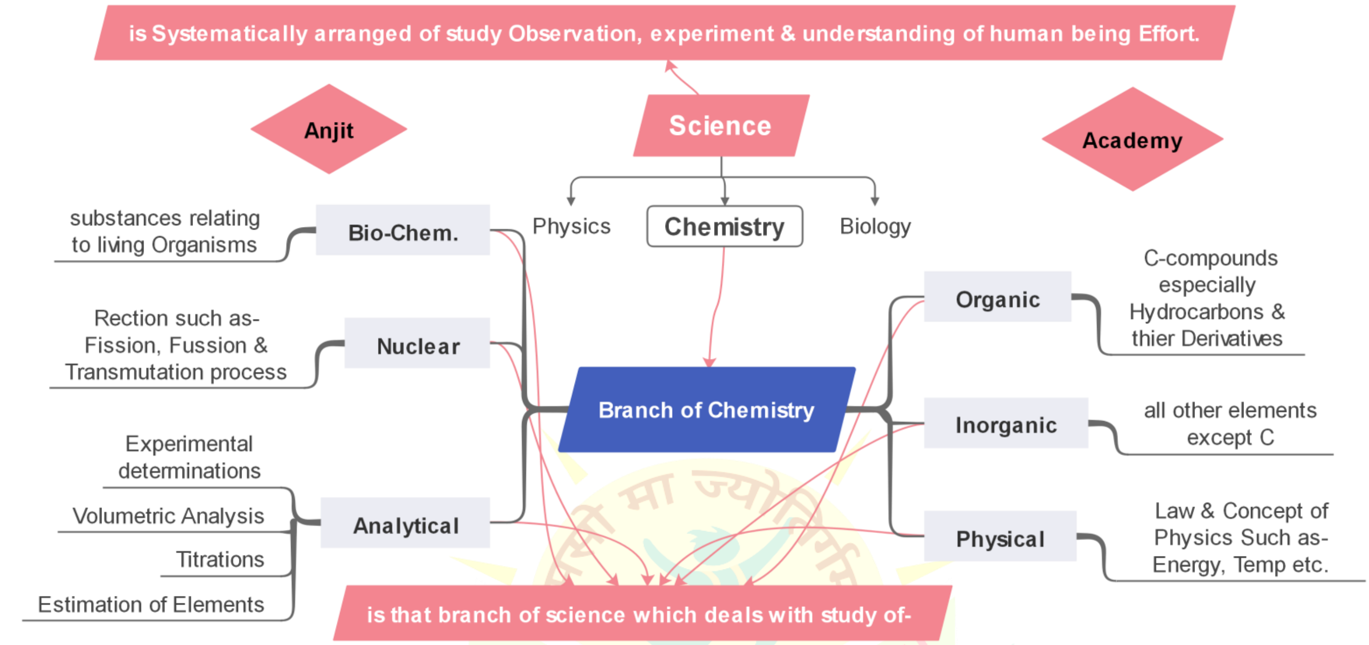
(iii) The use of preservatives has helped to preserve food products like jam, butter, squashes etc. for longer periods.

2. In Health and Sanitation:

(i) It has provided mankind with a large number of life-saving drugs. Today, dysentery and pneumonia are curable due to discovery of sulpha drugs and penicillin life-saving drugs. **Cisplatin** and **Taxol** have been found to be very effective for cancer therapy and **AZT** (Azidothymidine) is used for AIDS victims.

(ii) Disinfectants such as phenol are used to kill the micro-organisms present in drains, toilet, floors etc.

3. Application in Industry: Chemistry has played an important role in developing many industrially manufactured fertilizers, alkalis, acids, salts, dyes, polymers, drugs, soaps, detergents, metal alloys.



➢ Properties of Matter and Their Measurement-

S.N	Physical Change	Chemical Change
1.	No new substance is formed.	A new substance is formed.
2.	Properties of constituent elements/substance is retained.	Properties of constituent elements/substance changes.
3.	Change does not involve loss or gain of heat.	Loss or gain of heat may be involved in this reaction.
4.	This change is generally reversible.	This change is generally irreversible.
5.	For Example - Change of ice into water (melting).	For Example - Burning of coke $C + O_2 \rightarrow CO_2$

■ Physical Classification of Matter :

S.No	Properties	Solid	Liquid	Gas
1.	volume	Definite	Definite	Indefinite
2.	Shape	Definite	Indefinite	Indefinite
3.	I.M.F of attraction	Very high	Moderate	Negligible/Very low
4.	Arrangement of molecules	Orderly arranged	Free to move within the volume	Free to move every where
5.	Inter molecular space	Very small	Slightly greater	Very great
6.	Compressibility	Not compressible	Not compressible	Highly compressible
7.	Expansion on heating	Very little	Very little	Highly expand
8.	Rigidity	Very rigid	Not rigid known as fluid	Not rigid & known as fluid
9.	Fluidity	Can't flow	Can flow	Can flow
10.	Diffusion	They can diffuse due to K.E of liq./gases	Can diffuse And rate of diffusion is very fast	Can diffuse & rate of diffusion is very fast

■ Chemical Classification of matter :

Elements : An element is the simplest form of matter that cannot be split into simpler substances or built from simpler substances by any ordinary chemical or physical method. There are 118 elements known to us.

Compounds : A compound is a pure substance made up of two or more elements combined in a definite proportion by mass. Compounds are broadly classified into **inorganic** and **organic** compounds.

- Inorganic compounds are those, which are obtained from non-living sources such as minerals. For example, common salt, marble and limestone.
- Organic compounds are those, which occur in living sources such as plants and animals. They all contain carbon.

➢ Note : **" All Compounds are molecules , But Not all molecules are compounds. " Ex.-** Molecular hydrogen (H_2) & molecular Oxygen (O_2) are not compound b/coz each is composed of single element , But H_2O is compound.

➢ Differences between Compounds and Mixtures :

Compounds	Mixtures
1. In which, two or more elements are combined chemically.	1. In which, simply mixed and not combined chemically.
2. The elements are present in the fixed ratio by mass. This ratio cannot change.	2. Constituents are not present in fixed ratio. It can vary
3. Compounds are always homogeneous i.e., they have the same composition throughout.	3. It may be either homogeneous or heterogeneous in nature.
4 . Its constituents cannot be separated by physical methods	4. its Constituents can be separated by physical methods.
5. The constituents lose their identities .	5. The constituents do not lose their identities .

■ Mass and Weight :

Properties	Mass	Weight
1. meaning	Quantity of matter, contained in body.	Force applied on object due to the pull of gravity. $W = F_{external} = mg$
2. unit	Kg & Scalar	Newton ($1N = 1\ Kgm/s^2$) & Vector
3. Physical quantity	It remain same in any Location & Time.	Depend on location

Valancy of Cation & Anion : (a) Cations :

VALANCY – 1	
Ammonium	NH_4^+
Sodium	Na^+
Potassium	K^+
Rubidium	Rb^+
Cesium	Cs^+
Silver	Ag^+
Copper (Cuprous)	Cu^+
Gold (Aurous)	Au^+

VALANCY – 2	
Magnesium	Mg^{2+}
Calcium	Ca^{2+}
Stroncium	Sr^{2+}
Barium	Ba^{2+}
Zinc	Zn^{2+}
Cadmium	Cd^{2+}
Nickel	Ni^{2+}
Copper (Cupric)	Cu^{2+}
Mercury (Mercuric)	Hg^{2+}
Lead (Plumbus)	Pb^{2+}
Tin (Stannous)	Sn^{2+}
Iron (Ferrous)	Fe^{2+}

VALANCY – 3	
Iron (Ferric)	Fe^{3+}
Aluminium	Al^{3+}
Chromium	Cr^{3+}
Gold (Auric)	Au^{3+}

Inorganic Naming

Higher O.S ← Anjit → Lower O.S

-ate | -ic		-ite | -ous	
NO_3^-	Nitrate	NO_2^-	Nitrite
SO_4^{2-}	Sulphate	SO_3^{2-}	Sulphite
PO_4^{3-}	Phosphate	HPO_3^{2-}	Phosphite
CuO	Cupric Oxide	Cu_2O	Cuprous Oxide
H_2SO_4	Sulphuric acid	H_2SO_3	Sulphurous acid

(b) Anions :

VALANCY – 1	
Hydroxide	OH^-
Nitrate	NO_3^-
Nitrite	NO_2^-
Permanganate	MnO_4^-
Bisulphite	HSO_3^-
Bisulphate	HSO_4^-
Bicarbonate (Hydrogen carbonate)	HCO_3^-
Dihydrogen phosphate	$H_2PO_4^-$
Perchlorate	ClO_4^-
Chlorate	ClO_3^-
Chlorite	ClO_2^-
Hypochlorite	ClO^-
Iodate	IO_3^-
Meta aluminate	AlO_2^-
Meta borate	BO_2^-
Cyanide	CN^- $[:C\equiv N:]^-$
Isocyanide	NC^- $[^-\ddot{N}\equiv C:]$
Cyanate	CNO^- $[:\ddot{O}–C\equiv N]:$
Isocyanate	$[O=C=\ddot{N}^-]$
Thiocyanate	SCN^-
Formate	$HCOO^-$

VALANCY – 2	
Carbonate	CO_3^{2-}
Sulphate	SO_4^{2-}
Sulphite	SO_3^{2-}
Sulphide	S^{2-}
Thiosulphate	$S_2O_3^{2-}$
Tetrathionate	$S_4O_6^{2-}$
Oxalate	$C_2O_4^{2-}$
Silicate	SiO_3^{2-}
Hydrogen phosphate	HPO_4^{2-}
Manganate	MnO_4^{2-}
Chromate	CrO_4^{2-}
Dichromate	$Cr_2O_7^{2-}$
Zincate	ZnO_2^{2-}
Stannate	SnO_3^{2-}
Hexaflurosilicate (or silicofluorides)	SiF_6^{2-}
Tartrates	$C_4H_4O_6^{2-}$
Phosphite	HPO_3^{2-}
Chromate	CrO_4^{2-}
Pyroborate	$B_4O_7^{2-}$

VALANCY – 3	
Hexacyano ferrate (III) or Ferricynide	$[Fe(CN)_6]^{3-}$
Phosphate	PO_4^{3-}
Borate (orthoborate)	BO_3^{3-}
Arsenate	AsO_4^{3-}
Arsenite	AsO_3^{3-}
Nitride	N^{3-}
Phosphide	P^{3-}

VALANCY – 4	
Pyrophosphate	$P_2O_7^{4-}$
Hexacyano ferrate(II) or Ferrocyanide	$[Fe(CN)_6]^{4-}$

Units of Measurement : Fundamental Units : The quantities mass, length and time are called fundamental quantities and their units are known as fundamental units. There are **seven** basic units of measurement for the quantities :

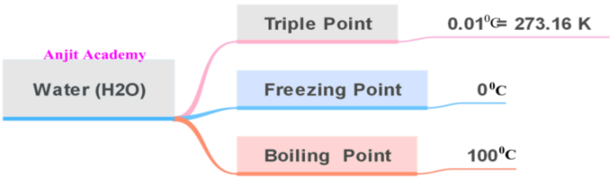
➢ Dimensional Analysis:

1. Volume : Volume has the units of (length)^3. So Volume has units of m^3 or cm^3 or dm^3. A common unit, litre (L) is not an SI unit, is used for measurement of volume of liquids. 1 L = 1000 mL = 1000 cm^3 = 10^3 cc = 1 dm^3

2. Density: Density of a substance is its amount of mass per unit volume.SI unit of density = SI unit of mass/SI unit of volume = kg/m^3 or kg m^{-3} This unit is quite large and a chemist often expresses density in g cm^{-3}.

3. Temperature: There are three common scales to measure temperature — °C (degree Celsius), °F (degree Fahrenheit) & K (kelvin). Here, K is the SI unit. K = °C + 273.15 & °F = 9/5 x °C + 32

➢ **Note—**Temp. below 0 °C (i.e. negative values) are possible in Celsius scale but in Kelvin scale, negative temp. is not possible.



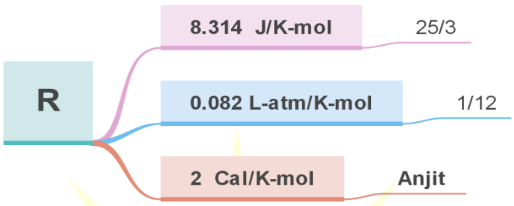
4. Pressure : 1 atm = 1.013 x 10^5 N/m^2 **or** Pascal's = 760 mm Hg = 76 cm Hg = 760 torr **[1 bar =** 10^5 Pascal's **]**

5. Triple Point : Three phase of matter (S,L,G) co-exit at the same point is called Triple point.

➢ Some Important Conversion :

1. 1 Cal. = 4.184 J = 4.18 x 10^7 erg
2. 1 inch = 2.54 cm = 25.4 mm
3. 1 cm = 10 mm = 10^{-2} m
4. 1 N = 10^5 dyne
5. 1 J = 10^7 erg
6. 1 ev = 1.6 x 10^{-19} C
7. 1esu = 3.33 x 10^{-10} C
8. 1 Pm = 10^{-12} m = 10^{-10} cm
9. 1 A^0 = 10^{-10} m = 10^{-8} cm = 0.1nm
10. C = 3 x 10 8 m/sec
11. h = 6.626 x 10^{-34} J-sec = 6.626 x 10^{-27} erg-sec
12. F = 96,500 C
13. N_A = 6.023 x 10^{23} /mol
14. 1 amu = 1.66 x 10^{-27} Kg
15. **charge of e$^-$** = 1.6 x 10^{-19} C = 4.8 x 10^{-10} esu

1. Value of **R**



Standard Measurments: Set by IUPAC as-

(i) STP : Standard Temp & Pr.
T = 0^{0}C , P = 1 bar, V = 22.4 L

(ii) NTP : Normal Temp & Pr
T = 20^{0}C, P = 1 atm, V = 24.05 L

(iii) SATP : Standard Ambient Temp & Pr.
T = 25^{0}C, P = 1 atm, V= 24.46 L

where, V = Vol. of 1 mole gas ∵ 1 atm = 1.013 bar

➢ Useful relation :

2. ln X = 2.303 $\log_{10}$ X
3. 1 L-atm = 24.21 Cal. = 101.32 J
4. 1' = 12"
5. π = 3.141
6. e = 2.71
7. V(STP) = 22.4 L

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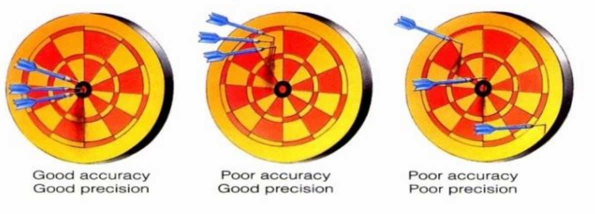
log(a/b) = log a - log b
log(a x b) = log a + log b
$\log_{10}10^x = x$
$\log_{10}(x^n) = n \log_{10} x$
$\log_{10}(1/a) = -\log_{10} a$
ln → natural logaritham base **e**

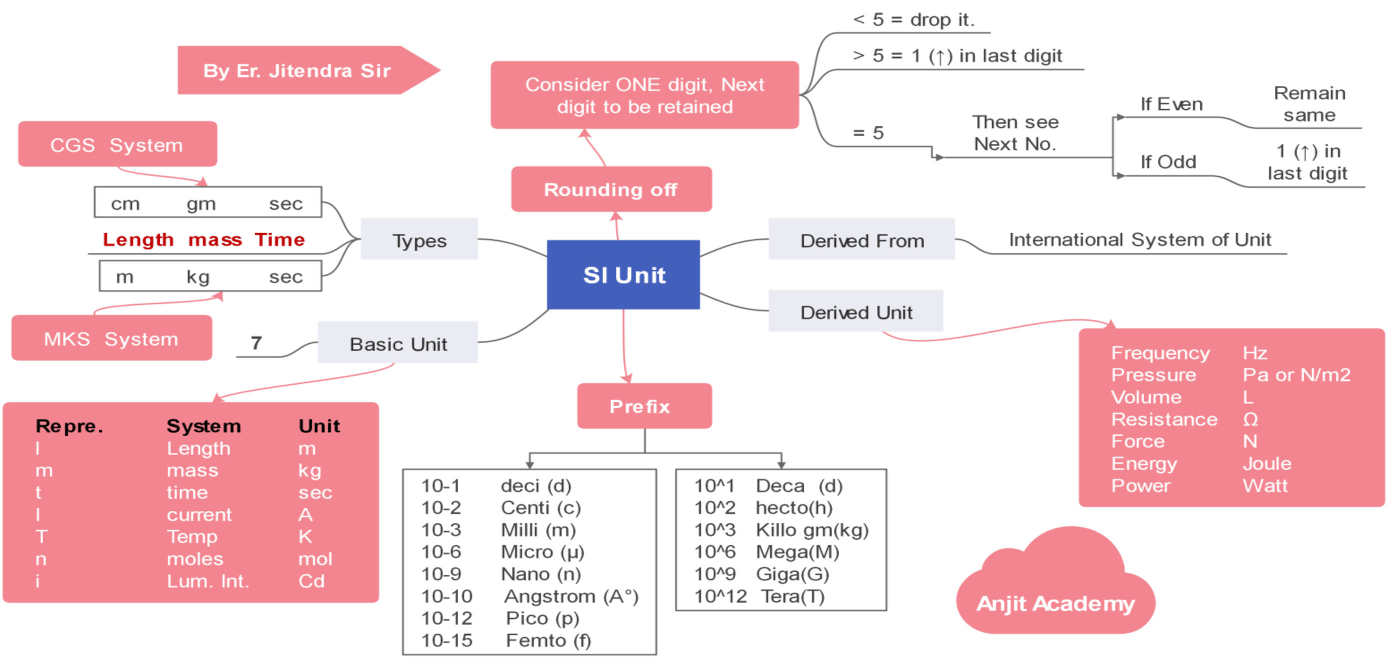
➢ **Atomicity :** is the no. of **gaseous atoms** of an element present in ONE molecule. **Ex.-** He → 1 H_2 → 2 O_3 → 3 But, **S_6 & I_2** → 0 b/coz it is exit in solid form.

➢ Precision & Accuracy:

Precision refers to the closeness of various measurements for the same quantity while **accuracy** refers to the agreement of a particular value to the true value of the result.



The International System of Units (SI) :



➢ Retention of Significant Figures - Rounding off Figures :

The rounding off procedure is applied to retain the required number of significant figures.

Example - 4.317 is rounded off to 4.32 ; 4.312 is rounded off to 4.31 8.375 is rounded off to 8.38 ; 8.365 is rounded off to 8.36

■ Scientific Notation :

Numbers are represented in $N \times 10^n$ form. Where, N = Digit term n = exponent having positive or negative value.

Examples- $12540000 = 1.254 \times 10^7$ & $0.00456 = 4.56 \times 10^{-3}$

Mathematical Operations of Scientific Notation:

Multiplication and Division: Example: $(7.0 \times 10^3) \times (8.0\times10^{-7}) = (7.0\times8.0) \times (10^{[3+(-7)]}) = 56.0 \times 10^{-4}$

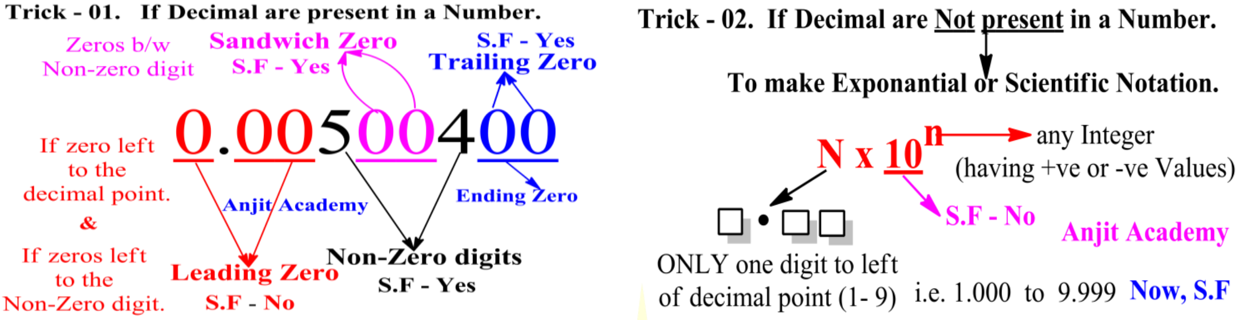
$(7.0 \times 10^3) / (8.0\times10^{-7}) = (7.0/8.0) \times (10^{[3-(-7)]}) = 0.875 \times 10^{10} = 0.9 \times 10^{10}$

Addition and Subtraction: Example- $(5 \times 10^3) + (8\times10^5) = (5 \times 10^3) + (800\times10^3) = (5+800) \times 10^3 = 805\times10^3$

Significant Figures : Total no. of digit in a number including last digit whose value is uncertain is called significant figures.

Ex.- dia. of measurement d = 56 ± 0.001 cm

There are certain rules for determining the number of significant figures. These are -



Trick - 03. Counting no. (fixed) = S.F → ∞ **Ex.-** 1 dozen = 12 or Kg = 1000g or $\pi = 3.14$ all have **S.F → ∞**

➢ **Note :** The decimal point does not count towards the number of significant figures.

Example- the number 345601 has six significant figures but can be written in different ways, as 345.601 or 0.345601 or 3.45601 all having same number of significant figures.

■ Addition & Subtraction of Significant Figures:

In addition or subtraction of the numbers having different precisions, the final result should be reported to the same number of decimal places as in the term having the least number of decimal places. For example: 3.52 + 2.3 + 6.24 = **12.06**

The final result has two decimal places but the answer has to be reported only up to one decimal place, i.e. the **ans** would be **12.0**

Subtraction of numbers can be done in the same way as the addition. for Example- 23.4730 - 12.11 = **11.3630**

The final result has four decimal places. But it has to be reported only up to two decimal places, i.e., the answer would be **11.36**

■ Multiplication & Division of Significant Figures :

In the multiplication or division, the final result should be reported upto the same number of significant figures as present in the least precise number.

Multiplication of Numbers : 2.2120 x 0.011 = **0.024332** ; According to the rule the final result = **0.024**

Division of Numbers : 4.2211 ÷ 3.76 = **1.12263** ; The correct answer = **1.12**

Laws of Chemical Combinations :

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

1. Law of Conservation of Mass : Given by Antoine Lavoisier in 1789.

It states that matter (mass) can neither be created nor destroyed. **In other words,**

In all physical and chemical changes, the total mass of the reactants is equal to that of the products.

The following experiments illustrate the truth of this law -

(**a**) When matter undergoes a physical change - Ice → Water

It is found that there is no change in weight though a physical change has taken place.

(**b**) When matter undergoes a chemical change -

For example, decomposition of Calcium Carbonate $CaCO_3 \xrightarrow{\Delta} CaO + CO_2$

40+12+48=100g 40+16=56g 12+32 =44g

During the above decomposition reaction, matter is neither gained nor lost.

Limitation : In nuclear reactions, some mass of reactant is converted into energy, so mass of reactant is always less than that of product.

Ex.- 1.7 gram of silver nitrate dissolved in 100 gram of water is taken. 0.585 gram of sodium chloride dissolved in 100 gram of water is added to it and chemical reaction occurs. 1.435 gm of AgCl and 0.85 gm $NaNO_3$ are formed. Show that these results illustrate the law of conservation of mass.

Sol. Total mass before chemical change = mass of $AgNO_3$ + Mass of NaCl + Mass of water = 1.70 + 0.585 + 200 = 202.285 gram

Total mass after the chemical reaction = mass of AgCl + Mass of $NaNO_3$ + Mass of water = 1.435 + 0.85 + 200 = 202.285 gram

Thus, in the given reaction Total mass of reactants =Total mass of the products.

2. Law of Definite Proportions 'or' Law of Constant Proportion : Proposed by Louis Proust in 1799.

which states that - A chemical compound is always contains fixed proportion of element by their weight.

For example, Water may be formed in a number of ways i.e., distilled water / River Water / Ice → H_2O = 2 : 16 = 1: 8 (fixed)

Limitations : In case of isotopes, ratio is not fixed

Ex.- 12_{CO_2} = 12: 32 = 3 : 8 **&** 14_{CO_2} = 14 : 32 = 7 : 16

Q. Irrespective of the source, pure sample of water always yields 88.89% mass of oxygen and 11.11% mass of hydrogen. This is explained by the law of:

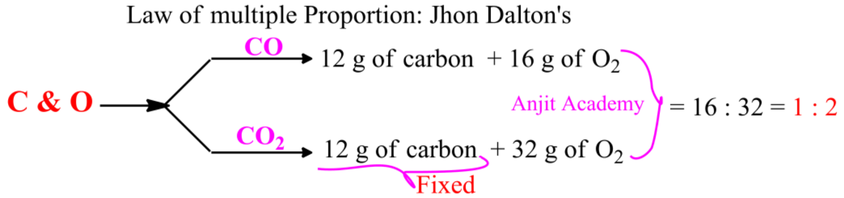
(a.) conservation of mass (b.) constant composition (c.) multiple proportion (d.) constant volume

As in water $\frac{\text{Hydrogen}}{\text{Oxygen}} = \frac{2}{16} \Rightarrow H:O = 1:8$ by mass | So oxygen = $\frac{8}{8+1} \times 100 = 88.89\%$ | Hydrogen = $\frac{1}{8+1} \times 100 = 11.11\%$

Both value always constant. Obey law of constant composition. Hence, (B) is correct.

3. Law of Multiple Proportions : Proposed by Dalton in 1803.

This law states that - When two elements combine together to form two or more chemical compounds, then the one element is fixed by mass of the other, bear a simple ratio to one another.



Limitations :

The law is not applicable, If element Exist in Isotopes. for Ex.- formation of HCl, Cl-isotopes combine Cl^{35} 1:35 & Cl^{37} 1:37 in HCl

Example- Two compounds each containing only tin and oxygen had the following composition. Show that these data illustrate the law of multiple proportion?

	Mass% of Tin	Mass % of oxygen
Compound A	78.77	21.23
Compound B	88.12	11.88

Sol. In compound A : 21.23 parts of oxygen combine with 78.77 parts of tin.

∴ 1 part of oxygen combine with 21.23 / 78.77 = 3.7 parts of Sn.

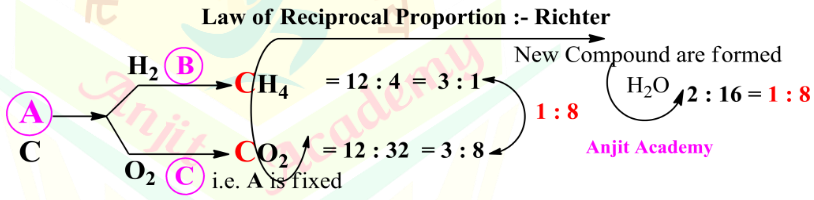
In compound B : 11.88 parts of oxygen combine with 88.12 parts of tin.

∴ 1 part of oxygen combined with 11.88/88.12 = 7.4 parts of tin.

Thus, the mass of Tin in compound A and B which combine with a fixed mass of oxygen are in the ratio 3.7 : 7.4 or 1 : 2. This is a simple ratio. Hence the data illustrate the law of multiple proportion.

4* Law of Reciprocal proportion : Proposed by Richter 1794 .**

If an element 'A' combined together with Two different element 'B' & 'C' to form two different compound in which element 'A' is fixed to others; when 'B' & 'C' combined together to form a new compound.



Ex.- One part of an element A combines with two parts of B (another element). Six parts of element C combine with four parts of element B. If A and C combines together, the ratio of their masses will be governed by:

(a) law of definite proportions (b) law of multiple proportions (c) law of reciprocal proportions (d) law of conservation of mass

Solution: $\frac{B}{A} = \frac{2}{1}; \frac{B}{C} = \frac{4}{6} = \frac{2}{3}$ from this when A combined with C the ratio is A/C = 1/3 Hence, (C) is correct.

5. Gay Lussac's Law of Gaseous Volumes : Given by Gay Lussac in 1808.

He developed Vol.-Vol. relations of gaseous equation. According to this law - " Volume of gases always combined with one another in a simple in a ratio by volume in a particular temperature and pressure".

The law may be illustrated by the following examples - Combination between hydrogen and chlorine:

e.g. $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$

1V 1V 2V All reactants and products have simple ratio 1 : 1 : 2

Ex-01. Ammonia is oxidized according to the following equation : $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$

How many litres of nitric oxide are formed when 90 litres of oxygen reacts with ammonia?

Ans. By Gay-Lussac's Law : $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$

$4NH_3$	$5O_2$	$4NO$
4 vols	5 vols	4 vols
4/5 vols	1 vol	4/5vols
4/5 × 90 lt	90 lt	4/5 × 90 lt
=72 lt	= 90 lt	= 72 lt

∴ The volume of nitric oxide formed at STP = 72 Lt

02. What volume of propane is burnt for every 100 cm3 of oxygen used in the reaction? $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

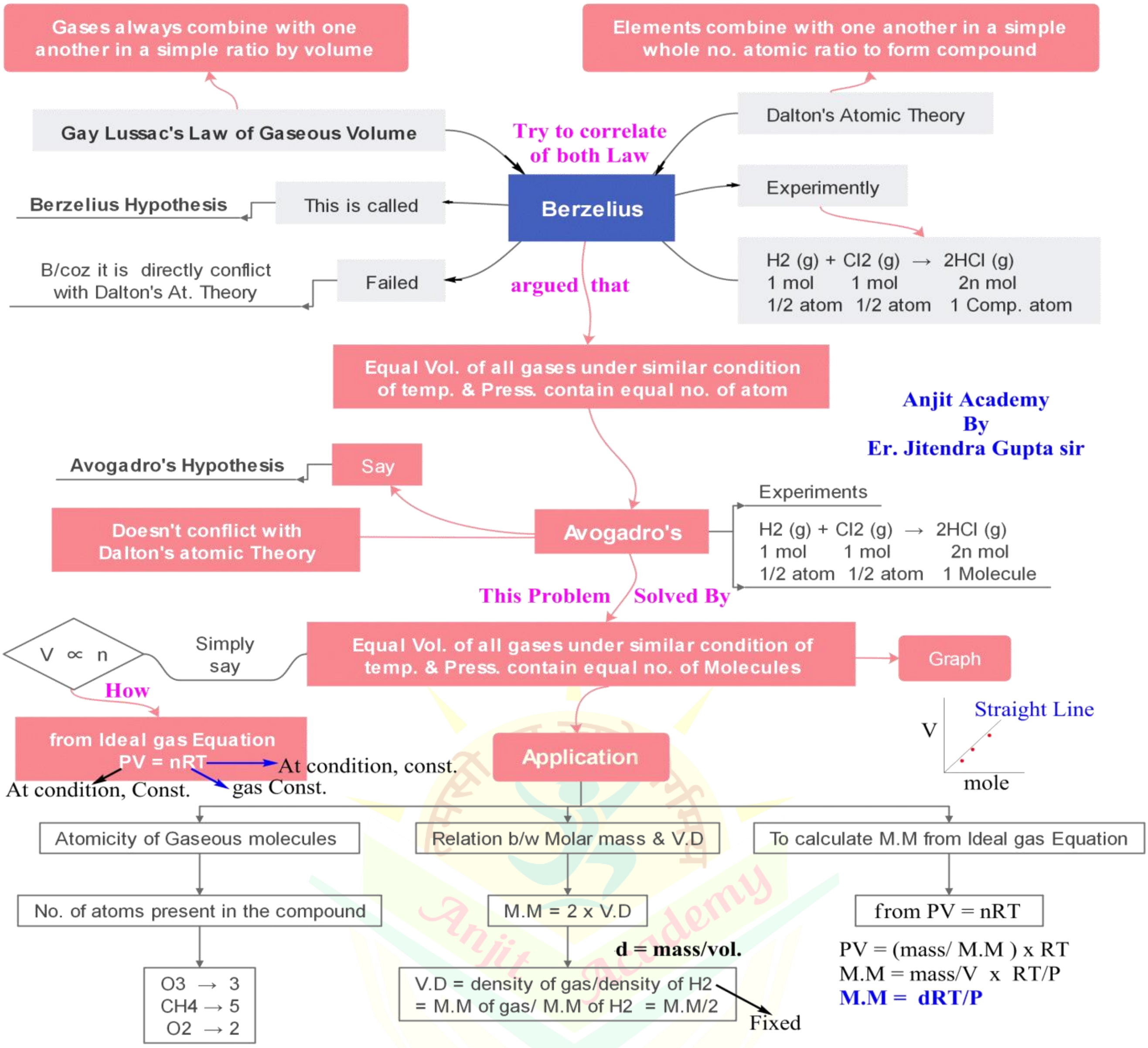
The volumes of gases are measured at room temperature and pressure.

Ans. By Gay-Lussac's law : $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

C_3H_8	$5O_2$	$3CO_2$
1 vol	5 vols	3 vols (div. by 5)
1/5 vol	1 vol	
1/5 × 100 cm3	100 cm3	
= 20 cm3	=100 cm3	

∴ The volume of propane burnt = 20 cm3.

6. Avogadro Law : In 1811, Given by Avogadro. According to this law - Equal volumes of gases at the similar condition of temperature and pressure contain equal number of molecules.



Ex-01. A gas cylinder filled with hydrogen gas holds 5 g of the gas. The same cylinder holds 85 g of gas X under the same temperature and pressure. Calculate : (i) vapour density of gas X (ii) molecular weight of gas X.

Ans. (i) Vapour density of gas X = $\dfrac{wt\ of\ gas\ X}{\text{wt of the same volume of hydrogen}} = \dfrac{85}{5} = 17$

(ii) Molecular weight of gas X = 2 × V.D = 2 × 17 = 34 amu.

✍ Dalton's Atomic Theory : based on the Law of Conservation of mass.

In 1808, Dalton published 'A New System of Chemical Philosophy' in which he proposed the following:

S.N	Main point of this Theory	Drawback's of this theory - It could not explain -	Modified Dalton's Atomic Theory
1.	Atom is made up of small individual particles.	Gay Lussac's law of gaseous volumes.	The element having same atomic no. but diff. mass no. (Isotope)
2.	Atom of same element to form shape, size & orientation.	Why ? different element to form different shape & size.	The element having same mass no. but diff. atomic no. (Isobar)
3.	Atom of different element to form diff. shape, size & orientation.	What is the force b/w atom & molecules in three states of matter ? S, L, G	Ratio of atom is fixed & integral, but may not be simple. e.g. In sugar molecules $C_6H_{22}O_{11}$, the ratio of C, H and O atoms is 12:22:11 which is not simple.
4.	Atom is always in whole no. ratio.		Atom is the smallest unit which takes part in chemical reactions.
5.	Atom can neither be created nor be destroyed.		Atom is no longer indivisible.

✍ Mole Concept :***

1. Atomic Mass : Atomic mass of an element is defined as the average relative mass of an atom of an element as compared to the mass of an atom of carbon -12 taken as 12.

■ **NOTE:** You Know : **A ≈ 2 Z**

H = 1	S = 32	Ca = 40	K = 39	N = 14

C = 12 O = 16	P = 31 Cl = 35.5	Ar = 40 Mg = 24	Ag = 108 Cu = 63.5	Na = 23 Al = 27

■ **u (Unified mass)** OR **amu (atomic mass unit) :** One atomic mass unit (amu) is equal to l/12th of the mass of an atom of carbon-12 isotope.

1 amu = $\frac{1}{12} \times \frac{12\ g}{NA}$ = 1/N_A =Reciprocal of Avogadro's no. = $\frac{1}{6.023 \times 10^{23}}$ = 1.67 x 10^{-24} g =mass of proton =1.67 x 10^{-27} kg

■ **Gram Atomic Mass :** Atomic mass of an element expressed in grams is the gram atomic mass or gram atom.
For example, the atomic mass of oxygen = 16 amu ; Therefore gram atomic mass of oxygen = 16 g

■ **Average Atomic Mass :** Most of the elements exist as isotopes ; it may be defined as the average relative mass of an atom of an element as compared to the mass of carbon atoms (C-12) taken as 12.

2. Molecular Mass : Molecular mass of a substance is defined as the average relative mass of its molecule as compared to the mass of an atom of C-12 taken as 12.

It is obtained by multiplying the atomic mass of each element by number of its atoms and adding them together.
For example, Molecular mass of methane (CH_4) = 12.011 u + 4 (1.008 u) = 16.043 u

■ **Gram Molecular Mass :** Molecular mass of a substance expressed in grams is called gram molecular mass.
e.g., the molecular mass of oxygen = 32 amu ; Therefore, gram molecular mass of oxygen = 32 g

3. Formula Mass : Sum of atomic masses of the elements present in one formula unit of a compound. It is used for the ionic compounds. Thus, formula mass of such as NaCl, KNO_3, Na_2CO_3 etc. Ex- NaCl = 23.0 u + 35.5 u = 58.5 u.

What is the mass of
- one atomof Cl ? = 35.5 amu.
- the atomic mass of Cl ? = $\frac{\text{Mass of an atom in amu}}{1\ \text{amu}} = \frac{35.5\ \text{amu}}{1\ \text{amu}} = 35.5$
- the gram atomic mass of Cl ? = [Mass of 1 Cl atom × NA] = 35.5 amu × NA = $\frac{35.5 \times NA\ \text{gram}}{NA}$ = 35.5 gram

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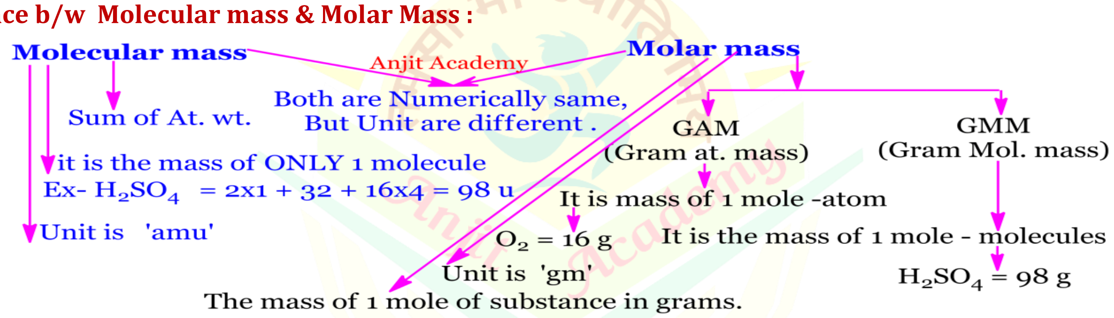
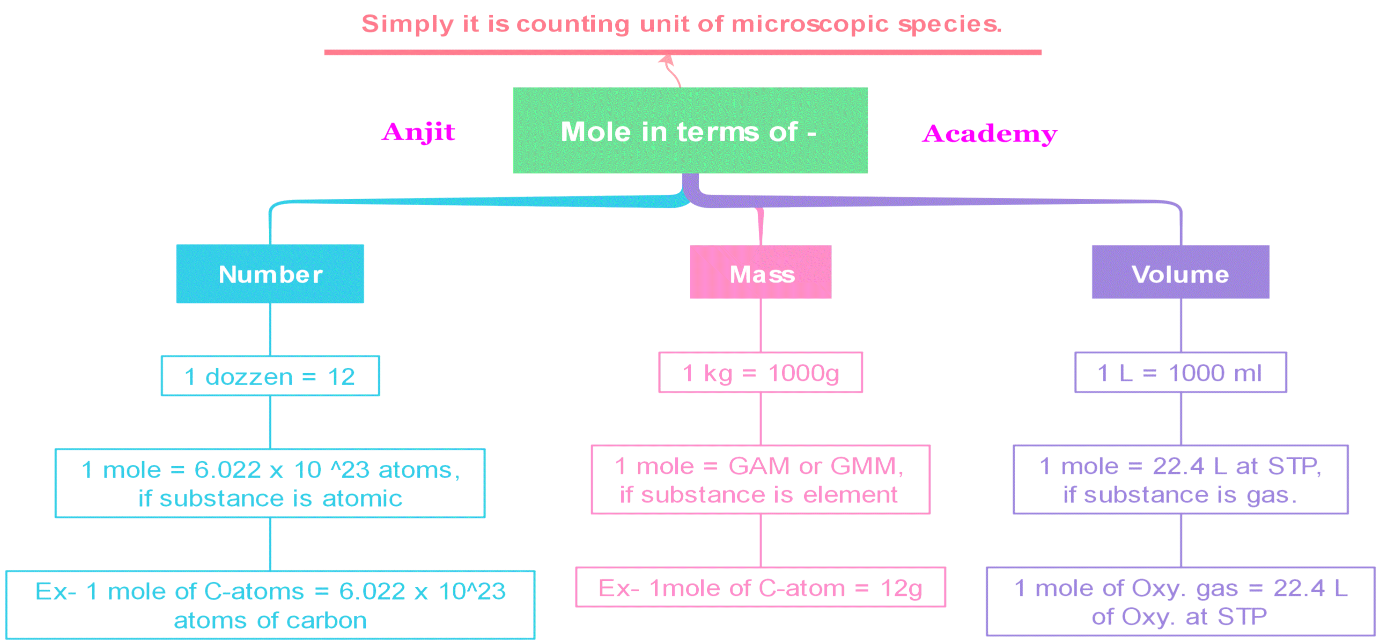
Difference b/w Molecular mass & Molar Mass :



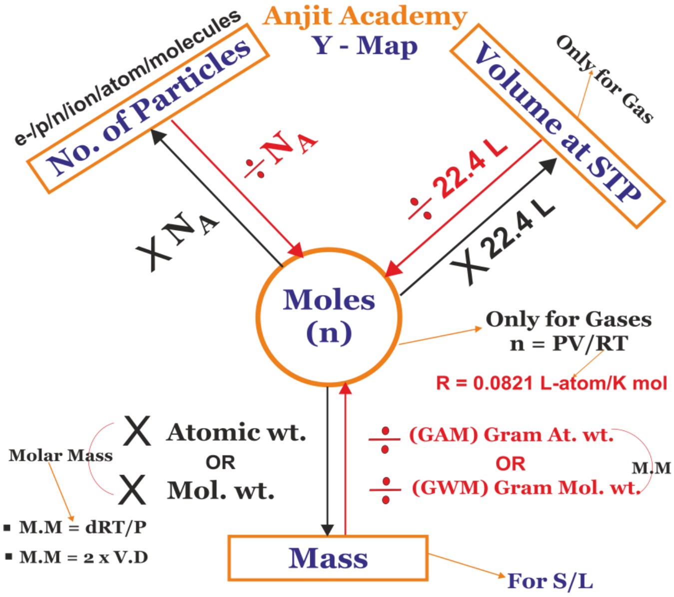
4. Mole : Mole is defined as the amount of a substance, which contains the same number of chemical units (atoms, molecules, ions or electrons) as there are atoms in exactly 12 grams of pure carbon-12.
A mole represents a collection of 6.022 x10^{23} (Avogadro's number) chemical units.. **The mass of one mole of a substance in grams is called its molar mass.**

■ **Mole can be defined in three ways-** We cannot measured of very small species entity like - e-/atom/molecules. So, need to measure physical properties that number is mole.



Note:
We can use the following relationship as per requirement of question.

$$\text{No. of mole} = \frac{\text{No. of particle}}{N_A} = \frac{\text{mass (gm)}}{[\text{gm at. or mol. mass}]} = \frac{V(l)\text{ occupied by a Gas at STP}}{22.7} = \frac{V(l)\text{ occupied by a Gas at 1atm and 273K}}{22.4}$$



Ex.- How many carbon atoms are present in 0.35 mol of $C_6H_{12}O_6$ -

Sol. 1 mol of $C_6H_{12}O_6$ has = 6 N_A atoms of C

∴ 0.35 mol of $C_6H_{12}O_6$ has = 6 × 0.35 N_A atoms of C

= 2.1 N_A atoms = 2.1 × 6.022 × 10^{23} = 1.26 × 10^{24} carbon atoms

Q. Total no. of valance electron in 4.2 g of N_3^- ion is

Total no. of **valance e-** = $\left(\frac{4.2}{42} \times N_A\right) \times 16$

Total no. of **e-** = $\left(\frac{4.2}{42} \times N_A\right) \times 22$

Concentration of Solutions (solute + solvent = Solution) : The concentration (amount of solute dissolve in per litre of sol^n) of a solution can be expressed in any of the following ways-

1. Strength (g/L) : strength = $\frac{\text{mass of solute in (g)}}{\text{Vol.of solution in L}}$

2. Percentage (%) : of compound is the relative mass of each of the constituents element in 100 parts of it.

Mass by Mass percent (w/w %) = $\frac{\text{Mass of Solute in gm}}{\text{Total mass of Solution}} \times 100$

amount of solute in g present in 100 g of solution.

Mass by Vol. percentage(w/V %) = $\frac{\text{Mass of Solute in gm}}{\text{Total Volume of Solution in ml}} \times 100$

Volume by Volume percent (V/V %) = $\frac{\text{Volume of Solute in ml}}{\text{Total volume of Solution in ml}} \times 100$

volume of the solute per 100 units of the volume of solution.

➢ **NOTE :** When V/V % → Vol. of Solution (Let always V= 100ml)

3. Mole fraction : Temp. Independent .

It is the ratio of number of moles of a particular component to the total number of moles of the solution.

For a solution containing n_2 moles of the **solute** dissolved in n_1 moles of the **solvent**,

$$n_2 = \frac{n2}{n2 + n1} \quad \& \quad n_1 = \frac{n1}{n1 + n2} \quad \text{from both equation, } n_1 + n_2 = 1$$

Hence, Sum of mole fraction is always equal to unity.

4. Parts per million (ppm) : The amount of solute in gram per million (10^6) gram of the solution.

ppm = $\frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6$ used in measure Atmospheric Pollution .

5. Parts per billion (ppb) : ppb = $\frac{\text{mass of solute}}{\text{mass of solution}} \times 10^9$

6. Molarity (M) → (mol/L) : It is defined as the number of moles of solute in 1 litre of the solution.

Measurements in Molarity can change with the change in temperature because solutions expand or contract accordingly; i.e Temp. dependent.

$$M = \frac{\text{no.of moles of solute}}{\text{Volume of solution (L)}} = \frac{n}{V(L)} = \frac{\text{wt.of solute} \times 1000}{\text{mol.wt.of solute} \times V(ml)\text{of solution}} = \frac{d\left(\frac{g}{ml}\right)\text{of solution} \times 1000}{\text{Mol.wt. of solute} \times \text{mass of solution (g)}}$$

$$= \frac{\text{Strength (g/L)}}{\text{Mol.mass (g/mol)}}$$

➢ **Note:**

Molar solution → **1M** i.e. Molarity is 1. | Semi molar → M/2 | Deci molar → M/10 | Centi molar → M/100

7. Molality (m) → (mol/kg) : It is defined as the number of moles of solute present in 1 kg of solvent.

It is Temp. independent. m = $\frac{\text{no.of moles of solute}}{\text{mass of soolvent in kg}} = \frac{\text{mass of solute(g)} \times 1000}{\text{mol.mass of solute (g)} \times \text{mass of Solvent (g)}}$

9. Formality (f) : formality is use similar to Molarity ; Use in case of ionic compound.

" no. of formoles of solute in per litre vol. of solution." f = $\frac{\text{mass of solute}}{\text{formula mass of solute} \times V(L)} = \frac{w}{M\text{ obs.} \times V(L)}$

$M_{obs.}$ → Observed Molecular mass or Experimentally Observed M.M (Used in Von't Hoff Factor - ***i***)

➢ **Relation b/w formality & Molarity :**

$\frac{f}{M} = \frac{M.M\,(Thero.)}{M.M\,(Obs.)} = i \implies f = M \times i \quad \because M = \frac{wt}{M.M\,(Thero.) \times V(L)}$ *Formula mass of NaCl = 58.5*

➢ **Molarity & Normality equation :** To calculate the Molarity of mixing two solution of same substance after Dilution.

The following equation is used: $M_1 \times V_1 = M_2 \times V_2$ where M_1= initial molarity, M_2= molarity of the new solution, V_1= initial volume and V_2= volume of the new solution. Similarly, Normality equation: $N_1 \times V_1 = N_2 \times V_2$

If Vol. V_1 of Soltion of Molarity M_1 is mixed with Volume V_2 of another Non-reacting Solution of molarity M_2, Then Molarity M_3 of mixed Solution.

$$\underset{\text{Sol}^n\ 1}{\boxed{M_1V_1}} + \underset{\text{Sol}^n\ 2}{\boxed{M_2V_2}} \longrightarrow \underset{\text{Mixed Sol}^n}{\boxed{(V_1+V_2)\,M_3}} \qquad M_3 = \frac{M_1V_1 + M_2V_2 + ----}{V_1 + V_2 + -------}$$

Ex.o1. 12.5 ml of a solution containing 6 g of a dibasic acid in 1L was found to be neutralized by 10 ml of a decinormal solution of NaOH. The molecular mass of the acid is-

Sol. **$N_1 \times V_1$ (acid) = $N_2 \times V_2$ (base)** $\implies N_1 \times 12.5 = \frac{1}{10} \times 10 \implies N_1 = \frac{1}{12.5}$

$\because$ **Sytrength** $= N \times eq.Mass \implies \frac{mass}{V(L)} = N \times eq.mass \implies \frac{6}{1} = \frac{1}{12.5} \times E \implies E = 75$

Hence, Molar mass = eq. mass x n-factor $\implies molar\ mass = 75 \times 2 = 150$

8 . Normality (N) → (g eq./L) :

$N = \frac{\text{gram equivalent of solute}}{\text{Vol. of solution(L)}}$; $d_{sol} = \frac{\text{mass}}{\text{Vol.}}$

$\frac{\text{mass of solute in (kg)}}{\text{Eq. mass of solute}}$

$\frac{\text{mol. mass of Solute ---}}{\text{n-factor }(n_f)}$ depends on acidity(a)/basicity(b)/no. of e-

Acidity ⟶ No. of displaceble H+ ion ⟶ Base

Basicity ⟶ No. of displaceble OH- ion ⟶ Acid

1. For Acid : It depends on Chemical Reaction

$H_2SO_4 \rightarrow 2H^+ + SO_4^{2-}$ $n_f = 2$

$H_2SO_4 \rightarrow H^+ + HSO_4^-$ $n_f = 1$

$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$ $n_f = 1$

2. For Base : It depends on Chemical equation

$Ba(OH)_2 \rightarrow Ba^{2+} + 2\,OH^-$ $n_f = 2$

$NaOH \rightarrow Na^+ + OH^-$ $n_f = 1$

3. Salt : total charge x no. of cations

$Al_2(SO_4)_3 \rightarrow 2\,Al^{3+} + 3\,SO_4^{2-}$ $n_f = 2 \times 3 = 6$

$Na_2CO_3 \rightarrow 2\,Na^+ + CO_3^{2-}$ $n_f = 2 \times 1 = 2$

4. Oxidising/ Redusing Agent : No. of e- transfered

KMnO4 — Is good oxidising agent

- Acidic medium (H^+) ⟶ Mn^{2+} 7-2 = 5e- = $n_f = 5$
- Basic medium (OH^-) ⟶ MnO_4^{2-} 7-6 = 1e- = $n_f = 1$
- Neutral/Mild basic ⟶ MnO_2 7- 4 = 3e- = $n_f = 3$

● $C_2O_4^{-2} \rightarrow 2CO_2 + 2e^-$ Oxlate ion $n_f = 2$

● $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2\,Cr^{3+} + 7H_2O$ Dichromate ion $n_f = 6$

H_2O_2

- act as Oxi. Agent, In Acidic Medium ⟶ $H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$ $n_f = 2$
- Red. Agent, In Acidic Medium ⟶ $H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$ $n_f = 2$
- But, in case of Disproportionation reaction ⟶ $H_2O_2 \rightarrow H_2O + 1/2\,O_2$ $n_f = 1$

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NOTE: 1. monoacidic base $NH_3 \rightarrow n_f = 1$ (a) $NH_3 + H_2O \rightarrow NH_4^+ + OH^-$

$H_3BO_3 \rightarrow n_f = 1$ (b) $B(OH)_3 + H_2O \rightarrow [B(OH)_4]^- + H^+$

2. $H_3PO_2 \rightarrow n_f = 1$ $H_3PO_3 \rightarrow n_f = 2$ $H_3PO_4 \rightarrow n_f = 3$

** No. of Oxy. - 1

3. Proof n_f depends on ONLY Chemical reaction - b/coz ome H is replaced in given reaction

$H_3PO_4 + NaOH \rightarrow NaH_2PO_4 + H_2O$ $n_f = 1$

4. Eq. wt. of element = At. wt. / valancy Ex- eq. wt. of Al = 27/3 = 9

OR, $N = \frac{\text{strength(g/L)}}{\text{eq. mass of solute}} = \frac{\text{wt. of solute (g) x 100}}{\text{eq. mass x Vol. (cm}^3)}$

➢ **Percentage Composition :**

The mass percentage of each constituent element present in any compound is called its percentage composition.

Ex- 01. Calculate the percentage of water of crystallization in washing soda $[Na_2CO_3 \cdot 10H_2O]$

Ans. Molecular weight of $Na_2CO_3 \cdot 10H_2O$ = 46 + 12 + 48 + 180 = 286 amu.

Molecular weight of 10 molecules of water = 10 × 18 = 180 amu

$\therefore$ %age of water of crystallisation $= \frac{182}{286} \times 100 = 62.93\%$.

02. Calculate which one of the following is a better potash fertiliser (a) Potassium phosphate (b) potassium nitrate

Ans. (a) **For potassium sulphate** Molecular weight of K_3PO_4 = 3(39) + 1(31) + 4(16) = 117 + 31 + 64 = 212

$\therefore$ %age of potassium in $K_3PO_4 = \frac{117}{212} \times 100 = 55.18\%$

(b) For potassium nitrate Molecular weight of KNO_3 = 1(39) + 1(14) + 3(16) = 39 + 14 +48 = 101

$\therefore$ %age of potassium in KNO3 $= \frac{39 \times 100}{101} = 38.61\%$

$\therefore$ Potassium phosphate (K = 55.18%) is a better potash fertiliser than potassium nitrate (K = 38.61%)

Empirical Formula(E .F) & Molecular Formula(M. F) :

Note :- For certain compounds the molecular formula and the empirical formula may be same. Like- CO ∴ **M. F = E. F**

If the vapour density of the substance is known, its **Mol. weight = 2 x Vapour Density**

Empirical Formula & Molecular Formula :

Empirical Formula: the simplest whole number ratio of various atoms present in a compound

Molecular Formula: the exact number of different types of atoms present in a molecule of a compound

Benzene

CH → C_6H_6 = 6 (CH)

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Molecular Formula = n x Empirical Formula

Value of n = 1, 2, 3, 4, 5, 6 etc. ← Integer or common factor or multiplying factor

Molecular Weight = n x Empirical formula weight

Example- 01. A chemical compound having C% = 24.27 , H% = 4.08 & Cl% = 71.65% and its M.M is 98.96 g then find empirical and molecular formula.

Element	% of element	At. wt.	No. of moles = $\frac{\%}{At.wt.}$	Whole no. ratio
C	24.27	12	2.02	1
H	4.08	1	4.08	2
Cl	71.65	35.5	2.02 **(div. by min.)**	1

Hence, Empirical Formula = $\mathbf{CH_2Cl}$

Molecular Formula = **2** x $[CH_2Cl]$ = $\mathbf{C_2H_4Cl_2}$ $\quad \because \; n = \frac{M.F.mass}{E.F.mass} = \frac{98.96}{12+2+35.5} = \frac{98.96}{49} = \mathbf{2}$

02. A hydrocarbon contains 82.8 % of carbon and has a relative molecular mass of 58. Write (i) its empirical formula (ii) its molecular formula (iii) its two possible structural formulae.

Element	%age wt.	At. wt.	Relative no. of moles	Simple ratio of atoms
C	82.8	12	82.8 ÷ 12 = 6.9	6.9 ÷ 6.9 = 1 or 2
H	17.2	1	17.2 ÷ 1 = 17.2	17.2 ÷ 6.9 = 2.5 or 5

∴ (i) Empirical formula of hydrocarbon = C_2H_5

(ii) Empirical formula weight of hydrocarbon = 2 × 12 + 1 × 5 = 29 & Molecular wt. of hydrocarbon = 58

∵ Molecular weight = n × Empirical formula weight ⟹ 58 = n × 29 ⟹ n = 2

(iii) Molecular formula of hydrocarbon = 2 × Empirical formula = 2 x C_2H_5 = C_4H_{10}

Possible structural formulae : n-Butane & Iso-Butane. (Only saturated)

Analytical Chemistry : % determination of element on Organic Compound -

1. % Estimation of C & H : Leibeg's method- $\quad C + O_2 \rightarrow CO_2$ (12g, 44g) $\quad H_2 + O_2 \rightarrow H_2O$ (2g, 18g)

$$C\,\% = \frac{12}{44} \times \frac{mass\ of\ CO2}{mass\ of\ Compound} \times 100 \qquad H\,\% = \frac{1\times2}{18} \times \frac{mass\ of\ H2O}{mass\ of\ Compound} \times 100$$

2. % Estimation of Sulphur - By $BaSO_4$ (137 + 32 + 64 = 233)

$$S\,\% = \frac{1\times32}{233} \times \frac{mass\ of\ BaSO4}{mass\ of\ Compound} \times 100$$

3. % Estimation of Phosphorous - by magnesium Pyrophosphate ($Mg_2P_2O_7$)

$$P\,\% = \frac{2\times31}{222} \times \frac{mass\ of\ Mg2P2O7}{mass\ of\ Compound} \times 100$$

4. % Estimation of Halogen (X - Cl, Br, I) - **By AgX : Carius Method -**

$$Cl\,\% = \frac{35.5}{143.5} \times \frac{mass\ of\ AgCl}{mass\ of\ Compound} \times 100$$

5. % Estimation of Nitrogen-

■ **Kjeldahl's Method is Not suitable in -**

(1) Nitrogen containing Ring - like- Pyridine, Quinoline etc.

(2) compound Containing $\mathbf{NO_2}$ (Nitrogen group) & $-N = N-$ (diazo group)

B/coz Above Compound do not completely digest with $(NH_4)_2SO_4$

Estimation of Nitrogen

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Duma's Method

- Estimation by N_2 gas
- This method is suitable for all organic compound (General applicability)
- $N\% = \frac{2 \times 14}{22{,}400} \times \frac{Vol.\ of\ N_2\ at\ STP\ (cm^3)\ or\ ml}{mass\ of\ Compound} \times 100$

Kjeldahl's Method

Estimation through ammonium sulphate ($(NH_4)_2SO_4$)

This method are suitable for Fertiliser, Drugs, Fast Food Packaging etc.

$N\% = \frac{1.4 \times N \times V}{Wt\ of\ Substance} = \frac{1.4 \times Meq.}{W}$

■ **Example -** The ammonia evolve from the treatment of 0.30g of an organic compound for the estimation of N_2 was passed in 100ml of 0.1M H_2SO_4. The excess of acid required 20 ml of 0.5 molar NaOH solution for complete neutralisation.

$$N\% = \frac{1.4 \times M_{eq.}}{W} = \frac{1.4 \times 10}{0.30} = 46.67$$

∵ H_2SO_4 0.1 M, 100ml, $N_1 \times V_1$ = 0.1 x 2 x 100 = 20

NaOH 0.5 M, 20ml, $N_2 \times V_2$ = 0.5 x 1 x 20 = 10

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Neutralised 20 - 10 = 10

6. % Estimation of oxygen :

O% = 100 - (C % + H % + Cl % + N % + ...)

Some Experimental Methods:

➢ Stoichiometry & Stoichiometric Calculations :

The word 'stoichiometry' is derived from two Greek words—Stoicheion (meaning element) and metron (meaning measure). Stoichiometry, thus deals with the calculation of masses (sometimes volume also) of the reactants and the products involved in a chemical reaction.

Example- How many moles of methane are required to produce 22 g CO_2 after combustion ?

Sol. $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$ $\because$ $44g\ CO_2$ *react with* $= 1\ mole\ CH_4$ $\therefore$ $22\ g\ CO_2$ *react with* $= \frac{1}{44} \times 22 = 0.5\ mol$

■ **NOTE:** *Balancing of any saturated Hydrocaron Tricks :* $C_nH_{2n+2} + \frac{3n+1}{2} O_2 \rightarrow n\,CO_2 + (n+1)H_2O$

➢ * Dulong Petit's law : *At. Wt x Specific Heat capacity = 6.4*

Example : The Specific heat of a metal is 0.16. Its approximate At. wt. would be- $\therefore$ At. wt. = 6.4/0.16 = 40

■ *** 1 g-ion = 1 mol** **Example- The charge on one gram ion of Al^{3+} ion.**

$\because$ *charge of* $1\ mole\ of\ e^- = N_A \times e^-$ $\therefore$ $Al^{3+}\ ion = 3\,N_A \times e^- = 3 \times 6.022 \times 10^{23} \times 1.6 \times 10^{-19}\ C$

➢Law of Isomorphism: According to Mitscherlich [year 1819].

When two or more compound forms similar type of crystals or able to form mixed crystals, they are known as isomorphs. For Examples: $MgSO_4.7H_2O$ (Epsom salt), $ZnSO_4.7H_2O$ (White vitriol) and $FeSO_4.7H_2O$ (Green Vitriol) are isomorphs of each other as their crystals posses same shape.

- The valency of elements that are similarly placed to that of other elements in their isomorphs are always same.
- In the above Example- **Fe, Zn & Mg** have same valency [2] & equal ratio of water molecule in each isomorphs.

➢Diffusion Method : It is based on Graham's Law of Diffusion.

The law states that *"under similar conditions of temperature and pressure, the rates of diffusion of gases are inversely proportional to the square roots of their densities (or molecular weights)."*

The rate of diffusion (r) of a gas at constant temperature is directly preoperational to its pressure

$r \propto P$ at constant temperature $r \propto 1/\sqrt{d}$ at constant temperature

$$\frac{r_1}{r_2} = \frac{P_2}{P_1} = \sqrt{\frac{M_2}{M_1}}$$

When molecular mass of a compound is expressed in grams, it is called Gram Molecular mass (GMM) or Gram Molecular weight (GMW).

➢ Iodometry & iodimetry : These two types of titrations are based on oxidation reduction in analytical chemistry..

- In iodimetry, a species is directly titrated with an *iodine solution*, but in iodometry a species is titrated with an *iodide* solution and then the released iodine is titrated with *thiosulphate.*
- Therefore, iodimetry is a direct method and iodometry is an indirect method.
- Iodometry can be used to quantify *oxidizing agents,* whereas iodimetry can be used to quantify *reducing agents.*

Iodimetry : This titration involves free iodine. Such direct estimation of iodine is called iodimetry.

This involves the titration of iodine solution with known sodium thiosulphate solution, whose normality is N. Let the volume of sodium thiosulphate used be V litre.

$I_2 + 2Na_2S_2O_3 \rightarrow 2NaI + Na_2S_4O_6$ $I_2 + 2e^- \rightarrow 2I^-$ (reduction) & $2I^- \rightarrow I_2 + 2e^-$ (oxidation)

(n=2) (n=1) the n-factor for the hypo solution *is 1*, whereas, when I write the balanced redox reaction, that is, $2S_2O_3^{2-} \longrightarrow S_4O_6^{2-} + 2e^-$ You can see that n-factor should be 2.

Equivalents of I_2 = Equivalents of $Na_2S_2O_3$ used = N × V

$\therefore$ Moles of $I_2 = \frac{N \times V}{2}$ & Mass of free I_2 in the solution $= \left[\frac{N \times V}{2} \times 254\right]g$

Example- $H_2S + I_2 \rightarrow 2H^+ + S + 2I^-$ $H_2SO_3 + I_2 \rightarrow SO_4^{2-} + 2I^- + 4H^+$ $2S_2O_3^{2-} + I_2 \rightarrow S_4O_4^{2-} + 2I^-$

Iodometry : is an indirect method of estimation of iodine. An oxidizing agent is made to react with excess of solid KI. The oxidizing agent oxidizes I^- to I_2.

Example- $Cr_2O_2^{2-} + 6I^- + 16H^+ \rightarrow 2Cr^{3+} + 3I_2 + 7H_2O$ $2Cu^{2+} + 4I^- \rightarrow 2CuI + I_2$

➢ Equivalent mass OR Chemical Equivalent :

It is defined as the number of parts by mass of the substances which combine/displace directly 1.008 parts by mass of Hydrogen **OR** 8 parts by mass of Oxygen OR 35.5 parts by mass of Chlorine.

Following method is used for determination of Eq. Mass:

1. Oxide formation Method : A known mass of the element is changed into oxide directly or indirectly.

Mass of Oxygen = mass of **oxide** - mass of element $\therefore$ Eq. mass of element $= \frac{mass\ of\ element}{mass\ of\ oxygen} \times 8$

2. Chloride formation method :

mass of Chlorine = mass of **chloride** - mass of element $\therefore$ Eq. mass of element $= \frac{mass\ of\ element}{mass\ of\ chlorine} \times 35.5$

3. Hydrogen Displacement method: This method is used for those element which can evolve hydrogen from acids i.e. Active Metals. The mass of librated hydrogen is determined using density of hydrogen(0.00009 at NTP).

$$\therefore \text{Eq. mass of element} = \frac{mass\ of\ element}{mass\ of\ Hydrogen} \times 1.008 = \frac{mass\ of\ element \times 1.008}{Vol.(ml) of\ Hydrogen displaced at\ NTP \times 0.00009}$$

$$= \frac{mass\ of\ element}{Vol.(ml) of\ Hydrogen displaced at\ NTP} \times 11200$$

4. Metal to metal displacement method : A more active metal can displace less active metal from its salt solution. if E1 & E2 are the Eq. masses of two elements and m1 & m2 their respective masses then - $\frac{m1}{m2} = \frac{E1}{E2}$

➢Percentage yield: **(i)** It is the ratio of actual yield of the reaction to the theoretical yield multiplied by 100.

(ii) % yield = $\frac{Actual\ Yield\ \ \mathbf{or}\ \ Actual\ amount\ of\ product}{Theoretical\ yield\ \ \mathbf{or}\ Calculated\ amount\ of\ product} \times 100$

Example- In the reaction $2\ NH_3(g) + 5\ F_2 \rightarrow N_2F_4 + 6HF$; 3.56 g N_2F_4 is obtained by mixing 2g NH_3 & 8 g F_2. The % yield of the production is-

Sol. $2\ NH_3(g) + 5\ F_2 \rightarrow N_2F_4 + 6HF$
34 g 190 g 104g

% yield = $\frac{3.56}{4.38} \times 100 = 81.28\ \%$

$\therefore\ amount\ of\ N_2F_4\ formed\ by\ 2\ g\ NH_3 = \frac{2}{34} \times 104 = 6.12\ g$

$\therefore\ amount\ of\ N_2F_4\ formed\ by\ 8\ g\ F_2 = \frac{8}{190} \times 104 = \mathbf{4.38\ g}$ [It is **L .R** so, Calculated amount]

➢Problems related with Mixture:

Ex. 4 gram of a mixture of $CaCO_3$ and Sand (SiO_2) is treated with an excess of HCl and 0.88 gm of CO_2 is produced. What is the percentage of$CaCO_3$ in the original mixture?

Ans. $CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$ & $SiO_2 + HCl \longrightarrow$ No reaction

Let, $CaCO_3$ = x gm then, 100 gm CaCO3 gives ⟶ 44 gm CO_2

x gm$CaCO_3$ gives⟶ 0.88 gmCO_2 $\Rightarrow \frac{100}{x} = \frac{44}{0.88} \Rightarrow x = 2\ gram$ ∴ % CaCO3 = $\frac{2}{4}$ x 100 = 50%

➢ Limiting Reagent [L .R] : In which, The reactant gets consumed first or limits; the amount of product is formed in the reaction is known as limiting reagent. The reactant which is not consumed completely in the reaction is called **excess** reactant.

■ **How to find L.R :-**
1. Write & balance Chemical equation.
2. Convert the amount of all species to **moles** = mass/M.Mass **OR** Use **Y**-map.
3. Divide the no. of moles by stoichiometric Coeff.
4. Whose value come smallest hence, It is L.R.

■ **Example-** 10g of hydrogen & 64g O_2 were filled in a steel vessel and exploded amount of water produced in this reaction will be- H_2 (g) + 1/2 O_2 (g) → H_2O *(l)*

Initial mole : *10/2 = 5 mol* *64/32 = 2 mol*
n/Stoci. coeff. : *5/1 = 5 mol* *2/0.5 = 4 mol (min. value so L.R)*
Final mole : *5-4 = 1 mol* *4-4 = 0* *1 x 4 = 4 mol* *(**Ans**.)*

POAC* : Power of Atom Conservation:-** Only Used relation no. of moles (Y-map) of atom.

- In this Method atom is always conserved i.e. Mass of reactant = mass of Product.
- This is follow the law of conservation of mass.
- Do not required writing the complete reaction or balancing.

Example - 01. Find out moles of CO2 formed combustion of 60 moles of Butane ?

$C_4H_{10} + O_2 \rightarrow 4\ CO_2 + H_2O$ 'C' is conserved so, **4 x n_{C4H10} = 1 x n_{CO2}** ⟹ **4 x 60 = n_{CO2}**

02. Calculate the amount of lime that can be produced by heating 100g of 90% pure substance.

sol. $\because\ CaCO_3 \rightarrow CaO + CO_2$ **apply POAC, Ca is Consumed,** $\therefore\ 1 \times n_{CaCO3} = 1 \times n_{CaO}$

$\Rightarrow \frac{wt.of\ CaCO_3}{M.M\ CaCO_3} = \frac{wt.of\ CaO}{M.M\ of\ CaO} \Rightarrow \frac{\frac{90}{100}\times 100}{100} = \frac{x}{56} \Rightarrow x = 50.4\ g$ $\therefore\ \boldsymbol{no. of\ moles\ CaO = \frac{50.4}{56} = 0.9\ mol}$

Equivalence Method* :**

Law of Equivalence: ' Always 1 eq. of a substance reacts with 1eq. of another substances & equivalent of each product is formed.'

aA + bB → cC + dD

No. of eq. of A = No. of eq. B = No. of Eq. c = No. of eq. D

S.N	Equivalant Concept	Mole Concept
1.	Balancing is not required	Balancing is required.
2.	No need of formation of product	Required formation of product
3.	No. of equivalants = wt/Eq. wt	No. of moles = wt/Mol.wt

Equivalance Method
m.eq.
1 m.eq. = 10^3 eq.
g.eq. (gram/eq.) or Simply No. of equivalance

No. of m Eq. = $\frac{wt}{eq.\ wt.} \times 10^3 = \frac{wt}{M.\ wt.} \times V.F \times 10^3 = \frac{wt}{M.\ wt.} \times V.F \times 10^3 \times \frac{V}{V} = N \times V\ (ml) = (M \times V.F) \times V(ml)$

$= no.of\ millimoles \times V.F = (no.of\ moles \times 1000) \times V.F = \frac{no.of\ particles}{N_A} \times 1000 \times V.F$

■ **NOTE:** **Actually → mEq.** $\xrightarrow{means}$ is 1/1000th of Chemical equivalent (wt/Eq.wt) $\xrightarrow{so}$ wt/Eq.wt x 10^3

Ex.- 01. How many gm of $K_2Cr_2O_7$ is present in 1L of its N/10 solution in acidic medium ?

Sol. $K_2Cr_2O_7 \xrightarrow{[H^+]} Cr^{3+}$ $so, n_f = 3 \times 2 = 6$ $\therefore no.of\ eq.of\ K_2Cr_2O_7 = no.of\ eq.of\ sol^n$

$\frac{wt}{eq.\ wt.} = N \times V\ (L) \Rightarrow \frac{wt}{M.\ wt./n_f} = N \times V\ (L) \Rightarrow \frac{wt}{294/6} = \frac{1}{10} \times 1 \Rightarrow wt = 4.9\ gm$

02. The no . of moles of $KMnO_4$ that will be needed to react with 1 mole of sulphide ion in acidic medium.

$\overset{+7}{KMnO_4} + \overset{+4}{SO_3^{2-}} \xrightarrow[Anjit]{[H^+]} Mn^{2+} + \overset{+6}{SO_4^{2-}}$ ∴ eq. of $KmnO_4$ = eq. of SO_3^{2-}

⇨ no. of moles x V.F = n x V.F ⇨n_{KmnO4} x 5 = 1 x 2 ⇨ n_{KmnO4} = 2/5

Sharp your Brain:

1. **Eq. mass of element** = $\frac{Atomic\ Mass}{nf}$ ⟹ At. Mass = Eq. mass x Valancy
2. **Eq. mass of substance undergoing disproportionation Redox Reaction-** n_f (n-factor) = $\frac{n1 \times n2}{n1+n2}$

3. **Strength of H_2O_2 :** Volume strength of H_2O_2 is expressed in terms of the vol. of Oxygen at NTP.

Vol. strength = 5.6 x N 'OR' = **11.2 M** **&** **% strength = $\frac{17}{56}$ x Vol. Strength**

4. **Strength = N x Eq. mass**
5. **Molar mass = Eq. mass x n_f**(n-factor)
6. **N = M x n-factor** $\because$ **N =** $\frac{g.eq.}{V(L)} = \frac{\frac{g}{Eq.mass}}{V(L)} = \frac{w}{\frac{M.M}{n-factor} \times V(L)} = \frac{w}{M.M \times V(L)}$ x n-factor = M x n-factor
7. If given density **d (g/ml) & w/w % (mass%) :** **M =** $\frac{\frac{w}{w}\% \times d \times 10}{M.M\ of\ solute}$ **&** **N =** $\frac{\frac{w}{w}\% \times d \times 10}{Eq.\ mass\ of\ solute}$
8. If says **w/v %** of Solution - **M =** $\frac{\frac{w}{V}\% \times 10}{M.M\ of\ solute}$ **&** **N =** $\frac{\frac{w}{V}\% \times 10}{Eq.\ mass\ of\ solute}$
9. If given **specific gravity & w/w %** of solute - **M =** $\frac{\frac{w}{w}\% \times specific\ gravity \times 10}{M.M\ of\ solute}$

$\because$ specific gravity = $\frac{density\ of\ any\ substance}{density\ of\ water(4^\circ C)}$

10. Relation **b/w N & M :** Normality of Solution **N = M x** $\frac{molar\ mass}{Eq.mass}$
11. Relation **b/w m & solubility : m =** $\frac{solubility \times 10}{M.M\ of\ solute}$ $\because$ solubility = $\frac{mass\ of\ solute}{mass\ of\ solvent}$ x 100
12. Relation **b/w m** & mole fraction of solute X_B $\Longrightarrow$ X_B = $\frac{m}{m + 55.5}$ $\because$ $X_A + X_B = 1$
13. molality (**m**) in terms of density(**d**) & Molarity(**M**) : **m =** $\frac{1000 \times M}{1000 \times d - M \times MA}$
14. molality (**m**) in terms of mole fraction & molecular mass : **m =** $\frac{XB \times 1000}{XA \times MA} = \frac{XB \times 1000}{(1-XB) \times MA}$

where, M_A → M. M of solvent M_B → M. M of solute X_B → mole fraction of solute X_A → mole fraction of solvent

15. **No. of moles (n) = M x V(L) =** $\frac{M\ x\ V(ml)}{100}$ $\because M = \frac{n}{V}$
16. If an element exists in Two Isotopes having atomic masses x & y with ratio p : q then find Average mass of isotopes- Av. At. mass = $\frac{p.x + q.y}{p+q}$

Solved Examples:

Q.1 Naturally occurring chlorine is 75.53% Cl^{35} which has an atomic mass of 34.969 amu and 24.47% Cl^{37} which has a mass of 36.966 amu. Calculate the average atomic mass of chlorine-

Sol. Av. at. mass = $\frac{\%\ of\ 1^{st}\ Isotopes \times its\ At.Mass + \%\ os\ 2^{nd}\ Isotopes \times its\ At.Mass}{100} = \frac{73.53 \times 34.969 + 24.47 \times 36.96}{100}$ **= 35.5 amu**

Q.2 Appropriate no. of significant digits $\frac{6.02 \times 10^{23} \times 4.00}{4.0 \times 10^{23}} = 6.02 \times 10^{23}$ $\therefore 6.0 \times 10^3$ **Ans.** (Min. S.F is Two)

Q.3 25 ml of solution of $Ba(OH)_2$ on titration with 0.1 M molar solution H_3PO_2 gave 35 ml. The molarity of $Ba(OH)_2$ solution was ?

Sol. For $Ba(OH)_2$ [$N_1 \times V_1$] = for H_3PO_2 [$N_2 \times V_2$] $\Longrightarrow$ $(M_1 \times a) \times V_1 = (M_2 \times b) \times V_2$ $\Longrightarrow$ $M_1 \times 2 \times 25 = 0.1 \times 1 \times 35$ $\Longrightarrow$ $M_1 = 0.07$ M

Q.4 What is the Normality of solution that contain 50g of H_2SO_4 dissolved in 15 L.

Sol. N = M x n-factor $\Longrightarrow$ N = $\frac{50/98}{15}$ x 2 = 0.068 N

Q.5 A 5.2 molal aqueous solution of methyl alcohol CH_3OH is supplied, What is the mole fraction of methyl alcohol in the solution?

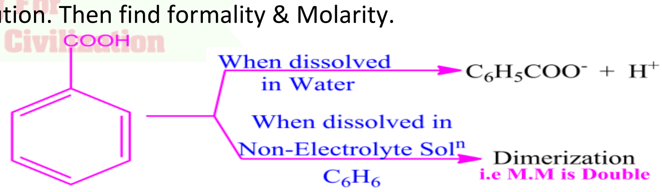
Sol. $x = \frac{m}{m+55.5} = \frac{5.2}{5.2+55.5}$ = 0.085

Q.6 2.44 g of benzoic acid is dissolved in benzene to form 500 ml solution. Then find formality & Molarity.

Sol. $M = \frac{w}{M.M_{acctual} \times V(L)} = \frac{2.44}{122 \times 500} = 0.04\ mol/L$

$\therefore\ f = i \times M = 0.5 \times 0.04 = 0.02\ \frac{mol}{L}$

$\because\ i = \frac{M.M_{theor}}{M.M_{observe}} = \frac{122}{244} = 0.5$



Q.7 A mixture of KBr and NaBr weighing 0.580g was treated with aqueous Ag^+ ion & all Br^- ion was recorded as 0.970 g of pure AgBr. What was the fraction by weight of KBr in Sample.

Sol. (KBr + NaBr) + Ag^+ → AgBr **'Br'** Consumed

Let, x g (0.560 - x)g 0.097 g so, $1 \times n_{KBr} + 1 \times n_{NaBr} = 1 \times n_{AgBr}$

Now, $1 \times \frac{x}{119} + 1 \times \frac{0.560-x}{103} = 1 \times \frac{0.970}{188}$ $\Longrightarrow$ x = 0.1332 g

so, fraction of KBr in mixture = 0.1332/0.560 = 0.237

Q.8 How many moles and grams of NaCl are present in 250ml of a 0.5 M NaCl solution ?

Sol. $\because\ no.of\ moles = \frac{M \times V(ml)}{1000} = \frac{0.5 \times 250}{1000} = 0.125$ $\therefore\ mass\ of\ NaCl = no.of\ moles \times mol.wt. = 58.5 \times 0.125 = 7.32\ g$

Q.9 Density of a 2.05 M solution of acetic acid in water is 1.02 g/ml. The molality of the solution is-

Sol. $\because m = \frac{M \times 1000}{1000 \times d - M \times M_B} = \frac{2.05 \times 1000}{1000 \times 1.02 - 2.05 \times 60} = 2.28\ mol/kg$

Q.10 A gaseous compound is composed of 85.7% by mass C & 14.3% by mass hydrogen. It's density is 2.28 g/L at 300K & 1 atm pr. Determine molecular formula of compound-

Sol. $\because\ d = \frac{PM}{RT}$ $\Longrightarrow M = \frac{dRT}{P} = \frac{2.28 \times 0.0821 \times 300}{1} = 56.15\ g/mol$ & **E.F** = $\frac{85.7}{12} : \frac{14.3}{1}$ = 7.14 : 14.3 = 1 : 2 so, E.F is **CH_2**

Now, n = $\frac{mol.mass}{E.F\ mass} = \frac{56.15}{(12+2)} = 4$ so, Mol. Formula = $4(CH_2)$ = ***C_4H_8***

Q.11 1g of Mg is burnt with 0.56g O_2 in closed vessel. Which reactant is left in excess & how much ?

Sol. Mg + 1/2 O_2 $\longrightarrow$ MgO

	Mg	1/2 O_2	MgO
Initial moles (n)	1/24 = 0.0416 mol	0.56/32 = 0.0175 mol	
n/stoi. coeff.	0.0416/1 = 0.0416 mol	0.0175/0.5 = 0.035 mol	**[min. value so, L.R]**
Final moles	(0.0416- 0.035) = 0.0066	(0.035 - 0.035)= 0 mol	1 x 0.035 = 0.035 mol

$\therefore mass\ of\ Mg\ left\ in\ excess = 0.0066\ \times 24 = 0.16\ g$

Q.12 Calculate molality of 1L solution of 93% H_2SO_4 (w/v) . The density of the solution is 1.84 g/ml.

Sol. mass of solute, in 1000ml of 93% H_2SO_4 = 93 g $\quad \therefore in\ 1000ml\ of\ 93\%\ H_2SO_4 = \frac{93}{100} \times 1000 = 930\ g$

Now, Mass of Solvent (H_2O) in 1000ml of solution = mass of Solution (d x v) - Mass of solute in 1000ml

= 1000 x 1.84 - 930 = 910 g $\quad \therefore\ m = \frac{930/98}{0.910}$ = 10.43 m

Q.13 The density of 3M sodium-thiosulphate solution each 1.25g/ml, Calculate-

(i) % by wt. of $Na_2S_2O_3$ **(ii)** mole fraction of $Na_2S_2O_3$ **(iii)** Molality of Na^+ & $S_2O_3^{2-}$

sol. Given data, d = 1.25g/ml, M = 3 mol/L means, mol = 3 (solute) $\because$ V = 1000ml = 1L (this is given)

(i) M = $\frac{\%\frac{w}{w} \times d \times 10}{M.M}$ $\Rightarrow$ w/w% = $\frac{M \times M.M}{d \times 10} = \frac{3 \times 158}{1.25 \times 10} = 37.92$

(ii) $x_{Na2S2O3} = \frac{n_{Na2S2O3}}{n_{Na2S2O3} + n_{H2O}} = \frac{3}{3+43.11} = \mathbf{0.0661}$

$\because mass\ of\ solvent\ (H_2O) = mass\ of\ solution\ (d\ \times v\ ) - mass\ of\ solute\ = \ 1.25\ \times 1000\ - 3\ \times 158 = 776\ g$

$\therefore\ moles\ of\ solvent\ (H_2O) = \frac{776}{18} = 43.11$

(iii) $Na_2S_2O_3 \rightarrow 2Na^+ + S_2O_3^{2-}$ $\quad \because m = \frac{3 \times 1000}{776} = 3.86\ m \quad \therefore\ Na^+\ =\ 2\ \times 3.86\ m\ \ \&\ \ S_2O_3^-\ =\ 1\ \times 3.86\ m$

Q.14 How many m.eq. of KCl are in a 50 ml does of 12% (w/v) KCl elixir ?

Sol. m.eq. = N x V(ml) = $\frac{(\%w/v) \times 10}{eq.\ wt.} \times V(ml) = \frac{12 \times 10}{75.5/1} \times 50\ =\ 80\ m\ mol$

Q.15. The strength of 10 volume of H_2O_2 solution is-

Sol. $\because\ M\ = \frac{Vol.\ strength}{11.2} \Rightarrow \frac{mass}{M.\ M\ \times V\ (L)} = \frac{Vol.strength}{11.2} \Rightarrow \frac{Strength}{M\ .\ M} = \frac{Vol.strength}{11.2} \Rightarrow \boldsymbol{Strength} = \frac{Vol.Strength}{11.2} \times M.M$

$= \frac{10}{11.2} \times 34 = 30.359\ g/L$

Q.16 In a victor Meyer determination of the relative molecular mass of benzene, the heating vessel was maintained at 120ºC. A mass of 0.1528 gm of benzene was used and the volume of displaced air collected over water at 15ºC, was 48 cm^3. The barometric pressure was 743 mm mercury. Calculate the relative molecular mass of benzene. The vapour pressure of water at 15ºC = 13 mm Hg.

Sol. Actual pressure of displaced air = 743 – 13 = 730 mm = $\frac{730}{760}$ atm $\quad \because$ 15 ºC = 15 + 273 = 288 K

V = 48 cm^3 = 48 × 10^{-3} lit. $\quad$ W = 0.1528 $\quad \therefore$ M = $\frac{WRT}{PV} = \frac{0.1528 \times 0.82 \times 288}{\frac{730}{760} \times 0.048}$ = 78.26

Hence, the molecular weight of benzene = 78.26

Q.17 There are 10 gm of mixture of NaCl and NaBr. If the amount of sodium is 25% of the weight of total mixture, calculate the amount of NaCl and NaBr present in the mixture. (Given, atomic weights of Na, Cl and Br are 23, 35.5 and 80 respectively).

Sol. Let in the given mixture NaCl present = x gm. $\quad \therefore$ in the given mixture NaBr present = (10 – x) gm.

The formula weight of NaCl = 23 + 35.5 = 58.5 and formula weight of NaBr = 23 + 80 = 103

$\therefore$ the total amount of Na present in the mixture = $\frac{23}{58.5}$ x + $\frac{23}{103}$ (10 – x) = 0.25 × 10 = 2.5;

on solving x = 1.5734 and 10 – x = 8.4266

$\therefore$ the amount of NaCl & NaBr present in the given mixture are 1.5734 gm and 8.4266 gm respectively.

Old IIT-JEE Objective type questions

Q.1 At 100ºC and 1 atm, if the density of liquid water is 1.0 g cm^{-3} and that of water vapor is 0.0006 g cm^{-3}, then the volume occupied by water molecules in 1 litre of steam at that temperature is -

(A) 6 cm^3 (B) 60 cm^3 (C) 0.6 cm^3 (D) 0.06 cm^3

Q.2 How many moles of electron weight one kilogram ?

(A) 6.023 × 10^{23} (B) $\frac{1}{9.108} \times 10^{31}$

(C) $\frac{6.023}{9.108} \times 10^{54}$ (D) $\frac{1}{9.108 \times 6.023} \times 10^8$

Q.3 A 2L solution (X) contain 0.02 mole of $[Co(NH_3)_5SO_4]Br$ and 0.02 mol $[Co(NH_3)_5\ Br]\ SO_4$. 1 L of this solution is - taken :

X + $AgNO_3$ (excess) $\rightarrow$ Y mol AgBr $\downarrow$

X + $BaCl_2$ (excess) $\rightarrow$ Z mol $BaSO_4 \downarrow$

Values of Y and Z are -

(A) 0.01, 0.02 (B) 0.02, 0.02

(C) 0.02, 0.01 (D) 0.01, 0.01

Q.4 In which of the following number of atoms are maximum ?

(A) 24 gms C (B) 27 gms of Al

(C) 56 gms of Fe (D) 108 gms of Ag

Q.5 Given that the abundances of isotopes ^{54}Fe, ^{56}Fe and ^{57}Fe are 5%, 90% and 5%, respectively, the atomic mass of Fe is -

(A) 55.85 (B) 55.95 (C) 55.75 (D) 56.05

Only single correct answer questions

Q.1 The atomic masses of two elements A and B are 20 and 40 respectively, if x gm of A contains y atoms, how many atoms are present in 2x gm of B ?

(A) y (B) 2y (C) $\frac{y}{2}$ (D) $\frac{y}{4}$

Q.2 If mole percentage of C–12 and C–14 in nature is 98% and 2% respectively, then the number of C–14 atoms in 12 g of carbon is -

(A) 1.2×10^{22} (B) 3.01×10^{22}
(C) 5.88×10^{23} (D) 6.02×10^{23}

Q.3 Total number of atoms of all elements present in 1 mole of ammonium dichromate is ?

(A) 14 (B) 19 (C) 6×10^{23} (D) 114×10^{23}

Q.4 25 grams of oleum contains 30% free SO_3. Strength of oleum is -

(A) 130% (B) 106.75 % (C) 115% (D) 110

Q.5 25 gram of A sample of oleum is labelled as 110%. The amount of H_2O which should be added to this sample to get 50 % H_2SO_4(w/w) is -(assuming **d** of H_2O = 1 g/ml)

(A) 25 gram (B) 20 gram (C) 30 gram (D) 27.5 gram

Q.6 The number of atoms contained in 11.2 L of SO_2 at N.T.P. are -

(A) $3/2 \times 6.02 \times 10^{23}$ (B) $2 \times 6.02 \times 10^{23}$
(C) 6.02×10^{23} (D) $4 \times 6.02 \times 10^{23}$

Q.7 The vapour density of gas A is four times that of B. If molecular mass of B is M, then molecular mass of A is -

(A) M (B) 4M (C) $\frac{M}{4}$ (D) 2M

Q.8 On analysis, a certain compound was found to contain iodine and oxygen in the ratio of 254 gm of iodine (at. mass 127) and 80 gm oxygen (at. mass 16). What is the formula of the compound -

(A) IO (B) I_2O (C) I_5O_3 (D) I_2O_5

Q.9 The hydrated salt $Na_2SO_4.nH_2O$, undergoes 55% loss in weight on heating and becomes anhydrous. The value of n will be

(A) 5 (B) 3 (C) 7 (D) 10

Q.10 The mass of oxygen that would be required to produce enough CO, which completely reduces 1.6 kg Fe_2O_3 (at. mass Fe = 56) is - ($Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$)

(A) 240 gm (B) 480 gm (C) 720 gm (D) 960 gm

Q.11 12g of Mg (atm. mass 24) will react completely with acid to give -

(A) One mol of H_2 (B) 1/2 mol of H_2
(C) 2/3 mol of O_2
(D) Both 1/2 mol of H_2 and 1/2 mol of O_2

Q.12 If one mole of ethanol (C_2H_5OH) completely burns to carbon dioxide and water, the weight of carbon dioxide formed is about - ($C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$)

(A) 22g (B) 45g (C) 66g (D) 88g

Q.13 Calculate the weight of lime (CaO) obtained by heating 200 kg of 95% pure lime stone ($CaCO_3$).

(A) 104.4 kg (B) 105.4 kg (C) 212.8 kg (D) 106.4 kg

Q.14 The mass of 70% H_2SO_4 required for neutralisation of 1 mol of NaOH

(A) 49 gm (B) 98 gm (C) 70 gm (D) 34.3 gm

Q.15 12 litre of H_2 and 11.2 litre of Cl_2 are mixed and exploded. The composition by volume of mixture is

(A) 24 litre of HCl (B) 0.8 litre Cl_2 and 20.8 lit HCl
(C) 0.8 litre H_2 & 22.4 litre HCl (D) 22.4 litre HCl

Q.16 What volume of 0.4-M $FeCl_3.6H_2O$ will contain 600 mg of Fe^{3+} ?

(A) 49.85 mL (B) 26.78 mL (C) 147.55 mL (D)87.65 mL

Q.17 The mole fraction of a given sample of I_2 in C_6H_6 is 0.2. The molality of I_2 in C_6H_6 is -

(A) 0.32 (B) 3.2 (C) 0.032 (D) 0.48

Q.18 1 mol of N_2 and 4 mol of H_2 are allowed to react in a vessel and after reaction, H_2O is added. Aqueous solution required 1 mol of HCl. Mol fraction of H_2 in the remaining gaseous mixture after reaction is -

(A) $\frac{1}{6}$ (B) $\frac{5}{6}$ (C) $\frac{1}{3}$ (D) None

Q.19 Calculate the weight of BaCl2 needed to prepare 250 mL of a solution having the same concentration of Cl^- ions as in a solution of KCl of concentration 80 g/L

(A) 27.92 g (B) 14.50 g (C) 22.52 g (D) 11.46 g

Q.20 0.2 mole of HCl and 0.1 mole of barium chloride were dissolved in water to produce a 500-mL solution. The molarity of the Cl^- ions is -

(A) 0.06 M (B) 0.09 M (C) 0.12 M (D) 0.80 M

Q.21 A sample of H_2SO_4 (density 1.8 g/ml) is 90% by weight. What is the volume of the acid that has to be used to make 1 litre of 0.2-M H_2SO_4?

(A) 16 mL (B) 10 mL (C) 12 mL (D) 18 mL

Q.22 A sample of Na_2CO_3. H_2O weighing 1.24 g is added to 200 mL of a 0.1-N H_2SO_4 solution. The resulting solution becomes -

(A) acidic (B) strongly acidic (C) alkaline (D) neutral

Q.23 125 mL of 10% NaOH (w/V) is added to 125 mL of 10% HCl (w/V). The resultant solution becomes-

(A) alkaline (B) strongly alkaline (C) acidic (D) neutral

Q.24 Equal volumes of 0.50 M of HCl, 0.25 M of NaOH and 0.75 M of NaCl are mixed. The molarity of the NaCl solution is -

(A) 0.75 M (B) 1/3 M (C) 0.50 M (D) 2.00 M

Q.25 How many grams of copper will be replaced in 2 L of a 1.50-M $CuSO_4$ solution if the latter is made to react with 27.0 g of Al ? ($3CuSO_4 + 2Al \rightarrow Al_2(SO_4)_3 + 3Cu$)

(A) 190 g (B) 95.25 g (C) 48 g (D) 10 g

Q.26 In which mode of expression, the concentration of a solution remains independent of temperature -

(A) Molarity (B) % w/V (C) Formality (D) Molality

Q.27 The largest number of molecules is in -

(A) 36 g of water (B) 28 g of CO_2
(C) 46 g of CH_3OH (D) 58 g of N_2O_5

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☎आपका परिश्रम + हमारा मार्गदर्शन = निश्चित सफलता ☎

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Old IIT-JEE Objective type questions

1-C	2-D	3-D	4-A	5-B

Only single correct answer type questions:

1-A	2-A	3-D	4-B	5-C	6-A	7-B	8-D
9-D	10-B	11-B	12-D	13-D	14-C	15-C	16-B
17-B	18-B	19-A	20-D	21-C	22-D	23-C	24-B
25-B	26-D	27-A					

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With Best Wishes