N-AVASTHI SIR

PHYSICS CHEMISTRY

HANDWRITTEN NOTES

JEE-MAIN
JEE-ADVANCED
NEET

VOL. 1

IIT-NEET.XYZ

29/4

MATTER

MATTER :-

Anything that occupies space and has some mass.

★ Classification of Matter :-

1) Physical classification :

	Shape	Volume
→ solid :	Definite	Definite
→ Liquid :	No definite	"
→ gases :	No. definite	No definite

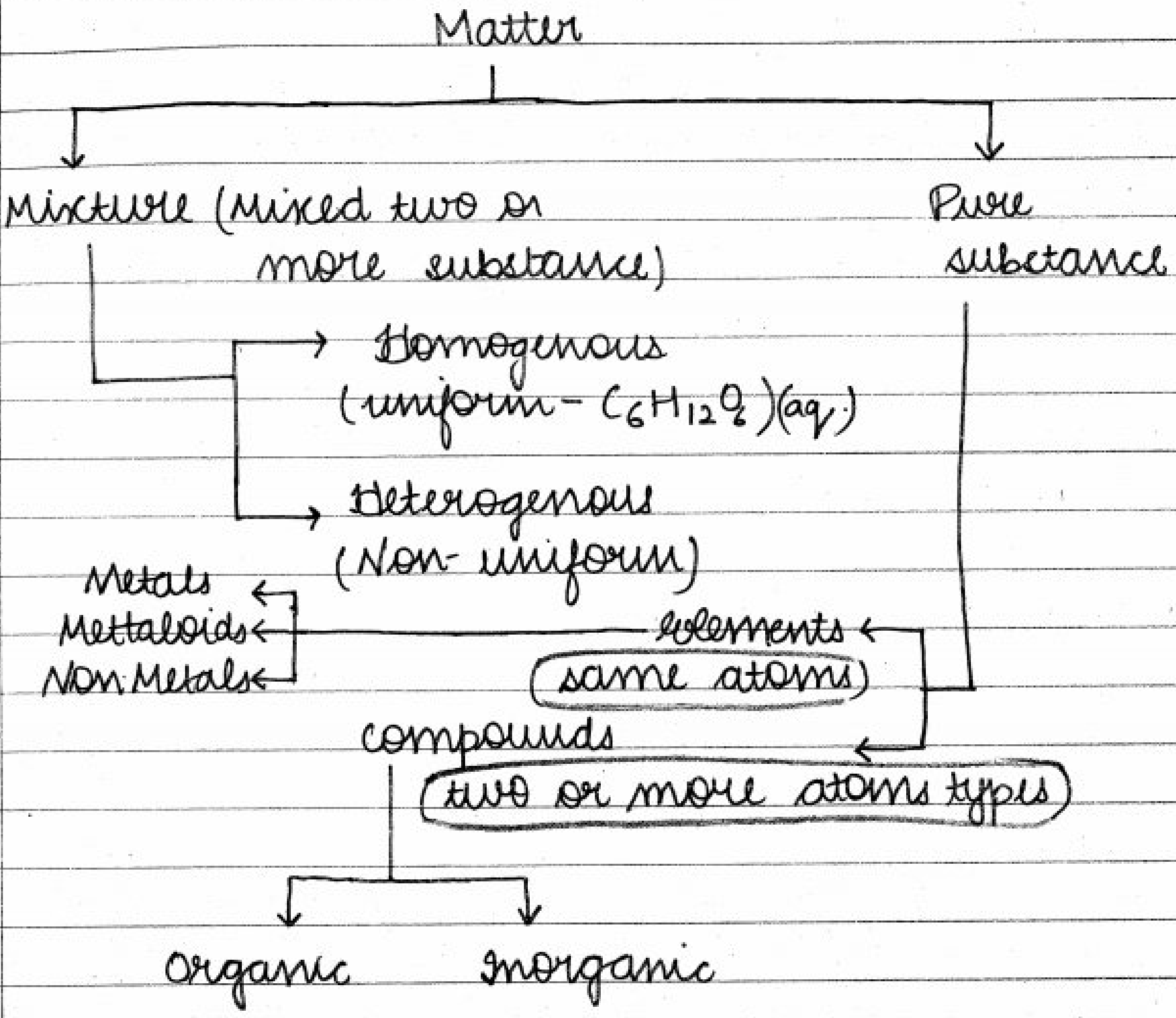
⇒ Plasma : gaseous ions at very high temperature.

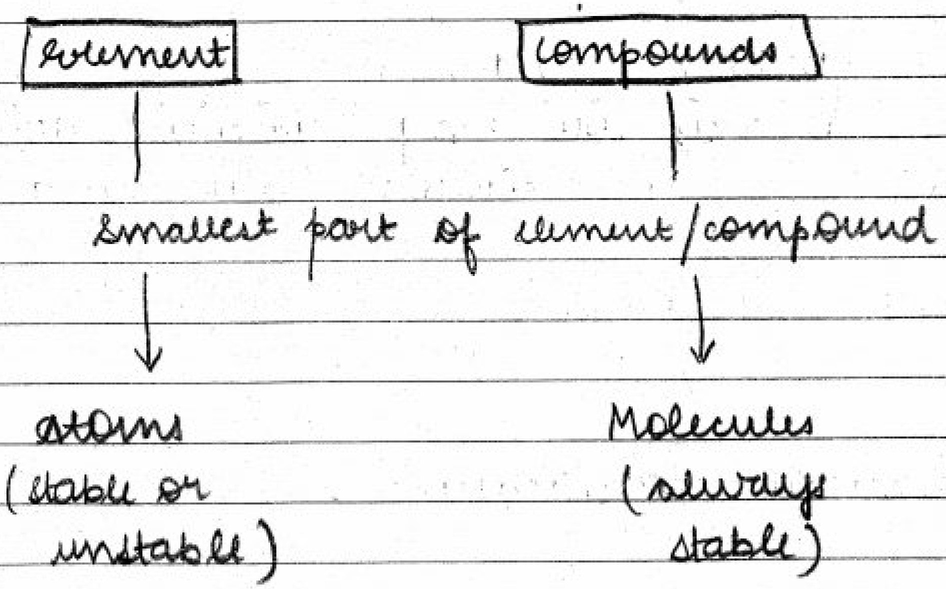
⇒ Bose-einstein Condensate :

At very low temp. gaseous atom merge or condense into single entity.

⇒ Fermi- condensate.

2) Chemical classification :


Matter
Mixture (Mixed two or more substance)
Pure subctance
Homogenous (uniform - $C_6H_{12}O_6$)(aq.)
Heterogenous (Non-uniform)
Metals
Mettaloids
Non Metals
elements (same atoms)
compounds (two or more atoms types)
Organic
Inorganic


element
compounds
smallest part of element/compound
atoms (stable or unstable)
Molecules (always stable)

MOLE CONCEPT.

★ **ATOM :**

smallest particle of an element that retains characteristics and chemical properties of that element.

Atoms may or mayn't have independent existence

$O(g)$ not stable at 1 atm.

$O_2(g)$ stable at 1 atm and room temperature.

★ **MOLECULE :**

smallest unit of a compound that retains composition and chemical properties of that compound.

H_2O — H_2O molecule contains 2 atoms of hydrogen and 1 atom of O.

—x—

Atomic Mass Unit (a.m.u.)

★ 1 a.m.u. is $\left(\frac{1}{12}\right)^{th}$ mass of $^{12}_{6}C$ atom

$$1\ \text{a.m.u.} = \frac{\text{Mass of Carbon-12}}{12}$$

$$\rightarrow \quad = \frac{1.992 \times 10^{-23}}{12} = \begin{array}{r} \\ \text{or} \end{array} \boxed{\begin{array}{l} 1.66 \times 10^{-24}\ g \\ 1.66 \times 10^{-27}\ kg. \end{array}}$$

★ Exact mass and atoms measured by "MASS spectometer"

★★ **ATOMIC MASS :**
Mass of an atom.

★★ **MOLECULAR MASS :**
Mass of. molecule.

★ Atomic mass of H atom = 1 a.m.u.
Molecular mass of Hydrogen = 2 a.m.u.

★★ **FORMULA MASS :**
Mass of ionic compound.
(They have strong forces between cation and anions)

like —

$NaCl$
Na_2SO_4
ZnS.

$^{A}_{Z}X$ $\quad$ $^{37}_{17}Cl$

★★ Z = Atomic number = No. of protons.
A = Atomic mass = Total no. of proton and nuetron

Mass of an atom = Mass of (e + p + n)
= Mass of (p+n)

★ $X \longrightarrow X^{+} + e^{-}$ → cation
★ $X + e^{-} \longrightarrow X^{1-}$ → anion

★★ **Relative Atomic Mass** : (NO UNIT)

$$\frac{\text{Mass of atom}}{\left(\frac{1}{12}\right)^{th} \text{ of mass of } ^{12}_{6}C} = \frac{\text{Mass of atom}}{1\,a.m.u.}$$

★★ **Relative Molecular Mass** : (NO UNIT)

$$\frac{\text{Molecular Mass}}{\left(\frac{1}{12}\right)^{th} \text{ of mass of } ^{12}_{6}C} = \frac{\text{Molecular Mass}}{1\,a.m.u.}$$

★ Relative atomic mass Na = 23
Relative molecular mass H = 2 {No units}

1 a.m.u = 1 u (unified mass)
= 1 Dalton

—x—

MOLE

★ A mole is a counting unit used by chemist / scientist.

Imp.

1 Mole = 6.022×10^{23} entities

★ entities may be —
Atoms / Molecules / electron / ion / p / n

★ Avogadro number is a pure no. experimentally it is 6.022×10^{23} (No unit)

★

★★★ NO. OF MOLES = $\frac{\text{given no. of entities}}{N_A}$

★★ N_A is Avogadro constant

It has unit $\boxed{mol^{-1}}$

It has unit because moles has unit → mol.

★ calculate total no. of neutrons in 5 mol of H_2O

→ One molecule of H_2O has 8 neutrons

$\boxed{5\times 6.022\times 10^{23}\times 8.}$

Total moles of protons in 5 Moles of H_2O = 5×10.

—x—

★ calculate total moles electrons in 3.011×10^{23} Mg ions ?

$\boxed{5}$

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★ AVOGADRO NO. :-

No. of $^{12}_{6}C$ atoms in exactly 12g carbon sample ($^{12}_{6}C$)

$$\boxed{\text{Avogadro no.} = \frac{12g}{\text{Mass of } ^{12}_{6}C \text{ atom}} = \frac{12}{1.992 \times 10^{-23}}}$$

$$= 6.022 \times 10^{+23}$$

—x—

$$1\ a.m.u = \frac{\text{Mass of } ^{12}_{6}C \text{ atom}}{12}$$

Numerically $1 a.m.u = \frac{1}{\text{Avogadro no.}} = \frac{1}{6.022 \times 10^{23}}$

$$= 1.66 \times 10^{-24} \text{ mol.}$$

—x—

★ Atomic Mass Of An Element

Average atomic mass of isotopes of that element.

$^{1}_{1}H$, $^{2}_{1}H$, $^{3}_{1}H$

atomic mass of $^{37}_{17}Cl$ = 37 a.m.u. (25%)

Atomic mass of $^{35}_{17}Cl$ = 35 a.m.u. (75%)

↓ [Natural abundance]

$$\text{Atomic mass of chlorine} = \frac{75 \times 35 + 37 \times 25}{100} = 35.5 \text{ a.m.u.}$$

Imp Atomic mass of Na = 23

atomic of C = 12

(Because their isotopes are .0000...1%)

$^{24}_{11}Na$ $^{23}_{11}Na$ $^{14}_{6}C$, $^{12}_{6}C$

Isotopes : same no. of protons

Different atomic mass.

$\boxed{^{1}_{1}H,\ ^{2}_{1}H,\ ^{3}_{1}H}$

—x—

MOLAR MASS

Mass of 1 mole of any substance.
(in g.)

Mass per mole

—x—

★★ Mole in terms of Mass :

No. of moles of subs. = $\frac{\text{given mass of sub (g)}}{\text{Molar mass (g/mol)}}$

Mole is S.I. unit of amount of substance.

★ Molar mass = Mass per mole of substance

Units of molar mass—
gram/mole $g\,mol^{-1}$

—x—

Mass of 1 mole substance =
Mass of 6.022×10^{23} atoms of Mg

$= 6.022 \times 10^{23} \times$ Mass of Mg atom

$= 6.022 \times 10^{23} \times 24$ a.m.u.

$= 6.022 \times 10^{23} \times 1.66 \times 10^{-24} \times 24$

$= \boxed{24\,g.}$

—x—

Important -

(GAM)

★ Mass of 1 mole of x atoms is gram atomic Mass

★ Mass of 1 mole of Y molecular is gram Molecular Mass (GMM)

★ Mass of 1 mole of Z formula unit is gram formulas Mass

—x—

★★★ $\boxed{\text{TOTAL NO. OF ATOMS}}$ =

No. of atoms in 1 molecule × Avogadro no. × No. of molecules.

★ Calculate no. of moles present in 120 ~~Mg~~ mg in Mg.

No. of moles = $\frac{120 \times 10^{-3}}{24} = \underline{5 \times 10^{-3} \text{ mol}}$.

★ given : 90 g $H_2O(l)$ sample :

Calculate – (i) no. of moles in H_2O molecule
(ii) " " oxygen atoms
(iii) " " hydrogen atom
(iv) " " moles of protons
(v) " " nuetrons.

(i) $\frac{90}{18} = 5 \text{ mol}$

(ii) $5 \times 6.022 \times 10^{23}$

(iii) $5 \times 6.022 \times 10 \times 2$

(iv) $10 \times 5 = 50 \text{ mol}$

(v) $8 \times N_A \times 5 = 40 N_A$

★★ given no. – 20g of CO_3^{2-}
Calculate no. of electrons.

$$\text{No. of moles} = \frac{20}{60} = \frac{1}{3}$$

~~Nof~~ No. of electrons in 1 molecule = ~~127~~ $6+24+2$

$= 32$

$$\therefore \text{No. of electrons} = \boxed{\frac{1}{3} \times N_A \times 32}$$

—x—

★ 1 mole is amount of substance that contains as many entities as $^{12}_{6}C$ atoms present in exactly 12g of carbon sample.

—x—

★★

1 g-atom = 1 mole of atoms
1 g-molecule = 1 mole of molecule
1 g-ion = 1 mole of ions

IDEAL GASES:

obey $\boxed{PV = nRT}$

P = Pressure
V = volume
n = no. of mole
R = gas constant
T = Temp (K)

for ideal gas :-

$\boxed{\text{volume of ideal gas} = \text{volume of vessel}}$

** Imp

1atm = 10135.25 Pa or N/m^2

1 bar = 10^5 Pa = 100 KPa

$\boxed{1\text{atm} = 1.01325 \text{ Bar}}$ 1 atm ≃ 1 Bar

—x—

$1 m^3 = 1000L$

$1L = 1000 cm^3$ or ml

$1ml = 1 cm^3$

unit of R :- atm·L/(mol·K)
(gas constant)

R = 0.0821 atm·l/(mol·K)

R = 8.314 J/mol·K
= 1.98 cal/mol·K
≃ 2 cal/mol·K.

$R = .0821 \times 101325 \times 10^{-3} \frac{N}{m^2} \times m^3/mol.K$

= 8.314 J/mol.K.

★ Calculate volume of 1 mole of gas (Molar volume) at P 1 atm & 273K.

① $V = \frac{nRT}{P} = \frac{.0821 \times 273}{1} = \boxed{22.4\,l}$ ★

★ volume of 1 mole of ideal gas 1 bar and 273 K

② $\boxed{V = 22.71\,L}$

③ $V = \frac{nRT}{P} = .0821 \times 300 = \boxed{24.63}\,l$.

—x—.

★ **STP** (Old) : Standard Temperature & Pressure

$P = 1\,atm$ $T = 0°C$ or $273K$.

★ **NTP** : Normal Temp. & Pressure

Both are same.

STP (New) = 1 bar, 273 K.

—x—

$$\boxed{\text{No. of moles of ideal gas} = \frac{\text{volume of ideal gas (in L) at atm and 273K}}{22.4\,l.}}$$

* Calculate no. of moles at 560 l at 1atm & 273 K.

Sol. Moles of gas = $\frac{560}{22.4} = 25$ mol.

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* Calculate moles of ideal gas in 8.21 L at 4 atm and 127°C

Sol. → $n = \frac{PV}{RT} = \frac{4 \times 8.21}{.082 \times 400}$

$= 1$ mol.

—x—

Note: →

At constant Temp and Pressure

$$\boxed{V \propto n}$$

$\boxed{\text{Mol \%} = \text{Volume \%}}$ for ideal gases.

Q. Calculate atoms in 112 l O_3 at 1atm & 273 K.

Sol.

$$n = \frac{112}{22.4} = 5 \text{ mol.}$$

$$\therefore \text{no. of atoms} = 3 \times 5 \times N_A$$

$$= \boxed{15 N_A}$$

Q. Calculate protons in 5.6 L of H_2 at NTP

$$n = \frac{5.6}{22.4} = \frac{1}{4}$$

$$\text{No. of Protons} = 2 \times \frac{1}{4} N_A = \boxed{\frac{N_A}{2}}$$

---

* No. of moles = $\dfrac{\text{Mass of gas}}{\text{Molar mass}}$

$$n = \frac{W}{M}$$

→ $PV = nRT$

$PV = \frac{W}{M} RT$ and Density $(\rho) = \frac{\cancel{M} W}{V}$

$$\therefore \boxed{\text{Density} = \frac{PM}{RT}}$$

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Mole Concept

In terms of ~~Mo~~ No.

$$n = \frac{\text{given entities}}{N_A}$$

In terms of Mass

$$n = \frac{\text{given mass}}{\text{Molar mass}}$$

In terms of volume

ideal gases

$$n = \frac{\text{V (in L) of gas at 1atm \& 273 K}}{22.4}$$

—x—

MIXTURE ANALYSIS

Mixture contains A and B

★ Wt. of mixture = Wt. of A + Wt. of B.

★★ % by weight or weight % of a substance =

$$\frac{\text{Wt. of A}}{\text{Wt. of mixture}} \times 100$$

★ Weight fraction = $\frac{\text{Wt. of A}}{\text{Wt. of mixture}}$

fraction % = Wt. fraction × 100

** **Note :-**

Moles of mixture = Moles of A + Moles of B

$$\text{Mole \% of A} = \frac{\text{Moles of A}}{\text{Moles of mixture}} \times 100$$

$$\text{Mol fraction of B} = \frac{\text{Mole of B}}{\text{Moles of mixture}}$$

Mol % = Mole fraction × 100

Q A mixture contains A 16g (M = 4g/mol) & 24 g of B. (M = 12g/m)

(i) Wt. % of A
(ii) Mol % of B

→ (i) $\frac{16}{40} \times 100 = 40\%$

(ii) $\frac{\text{Moles of B}}{\text{Moles of mix.}} \times 100$

$$\frac{24/12}{24/12 + 16/4} \times 100 = \frac{2}{6} \times 100$$

$$= \boxed{33.33\%}$$

—x—

FOR ISOTOPES :

Important ⇒ * $\boxed{\text{Atom\%} = \text{Mole\%}}$

* $\boxed{\text{Mole\%} = \text{\% of natural abundance}}$

—x—

Q In 100 g sample we have
75% of ^{35}Cl 25% of ^{37}Cl

$$\text{Molar mass of chlorine} = \frac{\text{Total wt of chlorine}}{\text{Total moles}}$$

$$= \frac{75 \times 35 + 25 \times 37}{100}$$

$$= \boxed{35.5\ g}$$

—x—

$$\text{Molar mass of element} = \frac{M_1x_1 + M_2x_2 + M_3x_3 \ldots\ldots}{x_1 + x_2 + x_3 \ldots\ldots}$$

$$= \boxed{\frac{\Sigma M_i x_i}{\Sigma x_i}}$$

→ M_i = Molar mass of isotope
x_i = Mole % of isotope

—

Q. Wt.% of $^{35}_{17}Cl$ (in natural sample) $= \dfrac{\text{Wt. of } ^{35}_{17}Cl}{\text{total wt. of chlorine}} \times 100$

$$= \frac{35 \times 75}{35.5 \times 100} \times 100$$

$$= 75 \times \left(\frac{35}{35.5}\right)$$

↓

Imp

$$\text{Wt\% of particular isotope} = \underline{\text{Mole\% of fraction}} \times \frac{\text{Molar mass of i}}{\text{Molar mass of element}}$$

[★] Similarly for mixture analysis:

Imp

$$\boxed{\text{Wt\% (or Mass\%) of a substance} = \text{Mol\% of a substance} \times \frac{\text{Molar mass of sub.}}{\text{average molar mass of mixture}}}$$

sample – 4% by wt H_2, 96% of O_2

Q. Calculate : (i) Mole % of H_2

(ii) Molar mass of gaseous mixture.

Ans. Let we have 100g of sample.

(i) Wt. of H_2 = 4g   Wt. of O_2 = 96g

Moles = $\frac{4}{2}$ = 2 mol   Moles = $\frac{96}{32}$ = 3 mol.

Mol% = $\frac{2}{5} \times 100$ = 40%

(ii) Molar mass of mixture = $\dfrac{\text{Wt. mixture}}{\text{total moles}}$

$= \frac{100}{5} = 20$ g/mol.

Q. sample contains :

80% mole% of H_2N_2 and 20 mol% O_2

(i) Molar mass of air = $\frac{\cancel{100}}{\frac{80}{28}+\frac{\cancel{20}}{\cancel{32}}} = \frac{\cancel{100}}{\frac{20}{7}+\frac{\cancel{5}}{8}} = \frac{5600}{195}$ g/mol

(ii) Wt% of N_2 $\frac{80}{100}\cancel{\times 100} \cdot \boxed{\cancel{80\%}}$

Molar mass = $\frac{80\times 28 + 20\times 32}{100} = 28.8$

Wt.% of nitrogen = $80 \times \frac{28}{28.8} = 77.77\%$

★ $H_2O(l)$ = Water
$H_2O(g)$ = steam/water vapour.

★ (1) At 1 atm and below 100°C	★ (2) At 1 atm and at 100°C both are in equilibrium	★ (3) At 1atm and above 100°C
↓		
$H_2O(l)$ present (stable)	Both stable (No. of moles remain constant)	$H_2O(g)$ is stable

Q. Two isotopes of boron are
B^{10} and B^{A}
% abundance of of B^{10} is 20%
average atomic wt. of boron = 10.8 amu

Value of A =

$$\frac{20\times10 + \cancel{x}\ 80\times x}{10.8} = 100$$

$$x = \frac{1080 - 200}{80} = \frac{880}{80} = \boxed{11}$$

Q. Atomic wt. Ag is 108 amu.
and it's isotopes are Ag^{107} and Ag^{109}

$$\Rightarrow \frac{107\times x + 109(109-x)}{100} = 108$$

$$x = 50\%$$

---

* [Q.] Calculate moles of 90ml H_2O
at 1atm 300 K.

~~PV = nRT~~

$$\frac{90\times}{100\times300\times.0821} = n$$

=

90 ml ($H_2O(l)$)
Mass of 90ml H_2O
= 90×1 = 90g
Moles = $\frac{90}{18}$ = 5 mol.
Ans.

[Q.] Calculate moles of H atom in 90 ml H_2O
at 1 atm and 127°C
($H_2O(g)$)

Solution : PV = nRT

$$n = \frac{PV}{RT} = \frac{1\times90}{1000\times.0821\times400}\times2 =$$

Q. $Fe(ScN)_3 \cdot xH_2O$ has 81% $Fe(ScN)_3$. find x.

$$\underline{56+(32+12+14)3} = 56+174 = \frac{232}{232+18x} \times 100$$

$$= 81$$

$$23200 = 81 \times 232 + 18x \times 81$$

$$19 \times 232 = 18x \times 81$$

$$\boxed{x = 3}$$

Q. $1\overset{5}{\cancel{8}}g$ $Na_2SO_4 \cdot xH_2O$ is heated,
the constant mass and residue (Na_2SO_4)
is 8 gram.

$$\frac{8}{18} = \frac{142 \cancel{+18x}}{142+18x}$$

$$142 \times 8 = 142 \times 1\cancel{8}5 - 1\overset{8}{\cancel{8}} \times 8x$$

$$x = \frac{1420}{1\cancel{8}5 \times 8} = \frac{142 \times 7}{15 \times 8}$$

$$= 6.9 = 7.$$

★ MOLECULAR FORMULA & EMPIRICAL FORMULA.

★ **Molecular formula**:
It represent actually no. of elements present in a molecule.

★ **Empirical formula**:
It represents simplest ratio (whole no.) of atoms of element in a molecule.

★ M.F. of Nitrobenzene – $C_6H_5ON_2$
E.F. ______ – "
M.F. – H_2O (Hydrogen Peroxide) – H_2O_2
E.F. – HO

★★ Molecular formula = $(\text{Empirical formula})_n$
$\{n = 1, 2, 3, \ldots\}$

$$n = \frac{\text{Molar mass of substance in M.f.}}{\text{Molar mass " in E.f.}}$$

Q Hydrazine contain 87.42% N and 12.58% H
Molar mass of Hydrazine 32 g.

Ans. Let 100 g.
87.42 g ~~A~~ = Moles of N = $\frac{87.42}{14} = 6.24$
12.58 g Moles of H = 12.58

$$\frac{\text{Mole of N}}{\text{Mole of } \cancel{O}\text{H}} = \frac{1}{2}$$

E.f. of NH_2

$$\therefore\ n = \frac{32}{16} = 2$$

$\boxed{N_2H_4}$

—x—

Q 7.3 g of organic compound contains 3.6 g C 0.7 g H, 1.4 g N and rest O (1.6g)

$\boxed{C_3H_7NO}$ - empirical formula.

★ Determination of E.F. and M.F

Step #1. calculate no. of moles of every element.

Step #2 Calculate molar ration of elements (Moles of elements divide by min. moles of element)

Step #3 If mole ratio is fractional then multiply by a suitable number and convert simplest ratio.

Step #4 Calculate ~~no.~~ n and find M.f.

★★ If sum of percentange of constituent elements of compound not equal to 100, then remaining % is %. of O.

★ Q. 60% C 8% H. find E.f.

Moles of C – $\frac{60}{12} = 5$
Moles of H – $\frac{8}{1} = 8$
Moles of O – $\frac{32}{16} = 2$

{In 100g}

$\boxed{C_5H_8O_2}$

Q. 1.4 g of organic compound burned in excess of air oxygen gave 4.4 g CO_2, 1.8 g H_2O. find. E.F.

$$CO_2 - \frac{4.4}{44} = .1 = n_{CO_2} = n_C$$

$$H_2O - \quad .1 = 2n_{H_2O} = \cancel{2}n_H$$

Moles of C – .1
Moles of H = .2

$\boxed{CH_2}$

Q. A sample of Ne consists 91.8% $^{20}_{10}Ne$ .5% $^{21}_{10}Ne$ and 7.7% $^{22}_{10}Ne$. Calculate atomic mass of neon.

$$\frac{20 \times 91.8 + .5 \times 21 + 7.7 \times 22}{100}$$

$$= \frac{18'36 + 10.5 + 169.4}{100}$$

$$= \boxed{20.159 \text{ amu}}$$

—x—

$\boxed{\text{imp}}$

If % not mentioned
↓

- ★ In compounds (like $C_6H_{12}O_6$) → Mass %
- ★ In isotopes → Mole%

—x—

★ DENSITY of a substance : $\boxed{\dfrac{\text{Mass}}{\text{Volume}}}$ Unit g/cm³ or g/ml or kg/m³

★ RELATIVE DENSITY of substance No unit

$$= \boxed{\dfrac{\text{Density of substance}}{\text{Density of reference/standard}}}$$

★★ FOR SOLIDS / LIQUIDS :

$$\text{Specific gravity} = \boxed{\dfrac{\text{Density of solid/liquid}}{\underbrace{\text{Density of } H_2O}_{\text{Fixed standard} - H_2O}(l) \text{ at } 4°c}} \quad (1\,g/cm^3)$$

↓ No Unit

★★ FOR GASES:

$$\text{Vapour Density of gaseous substance} = \frac{\text{Density (gas) at certain temp. \& pr.}}{\text{Density of } H_2(g) \text{ at same temp and pr.}}$$

We know $\boxed{D = \dfrac{PM}{RT}}$ $= \dfrac{PM_{gas}}{\cancel{RT}} \Big/ \dfrac{PM_{H_2}}{\cancel{RT}}$

$$= \frac{M_{gas}}{M_{H_2}}$$

Important V.D gas $= \boxed{\dfrac{M_{gas}}{2} \text{ (Molar mass)}}$

Q. gaseous sample 2mole N_2O_4 3 mole NO_2.
find V.D.

Sol.

$$\frac{(28+64)2+(46)3}{2\times5} = 92+6\cancel{8}\cancel{6}9$$

$$= \frac{16\cancel{0}1}{5} = \boxed{32.2}$$

Q. V.D. of mixture NO_2 & N_2O_4 is 27.6
cal. (i) Mol% of NO_2 (ii) Mass% of NO_2

Molar mass$_{(mix.)}$ = 55.2

$$46x + 9\overset{2}{\cancel{8}}(1-x) = 27.6\times2$$

$$46x + 9\cancel{8}^{2} - 9\cancel{8}^{2}x = 27.6\times2$$

$$32x = 50.4$$

$$x = \frac{50.4}{32}\ 22.8$$

$$46x = 36.8$$

(i) $x = \frac{36.8}{46}\times100$ = Mol%

(ii) Mass% of NO_2

~~Let~~ 100g mix. Total mass= 55.2

Mole persent of NO_2 = 20%

$$\therefore\ x\times\cancel{92}\,46 + 92\times4x = 55.2$$

$$46(9x) = 55.2$$

$$x = \frac{55.2}{46\times9}$$

$$\text{Mass}_{NO_2} = \frac{55.2}{46\times9}\times46$$

$$\text{Mass\%} = \frac{55.2/9}{55.2}\times100 = \boxed{\frac{100}{9}\%}$$

H.W.

Q. If any element $^{A}_{Z}X$ contains 2 isotopes $^{A_1}_{Z}X$ and $^{A_2}_{Z}X$ prove that

$$\text{Mol\% of } A_1 = 100 \times \left(\frac{M_{avg} - M_2}{M_1 - M_2}\right)$$

Sol. Let we have 100 moles sample

Moles of $A_1 = x$

$A_2 = 100 - x$.

$M_1 = xA_1$

$M_2 = A_2(100-x)$

$$M_{avg} = \frac{A_1x + A_2(100-x)}{100}$$

Mole% of $A_1 = x\%$

$$100\left(\frac{M_{avg} - M_2}{M_1 - M_2}\right) = 100($$

$$100\,M_{avg} = A_1x + A_2100 - A_2x$$

$$100(M_{avg} - A_2) = (M_1 - M_2)n_1$$

$$\frac{M\%_A M_1 + M\%_B M_2}{100} = M_{avg}$$

$$M\%_A M_1 + (100 - M\%_A)M_2 = M_{avg}100$$

$$M\%_A(M_1 - M_2 = (M_{avg} - M_2)100$$

$$M\%_A = 100\frac{(M_{avg} - M_2)}{M_1 - M_2}$$

Q. an inorganic salt has

$Na = 29.11\%$ $S = 40.51\%$ $O = 30.38\%$

find empirical formula -

In 100g sample -

Moles-Na $= \frac{29.11}{23} = 1.26$

$S = \frac{40.51}{32} = 1.26$

$O = \frac{30.38}{16} = 1.8$

Na : S : O $\boxed{2:2:3}$

$\boxed{Na_2S_2O_3}$

LAWS OF CHEMICAL COMBINATIONS

1) **Law of CONSERVATION OF MASS** :-

given by **Lavoisier.**

$A + B \rightarrow C$

Reactant (A, B) ↓ Product (C).

Law :- ★ Initial Mass of Reactants = Final Mass of Products.

In closed vessel. ⟵ ★ { Total Mass of it Before Reaction = Total Mass of it after Reaction }

★ TOTAL MASS INVOLVED IN REACTION REMAINS CONSTANT,

★ MASS NEITHER CREATED NOR DESTROYED IN A REACTION.

★ Law **not** applicable for NUCLEAR REACTION, where mass of reactant may not convert into energy.

(2) **Law of DEFINITION PROPORTION or DEFINITE COMPOSITION** :- given by PROUST

A compound has fix composition.

Compound CO_2 →

Mass% of C = 12/44 × 100

Mass% of O = 32/44 × 100

$\boxed{\text{Mass ratio} = 3:8}$

★ Mass% of elements in a compound is fixed and does not depend on method of preparation and source where it has been collected.

★ When we apply this law: isotopes of elements must be same.

(3) **Law of MULTIPLE PROPORTION** :- Given by DALTON

We consider **two** elements formed forming more than one product.

★ C, O

	CO	:	CO_2
★ If Mass of C fixed	12 : 16 g g	:	12 : 32 g g

Mass ratio of O in both compound (simple ratio) Natural NO. : 16 : 32

$\boxed{1:2}$

★ NO, NO_2, N_2O_5

If mass of N is fixed 14g then ratio of oxygen:

$NO : NO_2 : N_2O_5$

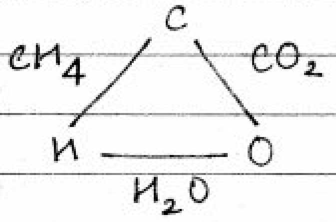
16 : 32 : 40

$\boxed{2:4:5}$

when two ~~or more~~ elements combine and formed two or more compounds then the mass of element combine with fix mass of other element bear a simple ratio.

(4) **Law of RECIPROCAL PROPORTION** : given by RICHTER

★ three elements are considered.



★1) select two elements — O and H
★2) Third element - C

★3) when both elements reacted separately with third element.

CO_2 : CH_4

Mass of O = 3~~6~~2 : 4

$\boxed{8:1}$

Mass ratio of O and H - $\boxed{8:1}$

★4) when both elements combined directly :

H_2O_2 / H_2O

Mass ratio of O and H.

$\boxed{8:1}$ or $\boxed{16:1}$

★5) It must be a multiple or same ratio

when two elements combined with fixed mass of third element, the ~~third~~ mass ratio of both element is either same or bear a simple ratio with a mass ratio of both element, when combined directly with each other.

GAY-LUSSAC'S गे-लुसाक के Law

Law of combining gases.

* applicable for ideal gases.

$\boxed{PV = nRT}$

at same temp and pressure:

$\boxed{V \propto n}$

gases combined in simple volume ratio when measured at same temp.
[also applicable for product when it is also gas]

[gaseous reaction]

	$H_2(g)$ +	$Cl_2(g)$ →	$2HCl$
Before reaction	1 vol.	1 vol.	-
After reaction	-	-	2 vol.

$$\begin{array}{ccccc} 2H_2(g) + & O_2(g) & \xrightarrow{\text{at } T > 100^\circ c} & 2H_2O(g) & (+822.4\ l) \\ 1\text{ vol} & 1\text{ vol} & \big\downarrow & 2\text{ vol} & \\ & & \text{at } T < 100^\circ c & & \end{array}$$

negligible vol. (18 ml)

AVOGADRO'S LAW :-

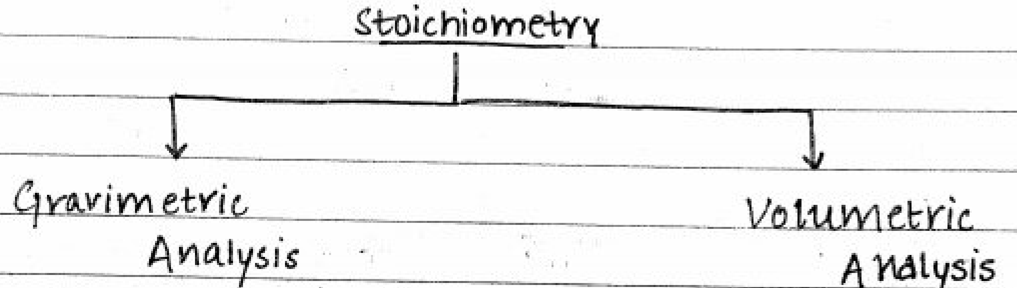
// Same volume of gases measured at same temperature and pressure contains same no. of molecules.

(using $PV = nRT$)

Stoichiometry

Stoichio : Elements
Metry : Measurement.

It is quantitative analysis (based on amount (mass/volume) of substance) of reactants and product in chemical reaction.



(i) Mass - Mass relationship
(ii) Mass - Volume "
(iii) Vol. - Volume "

(Titrimetric Analysis)

Balanced Reaction :

$$aA + bB \rightarrow cC + dD$$

A & B → Reactants
C & D → Products.

a, b, c, d = stoichiometric coefficient.

★ Vol. of CO_2 produced at 1atm & 273 K by 8g of O_2 completely reacted by carbon.

Sol. $C + O_2 \longrightarrow CO_2$

No. of moles of $O_2 = \frac{8}{32} = \frac{1}{4}$

∴ No. of moles of $CO_2 = \frac{1}{4}$

Volume at 1atm and 273 K $= \frac{1}{4} \times 22.4 = \boxed{5.6 L}$

	$C(s)$ +	$O_2(g)$ →	$CO_2(g)$
★ Stoichemetric amount	1 mol	1 mol	
	—	—	1 mol
★ Excess amount	$\boxed{5 mol}$	$\boxed{8 mol}$	
	—	3 mol	5 mol
	Limiting Reagent	Excess Reagent.	

★ **Excess Reagent**: Substance present in excess amount (more than stoichiometric amount).

or substance left after reaction.

★ **Limiting Reagent**: substance present in limiting amount.

- ★ L.R. is completely consumed in reaction
- ★ Product should be formed on basis of L.R.

★★ $3A + 2B \rightarrow 4C$

5mol 3mol

(Excess) (L.R.)

$5 \ast \frac{9}{2} = \frac{1}{2}$ mol — 6 mol.

$3A + 2B \rightarrow 4C$

$\underset{\text{mol}}{x}$ $\underset{\text{mol}}{y}$

If both present in stoichiometric amount :

$$\boxed{\frac{2}{3}x = y} \longrightarrow \boxed{\frac{x}{3} = \frac{y}{2}}$$

3 mole of A reacted with 2 mol B

x mole A —— $\boxed{\frac{2}{3}x}$ mol B.

$$\boxed{\frac{\text{Moles Given}}{\text{stoichiometric coefficient}}}$$

Steps to find L.R. :

(1) Balance the given reaction

(2) calculate moles of every element (reactant)

(3) Divide moles of reactant with their stoichiometric coefficients

(4) Reactant which has minimum value of $\boxed{\frac{\text{Moles}}{\text{S.C.}}}$ is L.R.

Percentage yeild :

$$\% \text{ yield} = \frac{\text{Actual yield of reaction}}{\text{Theoritical yield of reaction}} \times 100$$

$$= \frac{\text{Actual moles of product formed}}{\text{Moles of p. should be farmed as per stoichiometry}} \times 100$$

$$= \frac{\text{Actual mass of product formed}}{\text{Theoritical mass should be formed as per s.}} \times 100$$

* For normal reaction :

$$\boxed{\% \text{ yeild} \le 100}$$

Note * If % yield not given consider 100%.

Q $MgCO_3(s) \longrightarrow MgO(s) + CO_2(g)$

84 — 40

168 — 80

$\frac{90}{100} \times 80 = 72$

80

★ **Residue** :

Substance left (at the bottom of vessel) after reaction is called residue.

$$MgCO_3 \xrightarrow{\Delta} \underbrace{\boxed{MgO}}_{\text{Residue}} + CO_2 \uparrow$$

★ Impure sample of x :

Mass of Impure sample = $\boxed{\text{Mass of pure substance + Mass of impurity}}$

H.W.

Q.(1) What mass of P_4O_{10} was obtained when 1.24 g of P_4 reacts with 1.92 g of O_2.

Q.(2) In above problem only 1.42 g P_4O_{10} is obtained then calculate % yield of reaction.

Q.(3) PCl_5 reacts with water and forms H_3PO_4 and HCl according following reaction:

$$PCl_5 + 4H_2O \longrightarrow H_3PO_4 + 5HCl$$

If .4 moles of PCl_5 combines with 1.8 moles of water then find

(i) limiting reagent
(ii) Mass of HCl and H_3PO_4 produced
(iii) If mass of HCl formed is 65.7 g then calculate % yield.

Q.(4) calculate mass of $KClO_3$ required to produce 1.49 g KCl. ($KClO_3 \longrightarrow KCl + O_2$) – not balanced.
if % yield of reaction is:
ⓐ 100%
ⓑ 40%.

Q.(5) 2.5 g of an impure sample of $CaCO_3$ on treatment with excess HCl produce 0.88 g CO_2. What is the % purity of sample

Q.(6) On heating 80% pure $CaCO_3$ sample produced 1.76 g of CO_2. What mass of $CaCO_3$ is heated?

Sol.(1) $P_4 + 5O_2 \longrightarrow P_4O_{10}$

S.C. 1mol 5mol

N.O. of moles of $P_4 = \frac{1.24}{124} = .01$

No. of moles of $O_2 = \frac{1.92}{32} = .06$

L.R. ⇒ P_4

∴ Moles of $P_4O_{10} = .01$

Mass of $P_4O_{10} = .01 \times 284$

$= \boxed{2.84g}$ Ans

Sol(2) % yield = $\frac{\text{Actual mass}}{\text{Real Theoritical Mass}} \times 100$

$= \frac{1.42}{2.84} \times 100$

$= \boxed{50\%}$ yield Ans.

Sol(3) $PCl_5 + 4H_2O \longrightarrow H_3PO_4 + 5HCl$

S.C. → 1mol 4mol

.4mol 1.8mol

(i) Limiting Reagent — ✓ $\boxed{PCl_5}$ — .4 Ans

H_2O — $\frac{1.8}{4} = .45$

(ii) As L.R. is PCl_5

∴ Moles of HCl = $5 \times .4 = 2$

Mass of HCl = $2 \times 36.5 = \boxed{73g}$

Moles of $H_3PO_4 = .4$

Mass — $= .4 \times 98 = \boxed{39.2g}$

(iii) % yield = $\frac{65.7}{73} \times 100 = \boxed{90\%}$ Ans.

Sol. 4.

$2KClO_3 \longrightarrow 2KCl + 3O_2$

2mol

(a) If % yield = 100

Mass of 2mol $KClO_3$ – 245g

Mass of 2mol KCl – 74.5g × 2

$2 \times 74.5g$ — 245g

1.49 — $\frac{245}{74.5\times 2} \times 1.49 = \frac{245}{50} = \frac{\boxed{4.9g}}{2} = 2.45g$

(b) If 40% yield

$\frac{40}{100}\ast = \frac{1.49}{x}$

$x = \frac{149}{40\times 2} = \frac{3.725}{2}g$

Mass of $KClO_3$ required — $\frac{245}{74.5} \times \frac{3.725}{2}$

$= \frac{490}{40} = \frac{\boxed{12.25g}}{2}$ Ans. = 6.125g

Sol. 5

$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$

1mol 2mol

1mol $CaCO_3$ has mass = 100g

1 mol CO_2 has mass = 44g

44g — 100g

.88g — $\frac{100}{44} \times .88 = 2g$

∴ % Purity $= \frac{2}{2.5} \times 100 = \boxed{80\%}$ Ans

Sol. 6.

$$CaCO_3 \rightarrow CaO + CO_2$$

44g given by 100g of $CaCO_3$

1.76g given by $\frac{100}{44} \times 1.76 = 4g$

∴ Mass of $CaCO_3$ (impure) = $\frac{400}{80} = \boxed{5g}$ Ans.

Q. If an impure sample of $CaCO_3$ on treatment with HCl produces 1.76g CO_2 with 80% yield and % purity of $CaCO_3$ is 90%.

Sol

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$$

$$\frac{80}{100} x = 1.76$$

$$x = \frac{1.76 \times 100}{80}$$

$$100 — 44$$

$$\frac{1.76 \times 100}{80} — \frac{44 \times 1.76}{80}$$

$$\frac{100}{44} \times \frac{1.76 \times 100}{80}$$

$$= 5g$$

$$\frac{90}{100} x = 5$$

$$x = \frac{50}{9} = \boxed{5.5g}$$

Q. calculate mass of ~~of~~ Agcl formed when 95% pure 10g of $MgCl_2$ react with $AgNO_3$. % yield of reactn is 80%.

sol. $MgCl_2 + 2AgNO_3 \rightarrow 2AgCl + Mg(NO_3)_2$

$\frac{95}{100} \times 10$

9.5 g

95 — 287

9.5 — $\frac{287}{95} \times 9.5$

$= \boxed{28.8}$

$\frac{80}{100} \times 28.8$

$= 4 \times 5.65$

$= \boxed{22.60g}$ ans.

Mixture Analysis

★ Heating of carbonates :

$\boxed{\text{Ist group}}$:→

★ Only $\boxed{Li_2CO_3}$ decomposes : $Li_2CO_3 \xrightarrow{\Delta} Li_2O + CO_2$

$Na_2CO_3 \xrightarrow{\Delta}$ No reaction

But

$NaHCO_3 \xrightarrow{\Delta} Na_2CO_3 + CO_2\uparrow + H_2O\uparrow$

$\boxed{\text{II}^{nd}\text{ group}}$: all decompose

★ $CaCO_3 \xrightarrow{\Delta} CO_2 + CaO$

Q. 10g of $MgCO_3$. and Na_2CO_3 heated CO_2 liberated which has 2.463L at STP

(i) Wt% of $MgCO_3$

(ii) Wt of residue. 1.6g

sol. $MgCO_3 \longrightarrow MgO + CO_2\uparrow$

× $Na_2CO_3 \longrightarrow Na_2O + CO_2$ (No reaction)

.1 mol of MgO

(a) wt of $MgCO_3$ → $.1 \times 84$
$= 8.4$

(b) Wt of residue
$=$ Wt. of Na_2CO_3 + Wt. of MgO
$= 1.6 + .1 \times 40$
$= \boxed{5.6\ g}$

$\boxed{\text{Q.}}$ 26.8 g sample of $MgCO_3$, $CaCO_3$ and $NaHCO_3$ suffered loss of 11.9 gram (Δ). The residue on treatment with HCl gave 1.12 L at 1atm 273K

(i) Moles and mass of each subs.

(ii) Mass % of $CaCO_3$ in origin mixture.

sol.

$$MgCO_3 \xrightarrow{\Delta} MgO + CO_2\uparrow$$

$$CaCO_3 \xrightarrow{\Delta} CaO + CO_2\uparrow$$

3) $$2NaHCO_3 \xrightarrow{\Delta} Na_2CO_3 + CO_2\uparrow + H_2O\uparrow$$

$$Na_2CO_3 + HCl \rightarrow 2NaCl + H_2O + CO_2\uparrow$$

$$4 \times \cancel{3} \times 44(x+y+z) = 11.9$$
$$+ \quad 18$$

1·12 L means ·05 mole

$\Rightarrow$ ·05 mole means ·1 mole of $NaHCO_3$
(23+1+60)

a) Moles of

Wt. of $CaCO_3 + MgCO_3 \longrightarrow 26.8 - .1 \times 84$
$= 18.4\,g$.

$$3 \times 44\,(x + y + .05) + .05 \times 18 = 11.9$$

$$x + y + 0.05 = \frac{11 \times 4}{176} = \frac{1}{4}$$

$$x + y = .2$$

$$84x + 100y = 18.4$$

$$16y = 1.6$$

$$y = .1$$

$$x = .1$$

∴ Mass of			Mole
$NaHCO_3$	—	8.4	.1
$MgCO_3$	—	8.4	.1
$CaCO_3$	—	10	.1

(b) Mass % of $CaCO_3$ → $\frac{10}{26.8} \times 100$

$$= \frac{1000}{26.8} \approx \boxed{38\%} = 37.3\%$$

Q. [Good Problem]

A mixture of NaI and NaCl treated with H_2SO_4 producted same of Na_2SO_4 as that of the orginal mixture. calculate mass% of NaI

Sol

$$2NaI + H_2SO_4 \longrightarrow Na_2SO_4 + 2HI$$
$$2NaCl + H_2SO_4 \rightarrow Na_2SO_4 + 2HCl$$

NaI — x gram
NaCl — 100-x gram

$$\frac{x}{2\times150} = \text{Moles of } Na_2SO_4$$

$$\frac{100-x}{2\times162.5} = \text{Moles of } Na_2SO_4$$

$$\left(\frac{x}{300} + \frac{100-x}{\underset{117}{325}}\right)\times 142 = 100$$

$$25x + 300\times100 = \frac{100}{142}\times300\times325$$

$$25x = 300\left(\frac{100}{142}\times325 - 100\right)$$

$$x = \frac{300\times100\times183}{142}$$

$$-183x = \frac{100}{142}\times300\times117 - 300\times100$$

$$= 300\left(\frac{100\times117}{142} - 100\right)$$

$$-183x = \frac{300\times100\times-25}{142\times-183} = \boxed{28.86g}$$

Q. 2g mixture of $CuCl_2$ and $CuBr_2$ was dissolved in water and mixed with 1.8 gram AgCl.
after reaction the mixture of AgCl and AgBr was filtered, washed and dried.
Its Mass was found to be 2.25g
% of $CuBr_2$ in original mixture?

sol.

$$2AgCl + CuBr_2 \rightarrow 2AgBr + CuCl_2$$
$$x\,g. \qquad (2.25-1.8+x)\,g$$

$$\frac{x}{2\times143.5} = \frac{.45+x}{2\times188}$$

$$(2\times188)x = (.45+x)287$$
$$396x - 287x = .45\times287$$
$$\cancel{1}89x = .45\times287$$
$$x = \frac{.45\times287}{\cancel{1}89}$$

Moles of $\cancel{2}AgCl$ — $\frac{.45\times287\times2}{\cancel{1}89\times287}$

Moles of $CuBr_2$ — $\frac{.45\times287}{89\times143.5}\,\frac{1}{2} = \frac{.0101}{2}$

$$\frac{.0101}{2}\times226.3$$

Mass of $CuBr_2$ — 1.13

$$\text{Mass \%} = \frac{1.13}{2}\times100$$

$$= \boxed{56.5\,\%}$$

Q. When 2.86 g of a mixture of C_4H_8 (butene) & C_4H_{10} (Butane) Was burned in excess of Oxygen, 8.8 g of CO_2 and 4.14 g H_2O. Obt. Calculate Mass% of butane in original mixture.

Taking wt -

Sol. x

$$C_4H_{10} + O_2 \rightarrow 4CO_2 + 5H_2O$$

$$C_4H_8 + O_2 \rightarrow 4CO_2 + 4H_2O$$

$$\frac{x}{58} \times 4 = \cancel{4}\ \frac{8.\cancel{8}y}{44}$$

$$x = \cancel{4}\ \frac{58 \times .2}{4}$$

$$= 2.9$$

$$\frac{2.86 - x}{56} = \frac{8.8 - y}{44}$$

$$\frac{8.8}{44} - \frac{2.86 - x}{56} \times \frac{44}{44} = \frac{y}{44}$$

$$\frac{4x}{58} = .2 - \frac{2.86 - x}{56}$$

$$\frac{4x}{58} = \frac{11.2 - 2.86 + x}{56}$$

$$214x = 8.\overset{4}{\cancel{5}}4 \times 58$$

$$x = \frac{8.54}{224} \times 58$$

$$\text{Mass \%} = \frac{.03 \times 58}{2.86} \times 100 = .03 \times 58$$

Taking Moles :

★ When two reactants (elements) react with each other and forms two product.

Reactants	H_2 & O_2	H_2O	H_2O_2
	S & O_2	SO_2	SO_3
	Na & O	Na_2O	Na_2O_2
	C & O	CaO	CO_2

—x—

$C + \frac{1}{2}O_2 \rightarrow CO_{\cancel{2}}$ $\qquad$ $C + O_2 \rightarrow CO_2$

★ If we fix ~~mass~~ of
moles of $C = 1$

Conditions :-

(I) If $n_{O_2} < 0.5$. then only $CO_{\cancel{2}}$ is formed and C left after reaction

(II) If $n_{O_2} = 0.5$ then only CO is formed

(III) If $1 > n_{O_2} > .5$ then both CO and CO_2 are formed

(IV) If $n_{O_2} = 1$ then only CO_2 is formed

(V) If $n_{O_2} > 1$ then CO_2 is formed and O_2 left.

. * If moles of O_2 exist between .5 to 1 then both (CO and CO_2) formed.

$$\left\{ \begin{array}{l} C + \frac{1}{2}O \rightarrow CO \\ CO + \frac{1}{2}O_2 \rightarrow CO_2 \end{array} \right\}$$

* Suppose $n_{O_2} = .8$ mol

Mol. of C = 1

Mol of O_2 = .8

$C + \frac{1}{2}O_2 \rightarrow CO$

$x \quad \frac{x}{2}$

$- \quad - \quad x$ mol.

$C + O_2 \rightarrow CO_2$

$(1-x) \quad (1-x)$

$- \quad - \quad (1-x)$

Moles of O_2 consumed . $\frac{x}{2} + 1 - x$

$= .8$

$\therefore \boxed{x = .4}$

Moles of CO formed = $\boxed{.4}$

" CO_2 formed = $\boxed{.6}$

Q. Calculate mass of substance after reactn. If 96 g O_2 react with

(a) 24 g C

Moles of C = 2

Mols of O_2 = 3

Mass after reactn. — 2 Mol CO_2 + 1 mol O_2

= (88+32) g . 120 g.

(b) 48 g C

$n_C = 4$

$n_{O_2} = 3$

2 Mol CO_2 + 1 Mol CO_2 3 Mol CO_2 + 1 mol C

= (56+44) = 100 g. 56 88

(c) 84g of C

Mole of C = 7

n_{O_2} = 3

$\frac{3}{3.5}$ mol of CO = $\frac{3}{3.5} \times 28$

= $\frac{84}{3.5}$ = $\boxed{24g}$

3Mol CO_2 + 4Mol C

12g

(d) 72g of C

n_C = 6

n_{O_2} = 3

3Mol CO_2 +

168g

Q sodium reacts with oxygen and forms Na_2O and Na_2O_2, (Na_2O converts to Na_2O_2)
Calculate molar composition of final mixture if 4.6 kg Na reacts with 1792 L of O_2 at 1 atm and 273 K

None reactants left.

$2Na + \frac{1}{2}O_2 \rightarrow Na_2O$ $2Na + O_2 \rightarrow Na_2O_2$.

Sol. Mol of O_2 $\frac{1792}{22.4} = 80$ mol

Mol of Na. $\frac{4600}{23} . 200$

$2Na + \frac{1}{2}O_2 \rightarrow Na_2O$ $2Na + O_2 \rightarrow Na_2O_2$

x $\frac{x}{4}$ $(200-x)$ $\frac{-}{2}$

$\frac{x}{2}$ $\frac{200-x}{2}$

$$\frac{x}{4} + \frac{200-x}{2} = 80$$

$$-x + 400 = 320.$$

$$-x = -80$$

$$x = 80$$

Mol of Na_2O = 40

Mol. of Na_2O_2 = 60.

—x—

$\boxed{\text{POAC}}$

★ PRINCIPLE OF ATOM CONCENTRATION

→ No. of atoms of every element remains conserved in chemical reaction.

→ No. of moles of every element also remains constant.

$$CaCO_3 \longrightarrow CaO + CO_2$$

$\boxed{\text{POAC : Ca}}$ Moles of Ca atom in $CaCO_3$ = Moles of Ca atom in $CaCO_3$

$\boxed{\text{POAC : } O_2}$ Moles of O_2 in $CaCO_3$ = Moles of O in CaO + Moles of O in CO_2.

—x—

Q. 55.2 g of K_2CO_3 was reacted by a series of reagents. convert all the carbon to $K_2Zn_3(Fe(\overset{C}{\cancel{Z}}N)_6)_2$ (Molar Mass) = 698g)

Calculate the wt of the product.

$$K_2CO_3 \longrightarrow K_2Zn_3(Fe(CN)_6)_2$$

Sol. $\boxed{\text{POAC : C}}$ Moles of C in K_2CO_3 = Moles of C in product

$$\frac{55.2}{138} \times 1 = \frac{x}{698} \times 12$$

$x =$

Q $CaC_2 + H_2O \longrightarrow CaO + C_2H_2$

$C_2H_2 + H_2 \rightarrow C_2H_4$

$nC_2H_4 \longrightarrow (C_2H_4)_{An}$

Mass of ↑ produced from 6.4 kg CaC_2

sol. $\boxed{\text{POAC : C}}$ Moles of $CaC_2 \rightarrow \frac{6400}{64} =$ 100 moles.

Mass of product = $\boxed{2800\ g}$. 2.8 Kg.

Q. How much gases (in L) will be produced at 1 atm 0°C when 9 gram oxalic acid was heated with conc. H_2SO_4 (dehy. agent)

$(COOH)_2 \longrightarrow CO + CO_2 + H_2O$

sol. Moles of $(COOH)_2$ – $\frac{9}{90}$ = .1 mol.

.1 mol – CO 2.24 L

.1 mol – CO_2 2.24 L

$\boxed{4.48\ L}$ Ans.

Sequential Reaction

$A \rightarrow B \rightarrow C$

Reaction in which the product of on re^n is used as reactant of another reaction.

$A + 2B \rightarrow 3C + D$

★ **Factor label Method**

$A \rightarrow 2B$

x gm

$B + C \rightarrow 3D$

?

$$(x \text{ gm } A)\left(\frac{1 \text{ mol } A}{M_A \text{ gm } A}\right)\left(\frac{2 \text{ mol } B}{1 \text{ mol } A}\right)\left(\frac{3 \text{ mol } D}{1 \text{ mol } B}\right)\left(\frac{M_D}{1 \text{ mol } D}\right)$$

—×—

★ Calculate wt of H_2SO_4 formed from FeS_2 acc. to given reaction.

$$FeS_2 + \frac{5}{2}O_2 \xrightarrow{90\%} FeO + 2SO_2$$
$$SO_2 + \frac{1}{2}O_2 \xrightarrow{90\%} SO_3$$
$$SO_3 + H_2O \xrightarrow{90\%} H_2SO_4$$

If 12g FeS_2 is taken.

Sol. On balancing USING factor label Method.

$$\left(12g\ \cancel{FeS_2}\right)\left(\frac{\overset{FeS_2}{\cancel{1mol}}}{120g}\right)\left(\frac{2\,\cancel{mol}\ SO_2}{1\,\cancel{mol}\ \cancel{SO}FeS_2}\right)\left(\frac{1\,\cancel{mol\ SO_3}}{1\,\cancel{mol\ SO_2}}\right)\left(\frac{1\,\cancel{mol}\ H_2SO_4}{1\,\cancel{mol\ SO_3}}\right)$$
$$\left(\frac{98gm}{1\,\cancel{mol\ H_2SO_4}}\right)$$

$$= \frac{98 \times 2 \times 12}{120} = \boxed{19.6gm}$$

If yield is 90% in each reaction:

$$\boxed{19.6 \times \left(\frac{9}{10}\right)^3}$$

When atoms is conserved and yield is 100%, then.

Using [POAC],

Moles of s in FeS_2 = Moles of s in H_2SO_4

$$\left(\frac{12}{120}\right) \times 2 = \left(\frac{y}{98}\right) \times 1$$

$$y = \boxed{19.6\ gm}$$

* $KClO_4$ may be made from following series of reⁿ.

$$Cl_2 + 2KOH \rightarrow KCl + KClO + H_2O$$
$$3KClO \rightarrow 2KCl + KClO_3$$
$$4KClO_3 \rightarrow 3KClO_4 + KCl$$

If 568 g of Cl_2 is used then,

(a) Calculate mass of $KClO_4$
(b) Wt of KClO and $KClO_3$ formed.
(c) Wt. of KCl obtained during reaction
(d) calculate wt. of Cl_2 required if 69.25 g $KClO_4$ produced.

sol.

(a) $$(568g\ Cl_2)\left(\frac{1mol}{71gCl}\right)\left(\frac{1mol\ KClO}{1mol\ Cl_2}\right)\left(\frac{1\ KClO_3}{3mol\ KClO}\right)\left(\frac{3\ mol\ KClO_4}{4\ mol\ KClO_3}\right) \times \left(\frac{138.4}{1mol\ KClO_4}\right)$$

$$\frac{8 \times 3 \times 138.5}{12} = \boxed{276.8g}$$
$$= \boxed{277g}$$

(b) KClO $$\frac{568}{71} \times \frac{1}{8} \times 90.5$$ $$= \boxed{724g}$$

$KClO_3$ $$\frac{568}{71} \times \frac{1}{3} \times 122.5$$ $$= \frac{980}{3} = \boxed{326.6g}$$

(c) $$\frac{568}{71} \times \frac{2}{3} \times 74.5 = \frac{16}{3} \times 74.5$$
$$= 16 \times 24.8$$
$$= \boxed{396.8g}$$

$$\begin{pmatrix} 8 \times 74.5 \\ \frac{2}{3} \times 74.5 \end{pmatrix} +$$

$$= 14 \times 74.5 = (10403)\ gm.$$

(d)

$$568 — 277$$

$$\frac{568}{\cancel{277}_{4}} \times \cancel{69.25} — 69.25$$

$$= \boxed{142 gm} \text{ ans}$$

Other method :

$$(69.25\,g\,KClO_4)\left(\frac{1\,mol\,KClO_4}{138.5\,g}\right)\left(\frac{4\,mol\,KClO_3}{3\,mol\,KClO_4}\right)\left(\frac{3\,mol\,KClO}{1\,mol\,KClO_3}\right)\left(\frac{1\,mol\,Cl_2}{1\,mol\,KClO}\right)\left(\frac{71\,g}{1\,mol\,Cl_2}\right)$$

$$= \frac{69.25 \times 71 \times 4 \times 3}{138.5 \times 3}$$

$$= \boxed{142 gm}$$

—x—

Q. Nitric acid can be produced from a series of reactions :

$$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$$

$$2NO_{\cancel{2}} + O_2 \rightarrow 2NO_2$$

$$3NO_2 + H_2O \rightarrow 2HNO_3 + NO$$

(a) Calculate the wt NH_3 required to produce 1134g

(b) HNO_3 if % yield of 1^{st}, II, III reactn are. HNO_3 90, 80, 85%

(b) If 1.7g NH_3 is used for this reaction, calcult HNO_3 produced

Sol. (a) $\cancel{\frac{4}{17}}$ $\frac{1134}{63} \times \frac{3}{2} \times \frac{2}{2} \times \frac{4}{4} \times 17$

$$= \frac{567}{63} \times 3 \times 17$$

$$= \frac{567 \times 17}{21} \times \frac{9\cancel{0}}{10\cancel{0}} \times \frac{8\cancel{0}}{10\cancel{0}} \times \frac{85}{100} \times$$

$$= \frac{567 \times 17}{21} \times \frac{10000}{9 \times 8 \times 85} = \frac{63 \times 10000}{21 \times 5 \times 8}$$

$$= 3 \times 250 = 750g.$$

(b) $1.7 \times \frac{\cancel{4}}{17} \times \frac{4}{4} \times \frac{2}{2} \times \frac{2}{3} \times 63$

$$= \frac{20 \times 63}{300} = \boxed{4.2g}$$

—x—

- ★ concentration terms for liquid or aqueous solution :

solute + solvent ⇒ liquid solution
(In small amount) (in large amount)

- ★ Same physical state of solvent & solution.

$C_6H_6(l) + C_6H_5CH_3(l)$

↓

Liquid solⁿ of benzene and Toluene.

- ★ $C_6H_{12}O_6(aq) \rightarrow C_6H_{12}O_6(s)$ - solute
(aqueous) $H_2O(l)$ - solvent

$NaCl(l)$ = Melted NaCl
$NaCl(aq.)$ = NaCl dissolved in H_2O.

★ Wt % of solute (% by wt) $= \dfrac{\text{Wt of solute}}{\text{Wt of (solute + solvent)}} \times 100$

$$= \frac{W_{solute}}{W_{solution}} \times 100$$

★ Mole % of solute $= \dfrac{\text{Moles of solute}}{\text{Moles (solute+solvent)}} \times 100$

$$\underbrace{\hspace{3cm}}$$

For all $\boxed{\text{homogenous}}$ solution

★ $\% \frac{W}{W} = \dfrac{\text{Wt of solute}}{\text{Wt of sol}^n} \times 100$

Ex: 5% of KOH (aq) sol^n

→ 100g sol^n has 5g KOH.

★ $\% \frac{W}{V} = \dfrac{\text{Wt. of solute}}{\text{volume of sol}^n} \times 100$

Ex. – 10% $\left(\frac{W}{V}\right)$ HCl (aq.) sol^n

100ml sol^n has 10g HCl

—x—

★ Q. convert 10% $\left(\frac{W}{W}\right)$ sol^n into $\% \left(\frac{W}{V}\right)$

$\text{density}_{\text{sol}^n} = 1.05\ g/ml$

Sol. 100 gm sol^n contains 10g KOH

$\frac{100}{1.05}$ ~~g~~ ml of ~~KOH~~ sol^n cont. 10g KOH

$$\therefore \% \left(\frac{W}{V}\right)_{KOH} = \frac{10 \times 1.05}{100} \times 100 = 10.5\%$$

★ Q. convert 20% $\left(\frac{W}{V}\right)$ HCl into %$\left(\frac{W}{W}\right)$ (Density = 1.15g/ml)

100 ml contains 20 gm HCl
1.15 gm contains 20 gm HCl

$$\% \frac{W}{W} \text{ HCl} = \frac{20}{115} \times 100$$

$$= \frac{400}{23} = 17.\ \%$$

★ $\%\left(\frac{V}{V}\right) = \frac{\text{volume of solute}}{\text{volume of sol}^n} \times 100$

Q.X.

Q. 8% $\left(\frac{V}{V}\right)$ C_2H_5OH (aq) solⁿ
100 ml solⁿ contains 8 ml C_2H_5OH
? convert %$\left(\frac{W}{W}\right)$ given density alcohol = d_1 gm/ml
d_{H_2O} = d_2 gm/ml

Sol $\frac{8d_1}{8d_1 + (100-8)d_2} \times 100$

Q. convert 10% $\left(\frac{W}{W}\right)$ in same quest. in $\left(\frac{V}{V}\right)$

100gm sol. contains 10 gm C_2H_5OH
90gm Water

Ans → $\frac{\left(\frac{10}{d_1}\right)}{\left(\frac{10}{d_1}\right) + \left(\frac{90}{d_2}\right)} \times 100$

Parts Per Million

★ Concentration of solute in PPM :-

★ Con^n in PPM = $\frac{\text{Wt of solute}}{\text{Wt of solution}} \times 10^6$ } for not very dilute sol^n

★ Con^n in PPM = (Wt% of solute) × 10^4 }

Note :- For dilute solutions,

Wt. of solution ≃ Wt. of solvent

★ Con^n in PPM = $\frac{\text{Wt of solute}}{\text{wt of solvent}} \times 10^6$

★ Con^n in PPB = (Wt% of solute) × 10^9
(Parts Per Billion)

Q. 20gm solute is present with 200 ml solvent (d = .9 $\frac{gm}{ml}$to) calculate mass fraction of solute in solvention.

Q. Calculate the mass of HCl in 50 ml HCl (aq.) sol^n contains 40% by wt HCl (d = 1.2g/ml).

Q. What volume of 90% by wt C_2H_5OH sol^n (d = .8 must be used to prepare 150 cm^3 of 30% by wt C_2H_5OH sol^n (d = .96g/ml).

(1) Wt. of solvent $= 200 \times .9$

$= 180\text{gm}$

Mass fraction

$= \frac{20}{200}$

$= \boxed{.10}$

(2) Mass of solⁿ $= 50 \times 1.2$

Mass of solute $= .4 \times 1.2 \times 50$

$= \boxed{24\text{ gm}}$

(3) Let V ml solⁿ is required

$V \times .80 \times \frac{90}{100}$ = wt of C_2H_5OH

$= 150 \times .96 \times \frac{30}{100}$

$\boxed{V = 60\text{ml}}$.

— x —

MOLARITY (M) :-

$$\boxed{M = \frac{\text{No. of moles of solute}}{\text{volume of solution (in L)}}} \cancel{\times 100}$$

$$\boxed{M = \frac{n_{\text{solute}}}{V_{\text{solution}}\ (\text{ml})} \times 1000}$$

$$\left[M = \frac{W_{\text{solute}}}{\text{Molar mass of solute}} \times \frac{1000}{\text{volume}_{\text{sol}^n}\ (\text{ml})} \right]$$

MOLALITY (m) :-

$$\boxed{M = \frac{\text{No. of moles of solute}}{\text{Wt of solvent (in kg)}}}$$

$$\left[m = \frac{\text{No. of moles of solute}}{\text{Wt. of solvent (in g)}} \times 1000 \right]$$

↓ true

Remember:- Wt of solution = Wt. of solvent + Wt. of solute.

Q. Molarity (M) of 20% $\left(\frac{W}{V}\right)$ NaOH solⁿ?

Sol. 100 ml solⁿ has 20 gm NaOH

∴

$$M = \frac{1000}{2 \times 100} = \boxed{5M}$$

Q. Molality (m) of 20% $\left(\frac{W}{W}\right)$ NaOH solⁿ

Sol. 100gm solⁿ contains 20 gm NaOH.

$$m = \frac{1000}{2 \times 80} = \frac{100}{16} = 6.25\,m$$

Q. Mole fraction of sol~~vent~~ute in its 1 molal aq. solution

Sol. 1 molal solⁿ means $\frac{\text{1 mole solute}}{\text{1 kg solvent}}$

$$\frac{1}{1 + \frac{1000}{18}} = \boxed{\frac{18}{1018}} \quad (\because \text{Aqueous sol}^n)$$

Q. calculate molarity of 6.25 m (Molal) $NaOH_{(aq.)}$ solⁿ ($d_{sol.} = 1.2$ g/ml)

6.25 moles in 1000 g H_2O

Wt. of NaOH = 6.25 × 40 = 250~~0~~ g.

∴ Wt_{sol^n} = 250 + 1000 = (1250)g

∴ volume of solⁿ = $\frac{1250}{1.2}$ =

$\therefore$ Molarity of solution $= \dfrac{6.25}{1250} \times 1.2 \times 1000$

$= \dfrac{1.2 \times 1000}{200} = .006 \times 1000$

$= \boxed{6\ M}$

Q. Mole fraction of urea (NH_2CONH_2) in aq. solution is .02

calculate M and m

(given $d_{soln} = .99$ g/ml)

In 100 mole 2 mol of urea.

~~100~~ 98 mole of H_2O

Wt. of urea $= 2 \times 60 = 120g$

Wt. of $H_2O = 18(100-2) = 1800-36 = 17\overset{6}{4}4$ gm

Total wt. $= 1884$ gm

$V_{sol} = \dfrac{1884}{.99} =$

$\therefore \quad M = \dfrac{2}{1884} \times .99 \times 1000 = = \dfrac{.99}{\underset{4}{9\cancel{3}2}} \times 1000 = \dfrac{990}{\underset{4}{9\cancel{8}2}}$

$m = \dfrac{2}{1764} \times 1000 = \dfrac{1000}{882} =$

H.W.

Q. 40g NaOH was dissolved in water and formed 200 ml solⁿ. Calculate:

(i) $\% \left(\frac{W}{W}\right)$ (ii) $\% \left(\frac{W}{V}\right)$ (iii) M (iv) m (v) Mole fraction of solute

(vi) PPM of solute (given $d_{soln} = 1.5 g/ml$)

Sol. (i) Wt. of solution $= 200 \times 1.5$

$= 300 g$

$\% \left(\frac{W}{W}\right) NaOH = \frac{40}{300} \times 100 = \boxed{\left(\frac{40}{3}\right) \%}$

(ii) $\% \frac{W}{V} = \frac{40}{200} \times 100 = \boxed{20\%}$

(iii) $M = \frac{1}{200} \times 1000 = \boxed{5 M}$

(iv) $m = \frac{1}{260} \times 1000 = \boxed{\frac{100}{26} \%} = \boxed{3.9}\%$

(v) Mole fraction

Moles of $H_2O = \frac{260}{18}$

$$\frac{1}{1 + 260/18} = \frac{18}{\underset{78}{\cancel{260}}} = \boxed{\frac{9}{139}}$$

(vi) PPM of solute

$$= \frac{40}{300} \times 10^6 = \boxed{\frac{40 \times 10^4}{3} \%}$$

★ **DILUTION of a SOLUTION** :

addition of solvent ($H_2O(l)$)

$$M = \frac{n_{solute}}{V_{solution\,(L)}}$$

★ $\boxed{\begin{array}{l} M \times V(\text{in L}) = \text{Moles of solute} \\ M \times V(\text{in ml}) = \text{Millimoles of solute} \end{array}}$

Note
★★ In dilution process moles of solute remains constant.

$\therefore$ Moles of solute$_{\text{(before dilution)}}$ = Moles of solute$_{\text{(after dilution)}}$

$$\therefore \quad \boxed{M_1V_1 = M_2V_2}$$

M_1 : Molarity befor dil.
V_1 : Volume before dil. (initial)
M_2 : Molarity after dil.
V_2 : Volume after dil. (final)

where $\boxed{V_2 = V_1 + V_{H_2O}}$ — volume of water added.

$$\therefore \ M_1V_1 = M_2(V_1 + V_{H_2O})$$

—×—

↗ Molarity of a resultant when two or more than two solutions & a substance are mixed.

HCl (aq.) : M_1, V_1 (ml)
HCl (aq.) : M_2, V_2 (ml)

★ Molarity of resultant solution = $\frac{\text{Total m.moles of HCl}}{\text{Total vol. of sol}^n \text{ (ml)}}$

✓ $M_{sol^n} = \frac{M_1V_1 + M_2V_2}{(V_1+V_2)}$

★ ★ $M_{sol^n} = \frac{M_1V_1 + M_2V_2 + M_3V_3 \ldots\ldots}{V_1+V_2+V_3 \ldots\ldots} = \frac{\Sigma M_iV_i}{\Sigma V_i}$

↗ If two solutes are present in sol^n:

solutes : A & B
solvent : Water

$M_A = \frac{\text{Moles of A}}{\text{Volume sol}^n \text{(L)}}$ $M_B = \frac{\text{Moles of B}}{V_{sol^n}}$

Q. 100 ml of 50% $\left(\frac{W}{W}\right)$ $HCl_{(aq.)}$ sol^n (1.46g/ml) is mixed with 100 ml water. calculate molarity of final sol^n.

sol.

$Wt_{sol.} = 146g$

$Wt._{HCl} = \frac{50}{100} \times 146 = 73g.$

$n_{HCl} = 2$

$M_{final} = \frac{2}{200} \times 1000 = \boxed{10M}$

Q. How many ml of water should be added to 200 ml of .6 M HCl sol^n to dil. upto .2M.

sol.

M.

$n_{HCL} = \frac{.6 \times 200}{1000} =$

of $M = .2 = \frac{.6 \times 200}{1000 \times x} \times 1000$

$x = 600$

total volume

∴ $\boxed{600\text{ ml}}$ water must be added.

$\boxed{400\text{ ml}}$ ↗

Q. What vol. of 24.5% by wt H_2SO_4 sol^n (1.2g/ml) is required to prepare 1L of 1.5 M H_2SO_4 sol^n.

sol.

$n_{H_2SO_4} = 1.5$

∴ Wt. of $H_2SO_4 = 98 \times 1.5$

∴ $\frac{24.5}{100} \times x = 98 \times 1.5$

$x = \frac{98 \times 1.5}{.245}$

vol. required $= \frac{98 \times 1.5}{.245 \times 1.2} = \frac{98}{.49 \times 4} = \frac{24.5}{.49}$

$= \boxed{500\text{ml}}$

Q. 200 ml of .1 M HCl soln is mixed with 800 ml of .2 M HCl soln then M_{final}.

Sol. 1) milli n_{HCl} = 20

2) milli n_{HCl} = 160

$$M_{final} = \frac{180}{1000} = \boxed{.18\ M}$$

Q. 50 ml of 50% $\left(\frac{W}{W}\right)$ HCl soln (d= 1.4g/ml) is mixed with 100 ml of 70% $\left(\frac{W}{V}\right)$ HCl (aq) sol. and some amount of water is also needed in order to obtain 28% $\left(\frac{W}{W}\right)$ HCl soln (d= 1.25 g/ml).

What volume of water is added (in ml)?

Sol.

W_1 = 70g W_{1HCl} = 35 g

W_2 = 70 W_{2HCl} = 70 g

W_f =

$$28 = \frac{105}{W_f} \times 100$$

$$W_f = \frac{15}{4} \times 100 = 375 \text{ gm}$$

∴ W_{H_2O} = 270 gm

∴ V_{H_2O} = $\boxed{270\ ml}$

$$V_f = \frac{375 \times \cancel{1.25}}{1.25}$$

$$V_{H_2O} = 300 - 150$$

$$= \boxed{150\ ml}$$

Ans

—x—

Types of solution

Ideal solutions

sol^n of A (V_1)

sol^n of B (V_2)

$\boxed{V_f = V_1 + V_2}$

$\boxed{\Delta V_{mix} = V_f - (V_1 + V_2) = 0}$

Weak intermolecular forces

A.... B = A.... A = B.... B

Real solutions

$V_{mix} > 0$ or $V_{mix} < 0$

Volume of mix. may increase or decrease.

Weak intermolecular forces

(i) A.... B < A... A ; B... B

$V_f > \cancel{0} (V_1 + V_2)$, $\Delta V_m > 0$

(ii) A.... B > A... A ; B... B

$V_f < (V_1 + V_2)$, $\Delta V_m < 0$

Contraction in volume

expansion in volume

Both are negligible change

Remember ★

Eudio

If density not provided, then add volume
If density provided, then add mass and calculate volume.

{because then may be contraction or expansion)

★ Always Add Mass

Q. case I: If density$_{soln}$ is not provided

Q. If 10g of $H_2O(g)$ is mixed with 3.2g $CH_3OH(l)$ $(d = .891)_{ml}$ then find molarity.

sol. 10 ml H_2O

14 ml

Moles of $CH_3OH = \frac{3.2}{32} = .1$

$M = \frac{.1}{10+4} \times 1000 = \boxed{\frac{100}{14}}$.

Q. If density$_{soln}$ is ~~not~~ provided.

If 100ml of 1M H_2SO_4 solution (1.5g/ml) is mixed with 400 ml of water. Calculate molarity of final sol. (d = 1.5 g/ml)

sol. .27 M

Moles of $H_2SO_4 = .1$

$Mass_1 = 150g.$

$Mass_2 = 400g.$

$\therefore$ final volume $= \frac{550}{1.5}$ ml

$\therefore$ final molarity $= \frac{.1 \times 1.5 \times 1000}{550}$

$= \frac{15}{55} = \boxed{.27\,M}$

★ Types of electrolytes

:- Water as solvent

★(1) Strong electrolytes :

completely or almost completely ionised in water.

Ex. ★ Strong acids : HCl, HNO_3, $HClO_4$ etc.

★ strong bases : KOH, $NaOH$

★ All dissolved salts in water : KCl, $NaCl$, $NaNO_3$ etc.

★ Acid Base Neutralisation :

Acid (aq) + Base (aq.) ⟶ salt (aq.) + water

$HCl(aq.) + NaOH(aq.) \rightarrow NaCl(aq.) + H_2O(l)$

$HNO_3(aq.) + KOH(aq.) \rightarrow KNO_3(aq.) + H_2O(l)$

$Mg(OH)_2 + HCl\ aq. \rightarrow MgCl_2 + \cancel{H}\ H_2O$

$H_2SO_4 + 2KOH \rightarrow K_2SO_4 + 2H_2O$

★ Weak electrolytes

Partially ionized in water.

Weak acid :- CH_3COOH, C_6H_5COOH, $HCOOH$

Weak bases :- $NH_4OH(aq.)$, $C_6H_5NH_2(aq.)$

Acid : furnish H^+ ions (water)

Bases : furnish OH^- ions

↳ Ⓞⓡ accept H^+

$CH_3COOH(aq.)$ — c (concentration in mol/L)

$$\underset{c-x}{CH_3COOH(aq.)} \rightleftharpoons \underset{x}{CH_3COO^-(aq)} + \underset{x}{H^+(aq.)}$$

at equilibrium :

Important

★ Degree of dissociation (α) :
Fractional dissociation of reactants into products.

$\alpha = \frac{x}{c}$ % DOD = $\frac{x}{c} \times 100$

★ HCl : .1 M

$$\underset{—}{HCl(aq.)} \longrightarrow \underset{.1M}{H^+(aq.)} + \underset{.1M}{Cl^-(aq.)}$$

★ KOH : .02M

$$\underset{—}{KOH(aq)} \longrightarrow \underset{.02M}{K^+(aq.)} + \underset{.02M}{OH^-(aq)}$$

★ ★ If ionization of strong electrolyte is not provided then consider 100% ionization.

★ $AlCl_3(aq.) : 10^{-5}$ M

$$\underset{—}{AlCl_3} \longrightarrow \underset{10^{-5}M}{Al^{3+}} + \underset{3\times10^{-5}M}{3Cl^-}$$

Q. calculate conⁿ of K^+ and SO_4^{-2} in ·0.2 M of K_2SO_4 solution. consider 95% ionization of salt.

sol.

$$\begin{array}{ccccc} K_2SO_4 & \longrightarrow & 2K^+ & + & SO_4^{-1} \\ .2-x & & 2x & & x \\ \boxed{\cdot 01M} & & \boxed{.38M} & & \boxed{\cdot 19} \end{array}$$

$x = .2 \times \frac{95}{100} = .19$

as per given: ↑

Q. cal. concentration of 10^{-5} M $Ca_3(PO_4)_2(aq.)$ sol.

sol.

$$Ca_3(PO_4)_2 \longrightarrow 3Ca^{3+}(aq.) + 2PO_4^{3-}(aq.)$$

$$3 \times 10^{-5}\,M \qquad 2 \times 10^{-5}\,M.$$

★★ If solute reacted with each other

$\boxed{\begin{array}{l} HCl(aq.) \\ M_1, V_1 \end{array}}$ + $\boxed{\begin{array}{l} NaOH(aq.) \\ M_2, V_2 \end{array}}$

↘ ↙

Total vol of solⁿ · $V_1 + V_2$

★ conⁿ of NaCl [NaCl] = $\frac{\text{Moles of acid/base}}{\text{Total vol. of sol}^n (L)}$

Q. calculate conⁿ of NaCl

(i) If .1M of HCl (aq.) of 20ml is mixed with 10ml of .2M NaOH (aq.) sol.

(ii) If 60 ml of .1M ———— 10 ml of .2M NaOH(aq.) solⁿ.

Sol. (i) $HCl^{(aq)} + NaOH^{(aq.)} \rightarrow NaCl^{(aq.)} + H_2O\ (l)$

m.moles 2 2

— — 2

(ii) ~~4~~6 2

4 2

$[NaCl\ (aq.)] = \frac{\text{millimoles of NaCl}}{\text{volume soln (ml}}$

$= \frac{2}{60+10} = \frac{2}{70} = \frac{1}{35}$.

$[Cl^-\ (total] = \boxed{\frac{6}{70}} = \boxed{\frac{1}{35} + \frac{2}{35}} = \frac{3}{35}$.

Q. If 100 ml of .1M $Al_2(SO_4)_3$ (aq.) solution is mixed with 100 ml of .1M $AlCl_3$ (aq.) solⁿ. Calculate molarity of Al^{3+}, SO_4^{2-}, Cl^{-1}

Q. ② part —

If $d_{Al_2(SO_4)_3} = 1.30\ g/ml$

$d_{AlCl_3} = 1.40\ g/ml$

$d_{soln} = 1.32\ g/ml$

Calculate molarity of Al^{3+}

Sol. $Al_2(SO_4)_3 + Al_3Cl_3 \not\rightarrow X$

m.moles 10 10

$Al_2(SO_4)_3 \longrightarrow 2Al^{3+}(aq.) + 3(SO_4)^{2-}$

m.moles 10 20 30

$Al_3Cl_3 \longrightarrow 3Al^{3+} + 3Cl^{-1}$

10 30 30

$\therefore [Al^{3+}] = \frac{30}{200} = .15\,M$

$[SO_4^{2-}] = \frac{30}{200} = .15\,M$

$[Cl^{-1}] = \frac{30}{200} = .15\,M$

2 part :-

sol. Total mass of sol. = 130 + 140

= 270 g.

$V_{sol} = \frac{270}{1.32}$ ml

$\therefore [Al^{3+}] = \frac{30}{270} \times 1.32$

$= \boxed{.146\,M}$

Q. If 100 ml .1M $AlCl_3$(aq.) mixed with 200 ml .2 M $AlCl_3$(aq.) sol. then calculate molarity Al^{3+} Cl

$[Al^{3+}] = \frac{50}{300} = \frac{1}{6}\,M$

$[Cl^-] = \frac{150}{300} = \frac{1}{2} = .5\,M$.

Q. Calculate molality of $Al_2(SO_4)_3$, Al^{3+} and SO_4^{2-} in 1M $Al_2(SO_4)_3$ sol. (d = 2.342 g/ml)
(1 molar)

Sol. $\boxed{1L\ sol^n}$ $Al_2(SO_4)_3 \rightarrow 2Al^{3+} + 3SO_4^{2-}$

1 2 3

234.2 gm

Wt. of solvent = 2342 − 342 = 2000 gm

∴ Molality = $\frac{234.2}{}$

$= \frac{1}{2000} \times 1000 = \boxed{.5\,m}$

$m_{Al^{3+}} = \frac{2}{2} = \boxed{1m}$

$m_{SO_4^{2-}} = \frac{3}{2} = \boxed{1.5m}$

Q. Molarity? of NH_3 (d = .9 g/ml) which contain 17% NH_3 by weight.

$$M = \frac{17}{100} \times \frac{90}{17} \times \frac{1}{100} \times 1000$$

$$= \boxed{9M}$$

Q. The density of 1.35 m NaOH sol^n is 1.02 g/cm^3 calculate M.

$$\frac{1.35}{\frac{(1.35 \times 40 + 1000g)}{1.02}} \times 1000$$

$$= \frac{1.02 \times 1.35}{52 + 1000} \times 1000$$

$$= \frac{1.02 \times 1.35}{[illegible]} \times 1000$$

★ **Molarity of Pure Liquid :**

Pure Liquid :- $H_2O(l)$, $C_2H_6(l)$, $C_2H_5OH(l)$ etc.

MOLARITY : No. of moles of substance per litre.

★ Molarity of pure water :

Mass of 1l water = 1000 gm

$$\text{Moles} = \frac{1000}{18} = 55.55.$$

$$\therefore \quad M = \frac{55.55}{1}$$

★ Molarity of C_6H_6 , d = .88 g/ml

Mass of 1000 ml or 1L

$$\text{Moles} = \frac{880\cancel{0}}{78}$$

$$M = \frac{880\cancel{0}}{78} = \frac{112.9}{10} = \boxed{11.2\ M}$$

★ **Molarity of Ideal Gas :**

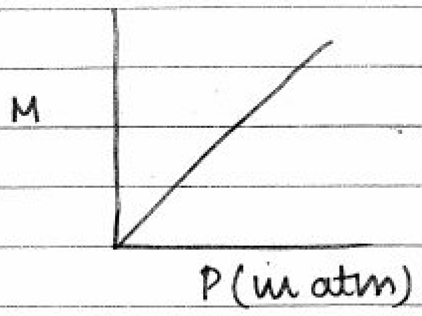
$$PV = nRT$$

$$M_{gas} = \frac{n}{V(\text{in } L)}$$

$$\therefore \quad \boxed{M_{gas} = \frac{P}{RT}}$$

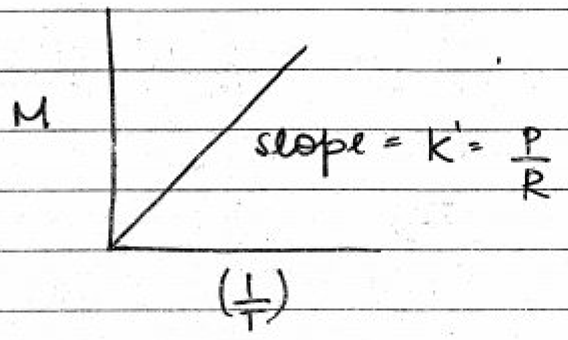
★ At constant temperature :-

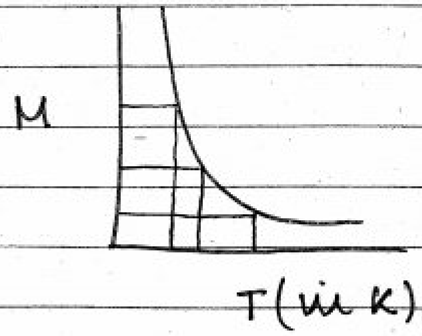
$\boxed{M = K.P}$ where $K = \frac{1}{RT}$ = constant



★ At constant Pressure :-

$\boxed{M = \frac{K'}{T}}$ where $K' = \frac{P}{R}$ = constant





∴ their Product = Constant

★ **Molarity of Gaseous Mixture** :

Mixture of gases A and B.

→ Molarity of mixture in terms of A

$$= \boxed{\frac{n_A}{V_{mix}(l)}}$$

→ Mixture's molarity in terms of B

$$= \boxed{\frac{n_B}{V_{mix}(l)}}$$

★ Molarity of **Air** :

★ If air contains only N_2 and O_2

(Mostly in terms of less conc^n) $\boxed{O_2}$

(M) in terms of O_2 = $\frac{n_{O_2}}{V_{air}(l)}$

Q. Calculate M_{air}. Given :- average molar mass 28.8 g. at 1 atm and 273 K

Sol.

$$28(x) + 32(1-x) = 28.8$$

$$-4 \cancel{+12}x = \cancel{12.8} \ -3.2$$

$$x = \frac{12.8}{12}$$

$$x = .8$$

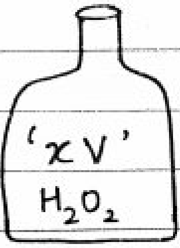
M_{air} in terms of O_2 $= \frac{.2}{\cancel{2} \times 22.4} \times 1$

$$= \boxed{\frac{1}{112}} \text{ M.}$$

Imp.

★ VOLUME STRENGTH OF H_2O_2

$$H_2O_2(aq) \longrightarrow H_2O(l) + \frac{1}{2}O_2(g)$$



xV :-

1 volume of H_2O_2 on decomposition gives x volume of O_2 at 1 atm & 273 K.

OR 1 litre of H_2O_2 given x l of O_2

1 litre of H_2O_2 gives $\frac{x}{22.4}$ moles of O_2

∴ 1 litre of H_2O_2 gives $\frac{x}{11.2}$ moles of H_2O_2

$$\therefore \text{Molarity of } H_2O_2 \text{ sol}^n = \frac{n_{H_2O_2}}{\text{Volume (L)}}$$

$$= \frac{\frac{x}{11.2}}{1}$$

$$= \boxed{x/11.2}$$

V. Imp

★★★

$$\boxed{\text{Molarity of } H_2O_2 \text{ sol}^n = \frac{\text{Volume strength of } H_2O_2}{11.2}}$$

★ Calculate the following concentration terms for "16.8 V" H_2O_2 solⁿ.

$d_{soln} = .951$ g/ml

(i) Molarity $= \frac{16.8}{11.2} = 1.5$ M

(ii) Molality =

Wt. of solvent $= -(34 \times 1.5) + 951$

$= 900$ gm

$\therefore \frac{1.5}{900} \times 1000 = \boxed{1.666 \text{ m}} = \boxed{1.67 \text{ m}}$

(iii) $\%\left(\frac{W}{V}\right) = \frac{51}{1000} \times 100 = \boxed{5.1\%}$

(iv) $\%\left(\frac{W}{W}\right) = \frac{51}{951} \times 100 = \boxed{5.37\%}$

(v) ppm of solute $= \boxed{5.37 \times 10^4}$

(vi) Mole fraction of solute =

$\frac{1.5}{1.5 + \frac{900}{18}} = \boxed{\frac{1.5}{51.5}}$

★ Temperature Dependence Of Concentration Terms.

★ Mass / Moles of substance not depend on temperature

★ Volume of substance (l/g) depends on temperature.

Temperature Dependent :-

$\%\left(\frac{W}{V}\right)$, Molarity

Temperature Independent :-

$\%\left(\frac{W}{W}\right)$, Molality, Vapour density

EUDIOMETRY

for ideal gases

EUDIOMETRY TUBE:

graduated tube (in mm) with Platinum terminal. They produce sparks.

- electric sparks used.

$$PV = nRT$$

At const temp. and pressure :- $\boxed{V \propto n}$

At const. temp and volume :- $\boxed{P \propto n}$

★★ In gaseous reaction, volume may increase, may decrease or remains constant. (large amount)

—×—

$$2NH_3 \rightarrow N_2 + 3H_2$$

V ml $\quad \frac{V}{2}$ ml $\quad \frac{3}{2}V$ ml $\qquad$ V ml Expansion

★ Expansion in volume = Volume of (Products − Reactants)

★ Contraction in volume = Volume of (Reactants − Products)

$$CO + \frac{1}{2}O_2 \rightarrow CO_2$$

$V \quad V/2 \quad V$ $\qquad$ $V/2$ ml contraction.

$$\underset{V}{H_2} + \underset{V}{Cl_2} \longrightarrow \underset{2V}{2HCl}$$

{ No change in volume }

★ **Remember :-**

★ Untill not speified N_2 is considered as inert gas.

★ Until not specified H_2O formed should be considered in liquid state. ($\because$ H_2O is liquid in room temp in tube)

★ Volume of solids and liquids are negligible as compared to gases ($H_2O(l)$)

★ No other gases are liquid.

★★

$$C_xH_y + \left(x+\frac{y}{4}\right)O_2 \rightarrow xCO_2 + \boxed{\frac{y}{2}H_2O}$$

$$a\,ml \qquad a\left(x+\frac{y}{4}\right)ml \qquad ax\,ml \qquad \cancel{a\left(\frac{y}{2}\right)ml}$$

→ Negligible volume (l)

Contraction :- $\left\{a + a\left(x+\frac{y}{2}\right)\right\} - ax - \cancel{\frac{ay}{2}}$

$$= \boxed{a + \frac{y}{2}}$$

★★ $$C_xH_yO_z + \left(x+\frac{y}{4}-\frac{z}{2}\right)O_2 \rightarrow xCO_2 + \frac{y}{2}H_2O$$

$$\begin{array}{cccc} 10\,ml & 10\,ml\left(x+\frac{y}{4}-\frac{z}{2}\right) & - & \\ - & - & 10x\,ml & - \end{array}$$

—

!!!

★★★

Gas/gases absorbed	Substance used.
★ $CO_2(g)$, $SO_2(g)$, $Cl_2(g)$	KOH
★ $CO(g)$	ammonical Cu_2Cl_2
★ $O_2(g)$	alkaline pyragallol solution
★ $O_3(g)$	Tarpentine oil
★ $HCl(g)$, $NH_3(g)$	H_2O

Note ★ anhydrous $CaCl_2/CuSO_4$ is used to absorb moisture.

★ In 50 ml mixture of CO_2 and CO

$$50\,ml \longrightarrow \boxed{\begin{array}{c}KOH\\ solution\end{array}} \longrightarrow \begin{array}{c}20\,ml\\ CO_2\end{array}$$

Contraction vol = 30 ml

(∵ 30 ml of CO present)

★ 15 ml O_3 and O_2

$$15\text{ ml} \longrightarrow \boxed{\begin{array}{c}\text{terpentine}\\ \text{oil}\end{array}} \longrightarrow \begin{array}{c}10\text{ ml}\\ O_2(g)\end{array}$$

contraction in vol = 5 ml

—x—

Q. 10 ml of CO is mixed with 100 ml O_2 and CO_2 formed. Calculate contraction in volume

sol.

$$CO + \tfrac{1}{2}O_2 \rightarrow CO_2$$

10ml 5ml →

— — 10ml.

(i) Contraction — ~~10ml~~ $\boxed{5\,ml}$

(ii) Yield — $\frac{75}{100} \times 10 = 7.5\,ml$.

Contraction — $(20 - (7.5 + 2.5 + 7.5)$

$= \boxed{2.5\,ml}$

$= 3.75\,ml$.

Q. 10 ml of gaseous hydrocarbon on complete combustion gave 30 ml of CO_2 and it was found that it required 40 ml of O_2. Calculate molecular formula of hydrocarbon.

sol.

$$C_xH_y + 4O_2 \rightarrow 3CO_2 + 2H_2O$$

10ml 40ml 30ml

Ans $x = 3$

$y = 4$

$\therefore \boxed{C_3H_4}$

Q. A 20 ml sample of hydrocarbon was exploded with excess amount of O_2. After explosion there was a volume contraction of 60 ml when resultant mixture tested with KOH another contraction of 80ml was observed. Calculate M.F.

Sol.

$$C_xH_y + O_2 \longrightarrow xCO_2 + \frac{y}{2}H_2O$$

$(x+\frac{y}{2})$

20ml $\quad (20)ml\,(x+\frac{y}{2}) \quad$ 20x ml $\quad \frac{20y}{2}$ ml

$$\cancel{\frac{20x}{x}} + \frac{20y}{2} = 20 + 20(x+\frac{y}{2}) - 60$$

??

$20x = 80$ $\quad$ ($\because$ CO_2 when react with KOH is evaporated).

$\boxed{x = 4}$

$\therefore \boxed{y = 8}$

Q. A 10L mixture of C_2H_6 and C_3H_8 at 1 atm and 273 K on complete combustion gave 24 L of CO_2. Calculate $V_{C_2H_6}$ and $V_{C_3H_8}$

Sol.

$$C_2H_6 + \frac{7}{2}O_2 \longrightarrow 2CO_2 + 3H_2O$$

(x) $\quad$ $(2x)$

$$C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$$

$(10-x)$ $\quad$ $3(10-x)$

$2x + 30 - 3x = 24$

$\boxed{x = 6\ L}$

Q. 40 ml of mixture of $H_2O(g)$, $CH_4(g)$ & $N_2(g)$ was exploded with 10ml of $O_2(g)$. after explosion and on cooling the gases occupied 36.5 ml. After treatment with KOH volume reduced by 3ml and again on treatment alk. pyrgallol sol., vol. further decreased by 1.5ml. Determine comp.

sol.

V_{CO_2} = 3 ~~33.5~~ ml $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

∴ V_{CH_4} = 3 ~~33.5~~ ml   6ml   3

V_{H} V_{O_2} used =

$H_2O + \frac{1}{2}O_2 \rightarrow H_2O$

8ml   4ml

$n_{H_2O} = \frac{8}{22.4}$

$\therefore V_{H_2O} = \frac{8}{22.4} \times 18 = \frac{144}{22.4} = 6.$

Ans H_2 : CH_4 : N_2

8ml : 3ml : 25.5ml

5 : 3ml : 32ml.

Ans →

Q. A sample of organic compound. (M = 110g/mol) having 10ml volume contains C,H, O and S was mixed with excess O_2 having V = 100ml contraction – 10ml on passing the gaseous mixture through KOH the vol. contracted to 70ml. If avg. molar mass of absorbed gas by KOH is 50.667. find M.f.

Sol. Volume of SO_2 and CO_2 → 70ml.

$$C_xH_yO_zS_k + \overset{(x+\frac{y}{4}+k-\frac{z}{2})}{O_2} \longrightarrow xCO_2 + \frac{y}{2}H_2O + kSO_2$$

$$10 \qquad \left(x+\frac{y}{4}+k\overset{-z/2}{\wedge}\right)10$$

$$10 \qquad 100 \qquad 20 \qquad 10$$

$$x+\frac{y}{4}-\frac{z}{2}+k = 10 \qquad \therefore \frac{y}{2} = 7$$

$$\boxed{y = 2z} \qquad \boxed{y = 14}$$

$10x + 10k = 30$

$(x+k) = 3$

$$44x + 64k = 50.667(x+k)$$

$$13.33k = 16.33x$$

$$13.33(3-x) = 16.33x$$

$$39.99 = 30x$$

$$x = 1.333 = 1.3\cancel{4} = \frac{\cancel{4}}{3}$$

$x:k = 2:1$ $\qquad k = 1.66 = \frac{5}{3}$

$\therefore$ vol. ratio $= \boxed{2:1}$

$$\therefore z = \left(\frac{14}{4}\cancel{\times 2} - 20\right)2$$

$$= \cancel{6}4 \qquad 7-$$

★ Experimental determination of molar mass of substance.

★ Dulong's & Petit's law:

→ Applicable for metals only.

→ Used to find approx. mass (molar) of metals

Heat : $q = M\,\boxed{C}\,\Delta T$

$M \rightarrow$ Mass of Metal

$C \rightarrow$ Specific heat. (cal/g°C or cal/g K)

★ **Law**

$$\boxed{\text{gram atomic mass of metal} \times \text{specific heat (cal/g°C)} = 6.4}$$

Q. 7.38 g of a sample of metal oxide (M_2O_x) is quantitively reduced to 6.84 g of pure metal. Specific heat of metal is .0332 cal/g°C. Calculate valency and acc. GAM of metal.

Sol. $\boxed{\text{POAC : } M_2}$

$$2\left(\frac{7.38}{2M+16x}\right) = \frac{6.84}{M}$$

But $M \times .0332 = 6.4$

$$M = \frac{6.4}{.0332}$$

$\therefore \quad 14.76M = 13.68M + 16 \times 6.84x$

$\therefore \quad x = 1.9$ (approx.)

$\therefore \boxed{\text{Valency} = 2}$

$\therefore$ Accurate GAM =

$$\frac{7.38 \times 2}{2M_1 + 16 \times 2} = \frac{6.84}{M_1}$$

$$\boxed{M = 202.66 \text{ g/mol}}$$

VICTOR MEYER METHOD

(used to foiund molar mass of volatile substance)

- ★ vapour of volatile substance is produced by heating
- ★ Formed vapour, displace air in vessel.
- ★ Displaced air collected over liquid surface $\{H_2O(l), Hg(l)\}$

★ At room temperature :
$Hg(l)$ is non volatile at room temperature
$H_2O(l)$ is volatile substance.

At equilibrium :

$$\underset{(H_2O)(l)}{\text{Rate of evaporation}} : \underset{(H_2O(l)}{\text{Rate of condensation}}$$

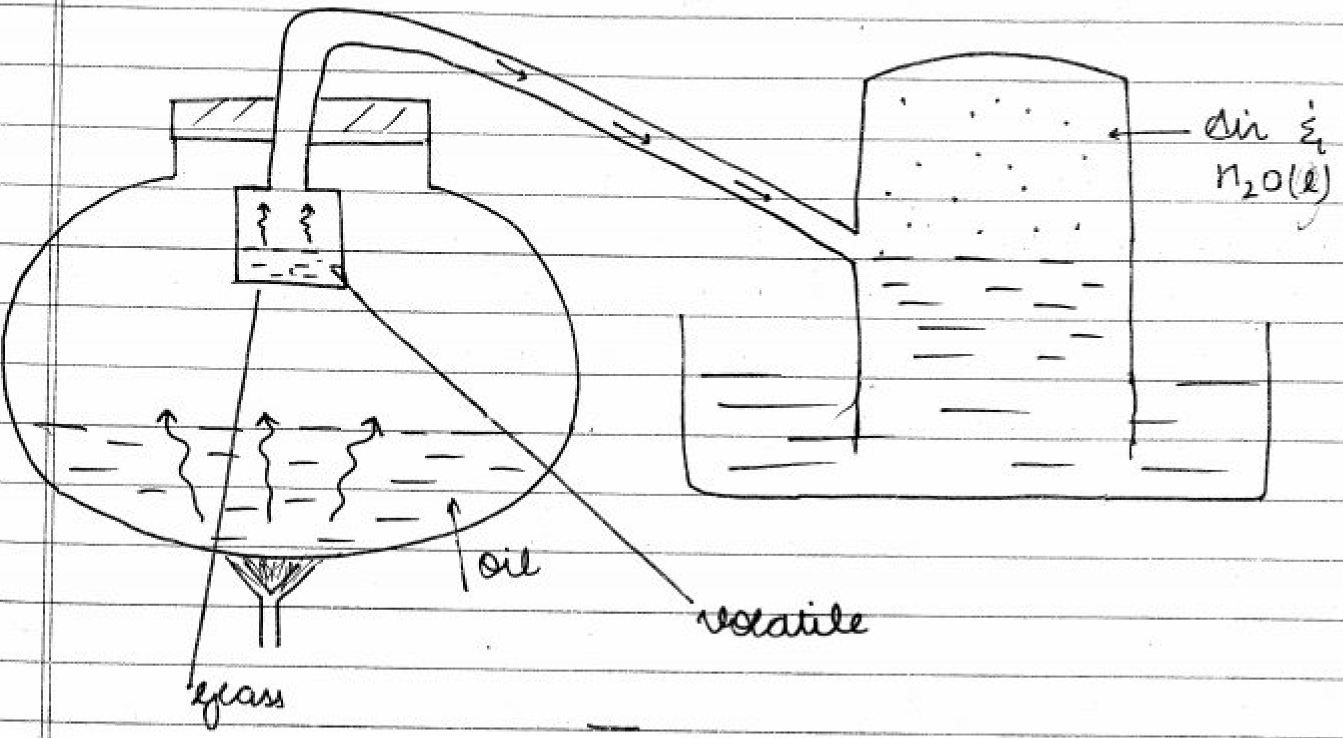
$$n_{H_2O(l)} = \text{constant}$$

$$n_{H_2O(g)} = \text{constant}$$

VAPOUR PRESSURE

The pressure exerted by the vapours on its liquid surface at equilibrium is called vapour pressure. of ~~or~~ liquid or aquones tension of water.

$\boxed{*}$ Hg makes no vapour pressure.



$P_{\text{wet air}} = P_{\text{dry air}} + \text{V. Pressure of } H_2O(l)$

$P_{\text{wet air}} =$

$$\boxed{P_{\text{dry air}} = P_{\text{wet air}} - \text{V. Pressure of } H_2O(l)}$$

$$P_{\cancel{\text{wet}}\,\text{air}} V_{\text{air}} = n_{\text{air}} RT$$

$$\boxed{n_{\text{air}} = n_{\text{substance}}}$$

$$\boxed{P_{\text{dry air}} V_{\text{air}} = \frac{W_{\text{substance}}}{M_{\text{substance}}} \times RT}$$

Wt. of displaced substance.

Molar Mass of Volatile substance

Page

- If air is cooled over Hg surface

$$\boxed{P_{air} = P_{dry\ air}}$$

Q. In a victor meyer method .2g of a volatile substance displaced 63.3 ml of air over water at 20°C and 740 mm Hg. Calculate the molar mass of substance.

given : V.P. of H_2O at 20°C = 18 mm

Sol.

$$P_{dry\ air} = P_{wet\ air} - V.P.$$

$$= 720 mm$$

$$\frac{10^{-3} \times 722 \times 63.3}{760} = \frac{.2}{M} \times .0821 \times 293$$

$$M = \underline{80}$$

★ SILVER SALT METHOD

{used to find molar mass of organic acids contain COOH group}.

Ex: Monoprotic acid - CH_3COOH
{Basicity : 1}

Step1 : Acid is reacted with aqueous solution of $AgNO_3$

$$RCOO\underline{H} + 2\underline{Ag}NO_3 \longrightarrow 2RCOOAg_2 + 2\underline{H}NO_3$$

$\boxed{\text{silver salt}}$

Step 2 $RCOOHOAg \xrightarrow{\Delta} Ag(s)$
x gm y gm

$\boxed{\text{POAC : Ag}}$

$$\frac{x}{M_{ACOOAg}} = \left(\frac{y}{108}\right) \times 1$$

★ Molar Mass of RCOOH = Molar mass of RCOOAg − Molar mass of Ag^+/Ag + Molar mass of H^+/H.

★ H_2A → Diprotic acid (Basicity = 2)

In diprotic acid :-

Molar mass of H_2A = Molar mass of Ag_2A −
2× Molar mass of Ag^+/Ag
+ 2× Molar mass of H^+/H

Tiprotic acid :-

$$H_3A + 3AgNO_3 \longrightarrow Ag_3A + 3HNO_3$$

$$Ag_3A \longrightarrow \cancel{3}Ag(s)$$

$\boxed{\text{POAC : Ag}}$

$$\left(\frac{x}{M_{Ag_3A}}\right) \times 3 = \left(\frac{y}{108}\right) \times 1$$

★ Molar mass of H_3A = Molar mass of Ag_3A −
3× Molar mass of Ag^+/Ag
$\underbrace{+ 3\times \text{Molar mass}}_{3(107)}$ of H^+/H

Q. .607g of silver salt of tribasic organic acid was quantitively reduced to .37g of pure silver.

Calculate the molar mass of acid.

Sol.

$$\frac{.607}{x} \times 3 = \frac{.37}{108}$$

$$x = 531.53$$

$\therefore$ Ans $531.53 - 3 \times 107$

$= \underline{210.53}$

ATOMIC STRUCTURE

-ve

★ Cathode Rays : J.J. Thomson.

Discovery of electrons

Cathode rays made up of stream of negatively charged particles called electrons.

+ve

★ Anode rays : Goldstein
(Canal rays)

Discovery of protons.

★ Discovery of Neutrons : J. Chadwick.

Found by artificial radioactivity

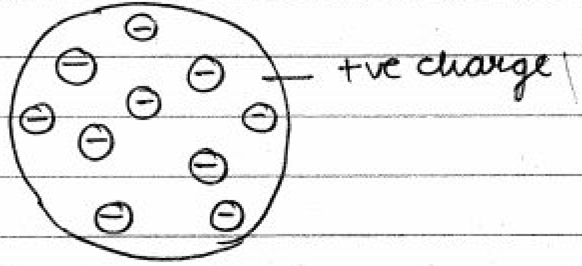
★ Mass of proton : 1.6726×10^{-24} g
★ Mass of neutron : 1.675×10^{-24} g
★ Mass of electron : 9.1×10^{-28} g

Millikan experiment :

★ charge of electron = 1.67×10^{-19} C

★ <u>Thomson's atomic model</u> :-
(watermelon model)
(Plum pudding model)

★ Overall atom is electrical neutral
★ +ve charge uniformly dispersed over entire sphere.
★ electrons are embedded.



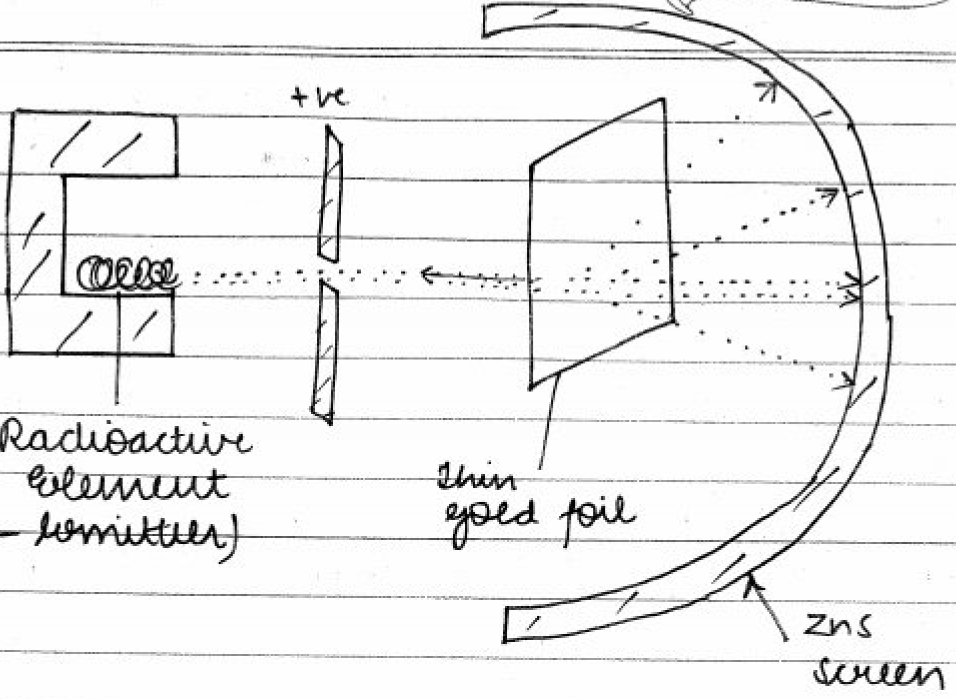
mass and +ve
charge uniformly dispersed.

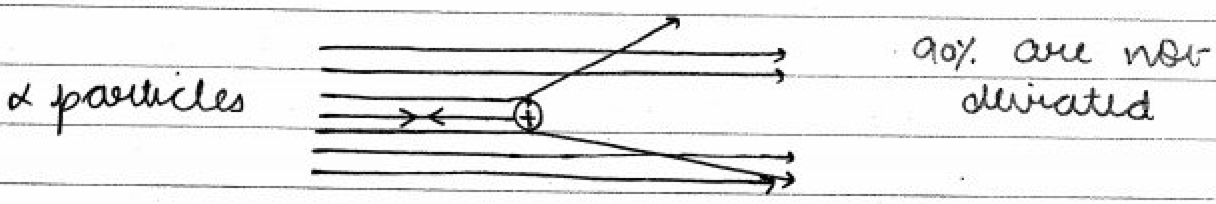
★ <u>Rutherford α-particle scattering</u> :

★ α particles are used.
★ Radioactive element used.
(Polonium)
Radium

★ α particle (positive charge)
:- ⇒ $^{4}_{2}He^{2+}$ (double ionized He nuclei)

Rutherford Atomic Model





Observations:

★ (1) Most of the α particles are not deviated, passed through gold foil without any deflection.

★ (2) Few α-particles deviated with small angles.

★ (3) Very few particles ~~very~~ with large angles.

★ (4) Very few particles rebounded (angle of deviation = 180°)

Conclusion:

★ Almost volume of atom is chargeless/empty

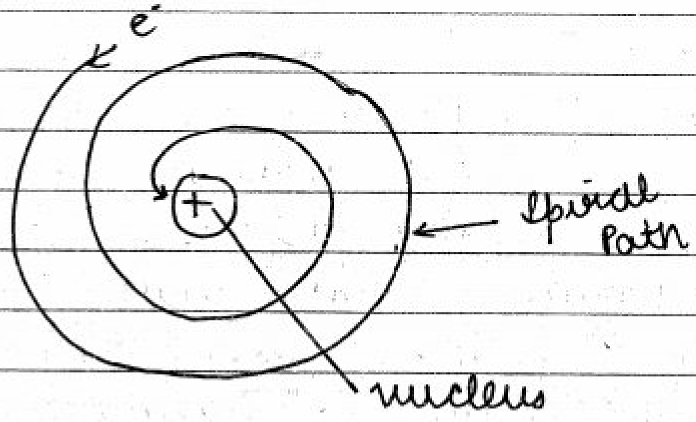
★ Positive charge heavy dense core present in atom is called nucleus.

Rutherford Atomic model.

- Overall atom is electrical neutral
(Total +ve charge = Total -ve charge)
- Almost volume of atom is empty.
- Heavy dense +ve charge nucleus charge present in centre of atom.
- electrons revolve outside the nucleus with very high speed and at different distances.

- Nucleus and electrons are held together due to electrostatic force of attraction.

- Drawback:



- Acc. to Electrodynamic's theory when electron revolve outside +ve nucleus it should emit radiations
Due to radiations, energy is lost.
So electrons must follow spiral path and fall in nucleus.

- Was unable to explain discontinuous spectrum.

** Mass of atom = Mass of $(n + P + e^-)$
$\simeq$ Mass of $(n + p)$

** Radius of Nucleus

$$\boxed{r = r_0 A^{1/3}}$$

$$\boxed{\text{where } r_0 = 1.33 \times 10^{-13}\ \text{cm}}$$

A = atomic mass no.

$$\boxed{\text{Radius of nucleus} = \text{Radius of atom} \times 10^{-5}}$$

* COULOMBIC FORCE OF ATTRACTION/ REPULSION.

→ Repulsion between same charge particles.
→ Attraction between opp. charge particle

$F \propto q_1 q_2$

$F \propto \frac{1}{r^2}$

K = constant

$K = 10^9 \times 9\ Nm^2/C^{-2}$

$F = \frac{K q_1 q_2}{r^2}$ where $\boxed{\begin{matrix} q_1 = Z_1 . e \\ q_2 = Z_2 . e \end{matrix}}$ $= 10^{18} \times 9$ dyne cm^2 C^{-2}

$= 1$ dyne cm^{-2} $(e.s.u.)^{-}$

If $\overset{--/}{++}$, force of repulsion.

If $\underset{+-}{-+/}$ force of attraction.

1 ~~e.s.u~~ C = 3×10^9 ~~dy~~ e.su

Energy <
- Work (W= f×d)
- Heat (m.c.Δt)

Total energy of particles / system :

Total energy = Kinetic energy + Potential energy.

★ K.E. of particles :- due to motion.

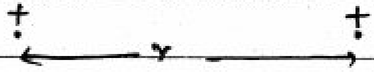
$\boxed{K.E. = \frac{1}{2}mv^2}$ m = mass
v = velocity.

★ P.E. of particle :- energy possessed by an object by virtue of its position.

$\boxed{\text{Potential energy} = \frac{Kq_1q_2}{r}}$ $\boxed{F \times d}$

$\boxed{\text{Change in P.E.} = \text{Work done}}$

Repulsive forces.



Imp. ★ distance between charge particle increases ↑
P.E ↓ decreases

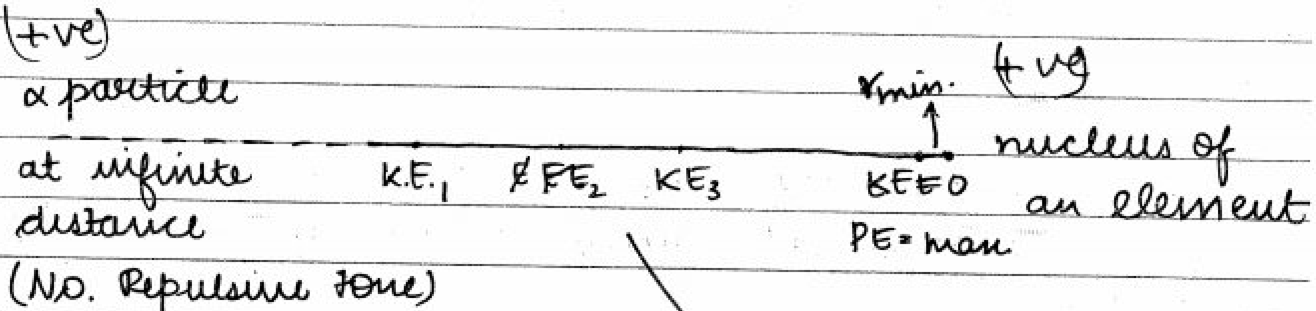
★ ________________ decrease
PE ↑ increases

Attractive force

$$\boxed{P.E. = -\frac{Kq_1q_2}{r}}$$

- ★ distance r increased, PE increased ↑
- ★ distance r decreased, PE decreased ↓

★ CALCULATION OF DISTANCE OF CLOSEST APPROACH



$(K.E.)$ max. of α particle $= \frac{1}{2}mv^2$.

$KE_1 > KE_2 > KE_3$

M = Mass of α particle
V = Velocity of α particle

K.E.
$$\left\{\begin{aligned} K.E._{(max)} &= P.E._{(max)} \\ \frac{1}{2}mv^2 &= \frac{Kq_1q_2}{r_{min}} \end{aligned}\right\}$$

Mass of α-particle — 4 a.m.u.

★ $\boxed{r_{min.} = \frac{2Kq_1q_2}{mV^2}}$

★ $r_{min.}$ = Minimum distance (closest) between particles

∴ $\boxed{r_{minimum} \simeq r_{nucleus}}$ — radius of nucleus of the element.

★ We call it "head-on collision" but it actually do not collide.

Diameter of atom = 10^{-8} cm
Diameter of nucleus = 10^{-13} cm

} Roughly

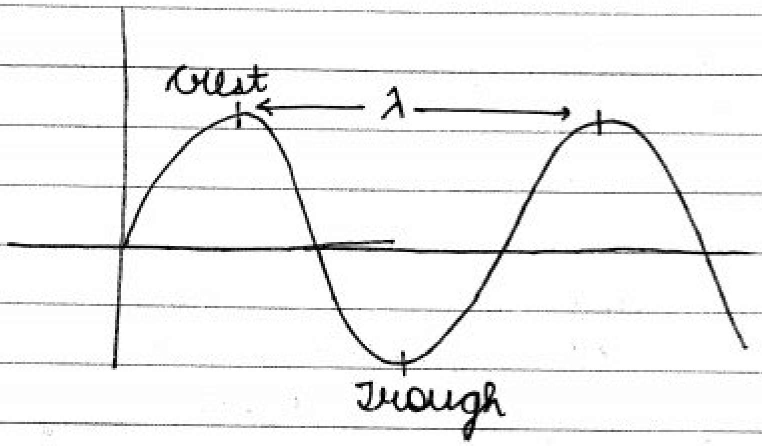
Volume — Vol. nucleus : Vol. atom

$\boxed{1 : 10^{-15}}$

★ Volume of nucleus is very small as compare to atom.

★ **WAVE** :

Mode of energy transmission from one place to another without displacement of matter.



★ **Wavelength (λ)** :-

distance between two successive peak

Unit :- Metre

$1\ \mathring{A} = 10^{-10}\ m$

$1\ nm = 10^{-9}\ m$

$1\ Pm = 10^{-12}\ m$

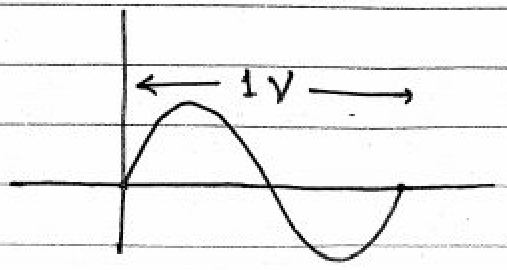
★ **Wave number ($\bar{\nu}$)** :

$\boxed{nu}$ $\quad \boxed{\bar{\nu} = \frac{1}{\lambda}} \quad \boxed{m^{-1} = \text{Unit}}$

★ **Frequency (ν)** : $\boxed{\text{Hertz or } s^{-}}$

No. of cycles per second is called frequency.

No. of waves passed through a point



* **Speed of wave :**

$\boxed{C = \lambda \nu}$ ms^{-1}

* **amplitude :**

Maximum distance from midpoint of one crest or trough.

* **Time period :**

Time required for completion of one oscillation.

$\boxed{T = \frac{1}{\nu}}$ $\boxed{sec}$

* MAXWELL THEORY OF ELECTROMAGNETIC RADIATION :

* Electromagnetic radiations are emission and transmission of energy in form of wave.

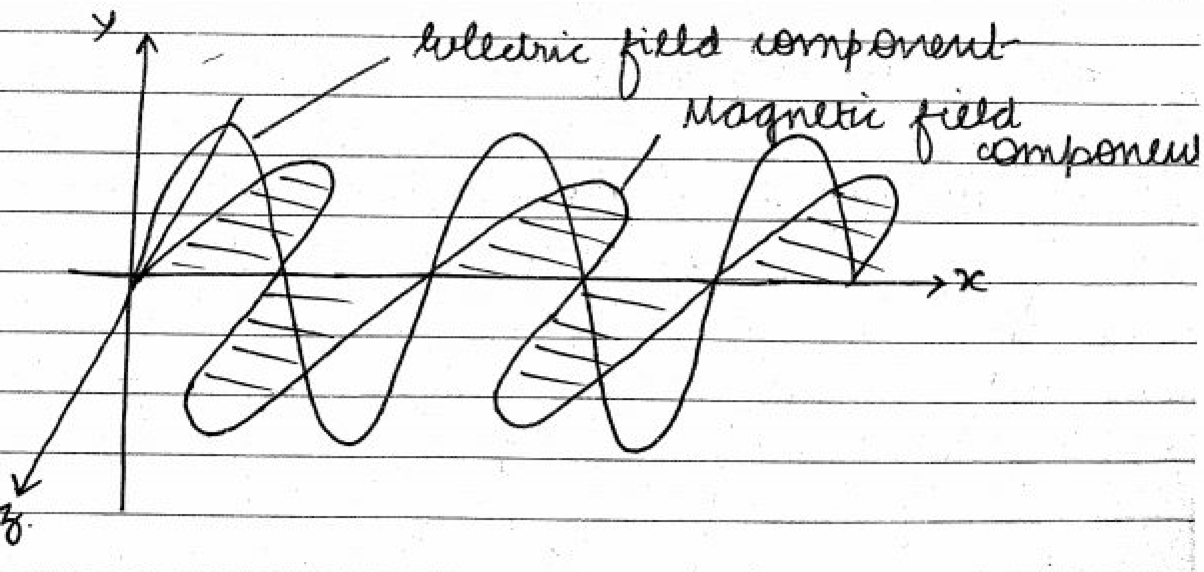
→ No matter displaced.

* Electromagnetic radiation has two component :

1) Electric field component
2) Magnetic field component.

* Both component have same frequency, speed and wavelength

* But they travel or propogate in perpendicular planes.



★ Light is type of electromagnetic radiation.

★ Speed of electromagnetic radiation in vacuum is 3×10^8 m/s

★★ **Maxwell theory was unable to explain :**

(1) Planck's theory.
(2) Photoelectric effect
(3) Black body radiations.

Electromagnetic Radiations

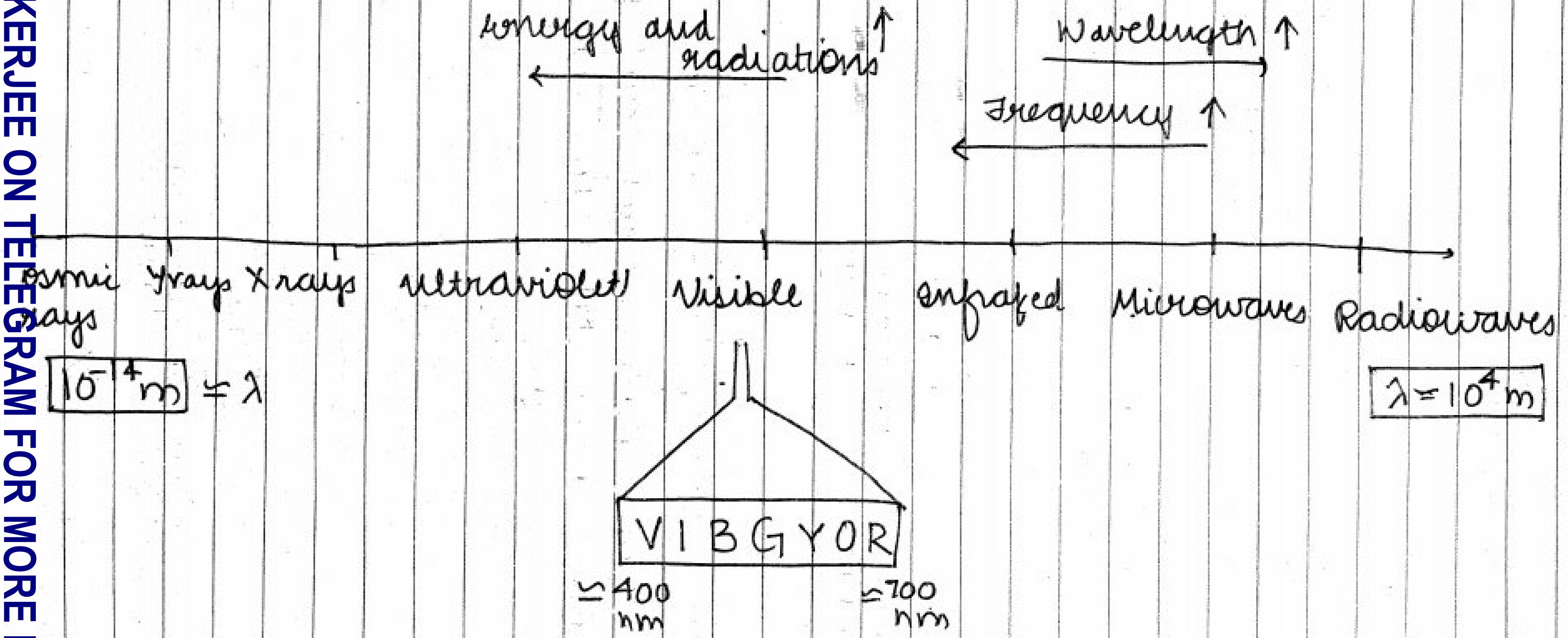
★ $E = h\nu$

Imp → Energy $\propto$ Frequency.
→ Energy $\propto \frac{1}{\text{Wavelength}}$

★ Classical Theory

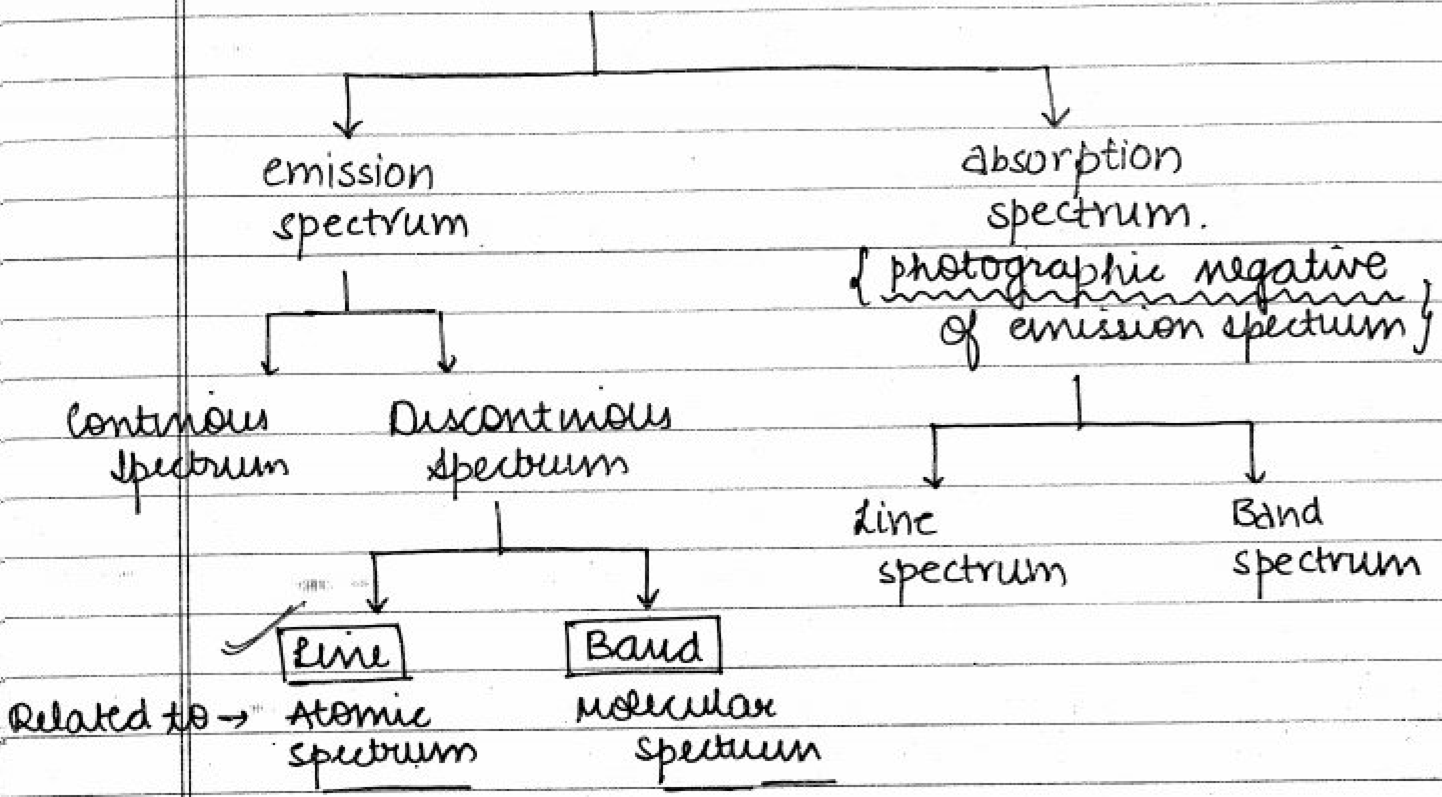
Energy of electromagnetic radiation is dependent on intensity.

Electromagnetic Radiations



SPECTRUM

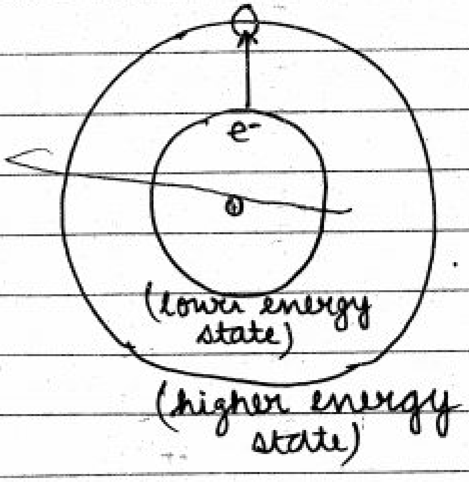
★ Types Of spectrum :



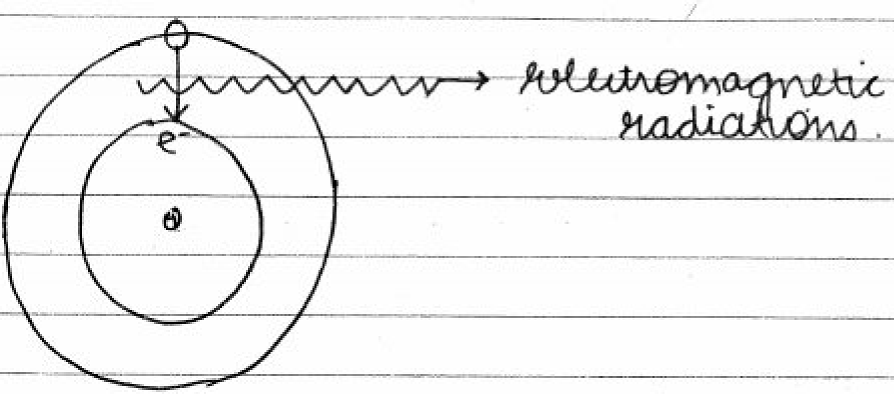
- emission spectrum
  - Continuous Spectrum
  - Discontinuous Spectrum
    - Line — Related to → Atomic spectrum
    - Band — Molecular Spectrum
- absorption spectrum. {photographic negative of emission spectrum}
  - Line spectrum
  - Band spectrum

★ DURING HEATING of substance :

Electrons absorb energy and promoted from lower state to higher energy state

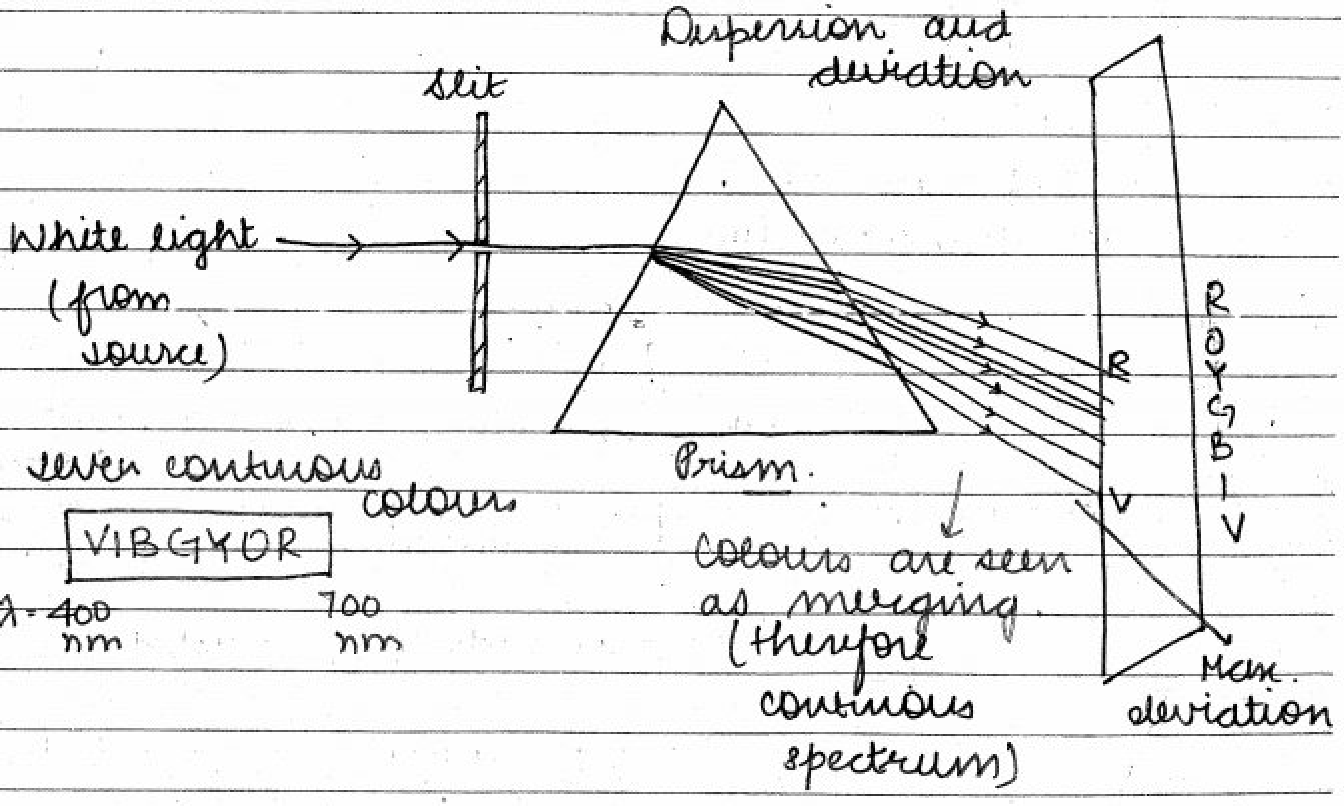


- **DURING COOLING** of substance
energy released by electrons and transferred from higher state to lower state (energy)



★ EMISSION : ——

- **Continous Emission Spectrum** :-



seven continuous colours

VIBGYOR

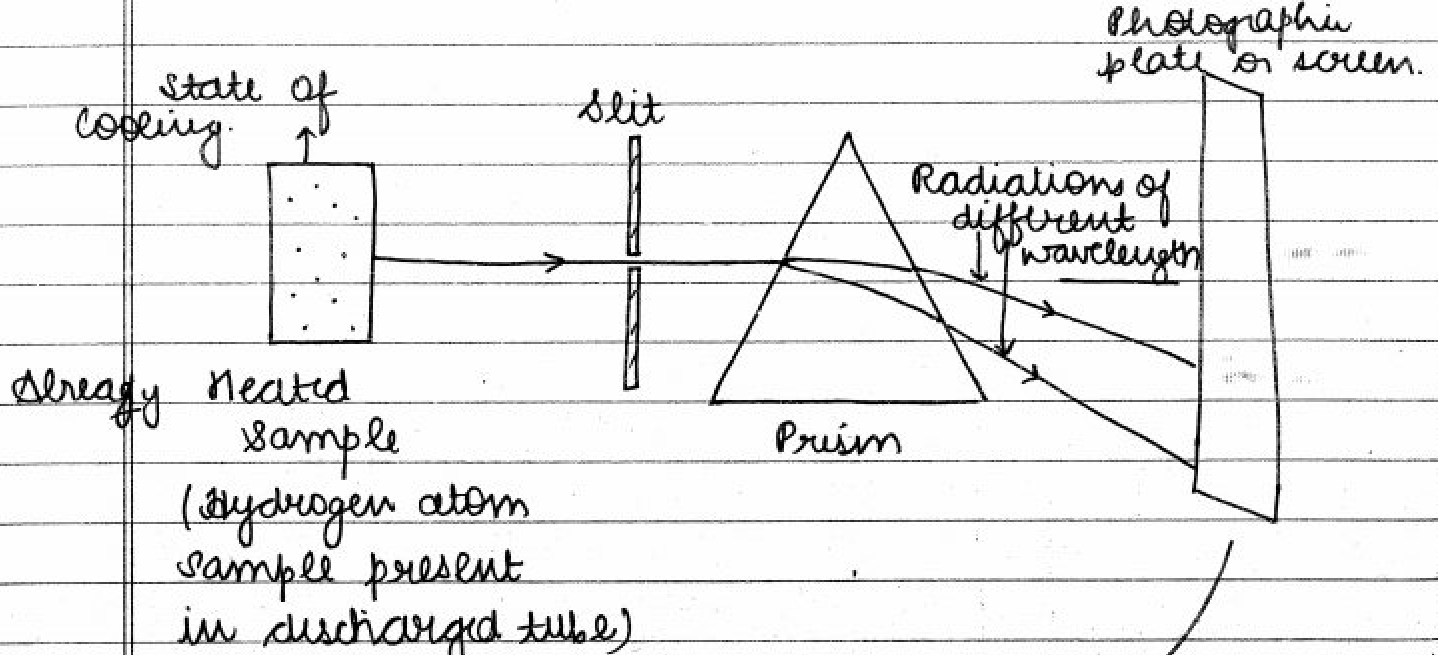
$\lambda = 400$ nm 700 nm

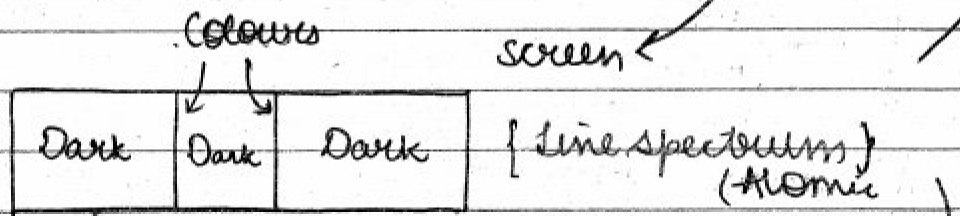
Imp

★ $\boxed{\text{Angle of deviation} \propto \text{Frequency of Radiation} \propto \frac{1}{\lambda}}$

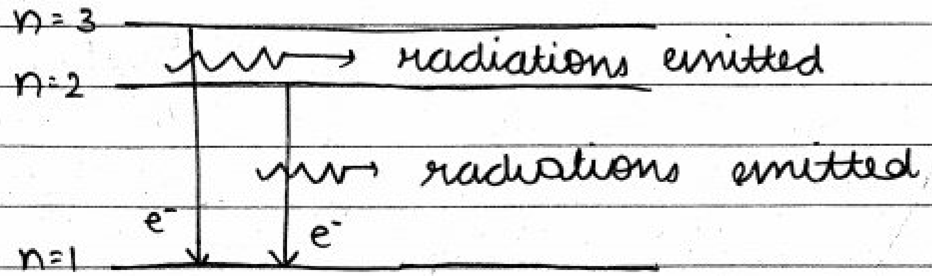
- ★ when white light from source is passed through prism, dispersion and deviation of radiations takes place and impression produced on screen.

★ DISCONTINOUS EMISSION SPECTRUM.



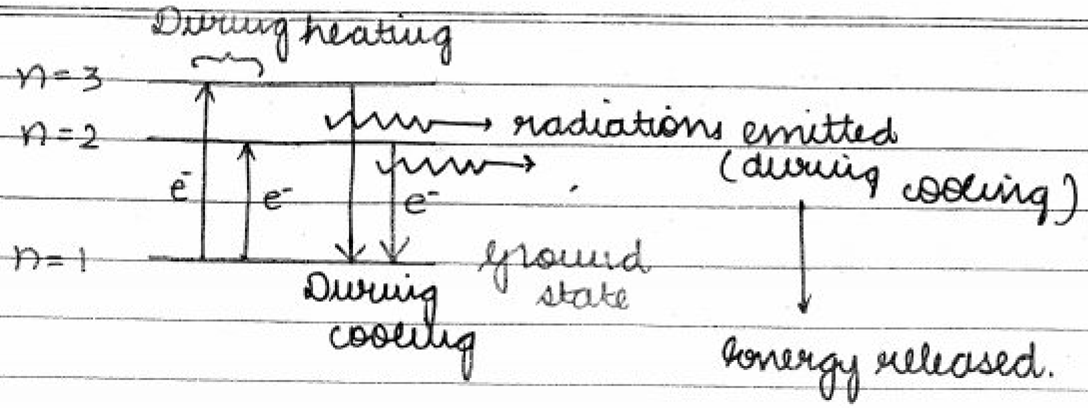


★ Each line in spectrum represent individual e^- transition.



{ energy released by e^- in terms of radiations and transferred from higher to lower energy levels }

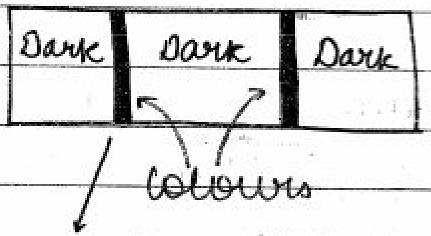
*



* If heated sample contains atoms then observed spectrum in line spectrum.

* BAND SPECTRUM (Molecular Spectrum)

(Multielectronic transition at different energy levels)

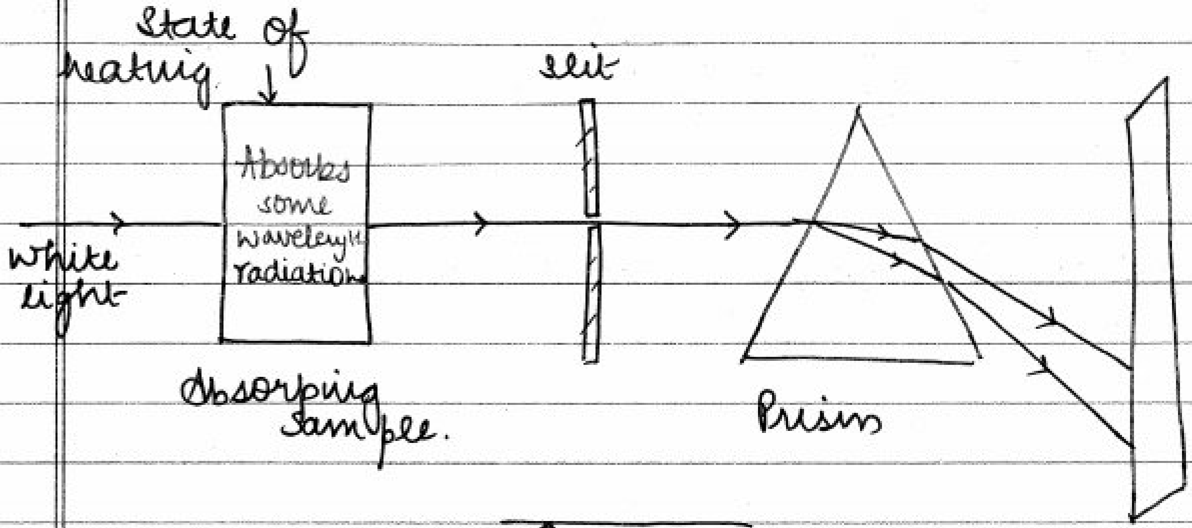


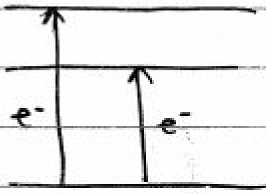
Lines are very close and fine (appear to be bands)

* If heated sample contains molecule then observed spectrum is called band spectrum.

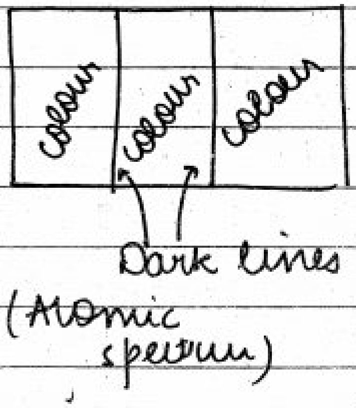
∵ (line spectrum is seen by atoms of elements only)

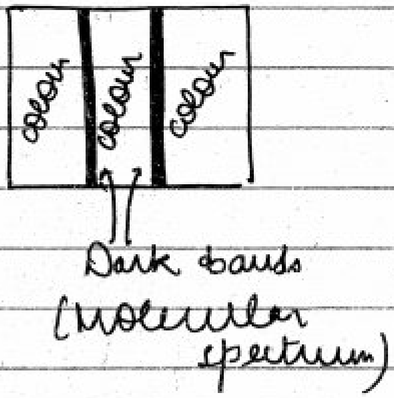
Absorption Spectrum :-





Radiations absorbed by electrons.





* Every element has unique line spectrum Different elements have different line spectrum and their spectrums are called "fingerprints of elements."

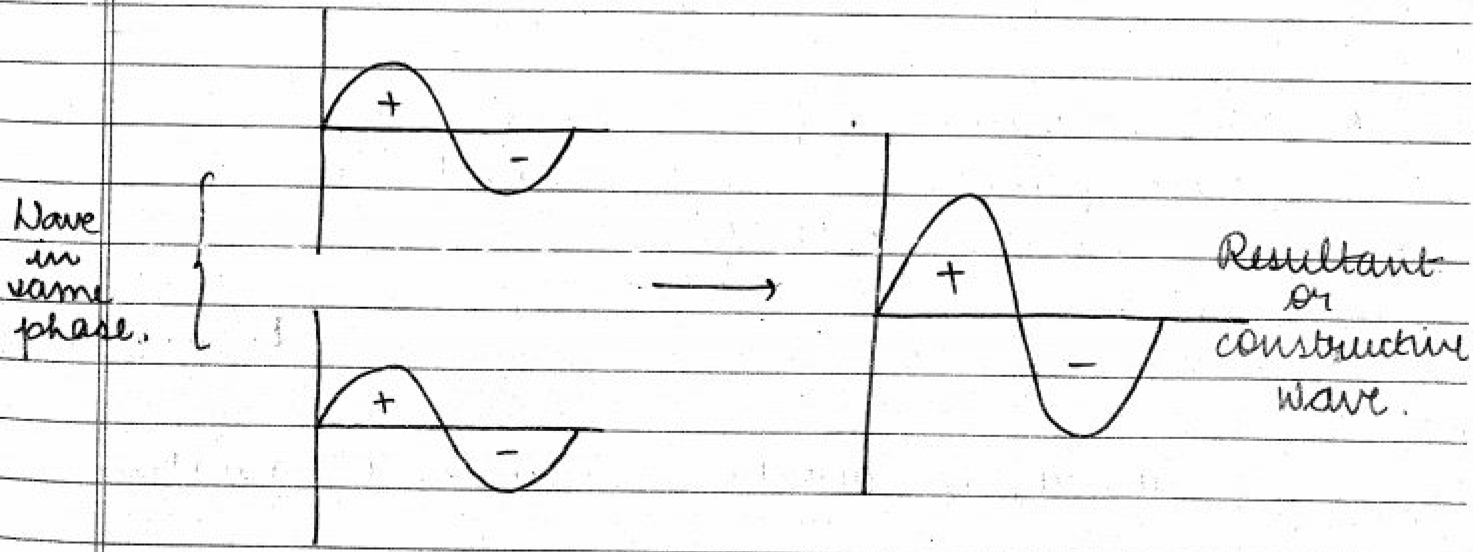
★ collection of radiations of different wavelenght on photographic plate is called spectrum.

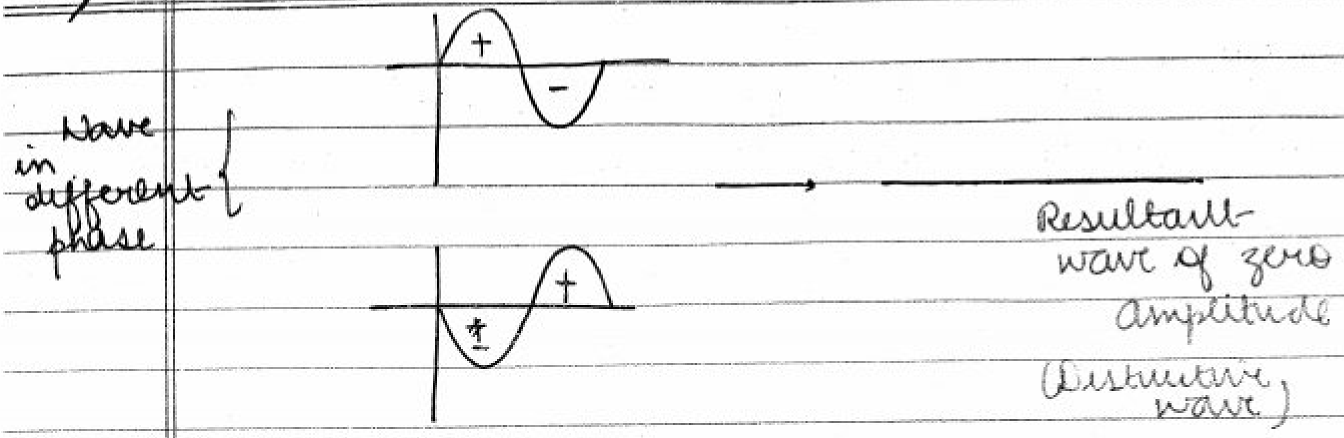
★ Diffractions & Interference of Waves:

(addition and subtraction of waves)

Spreading of wave, when it is passed through a small hole

same phase waves are added, Opp. phase waves are substracted.





* Planck's Quantum Theory

- * atoms or molecules of a substance can emit or absorb electromagnetic radiations.
- * atoms/molecules emit energy in form of small packets, these small packets are called 'QUANTA'
- * Smallest packet of energy is called 'QUANTUM'

energy of quanta ∝ Frequency of radiations
(E) (ν)

(h) → Planck's constant

$\boxed{E = h\nu}$

$$\boxed{\begin{aligned} h &= 6.626 \times 10^{-34}\ \text{J.s} \\ &= 6.63 \times 10^{-34}\ \text{Js} \end{aligned}}$$

- * energy of one quanta = $h\nu$
- * energy of n quanta = $n(h\nu)$

(n : 1, 2, 3, ...) n → No fractional value.

- Quantum is the smallest packet that can emit or absorb by a substance in form of electromagnetic radiations.

- Atoms/molecules could emit/absorb energy in discrete quantity, not in arbitrary amount.

  ↓ Integral value of n

- In case of light, quantums are called **PHOTONS**

$$E = h\nu$$

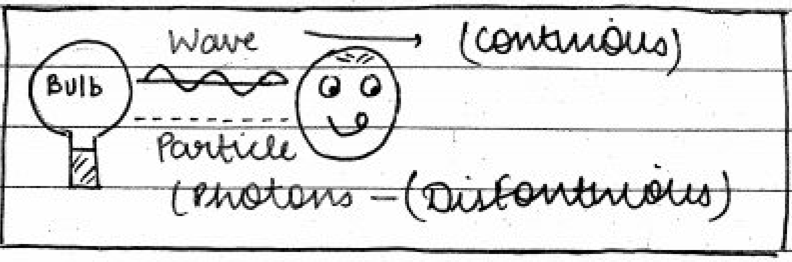
$$\boxed{c = \nu\lambda}$$

energy of one photon/quantum :-

$$\boxed{E = \frac{hc}{\lambda}}$$

energy of n photons :- $E = \frac{n(hc)}{\nu}$.

- Planck's theory supported ~~particle~~ 'nature' of electromagnetic radiation.

- 

DUAL NATURE OF electromagnetic radiation.

* Black Body Radiations

- * a body that can emit or absorbs of all frequencies is called 'black body'.
- * Radiations from black body is called 'Black body Radiations'.

Heating of substance {Black body}

$\boxed{Fe}$

↳ Red → yellow → Blue → white.

- * amount of emitted energy depends on wavelength of radiations
- * supported - Planck's Theory

Do not supported by - Maxwell theory.

Q. E.R. emitted by FM at a frequency of 1240 KHz. What is wavelength ?

Sol.

$$3 \times 10^{8} m/s = 1240 \times 1000 \times \lambda$$

$$\frac{3}{1240} \times 10^{5} = \lambda$$

$$\boxed{2.4 \times 10^{2} m}$$

Q. calculate frequency of E.R. having wave no. $5 m^{-1}$.

Sol.

$$\lambda = \frac{1}{\bar{\nu}}$$

$$3 \times 10^{8} = \frac{1}{5} \times \nu$$

$$\boxed{15 \times 10^{8} s^{-1}} = \nu.$$

Q. What is the ratio of energies E_1, E_2 of two radiation one with wavelength 6000 Å 2000 Å (λ_2)

$$\frac{E_1}{E_2} = \frac{\lambda_2}{\lambda_1} = \frac{1}{3} \text{ Ans.}$$

$E = hc/\lambda$.

Q. calculate no. of photons of light having $\lambda = 100 nm$ which can provide 1 KJ energy. ($h = 6.63 \times 10^{-34} J.s$)

Sol.

$$n = \frac{E \times \lambda}{hc} = \frac{1000}{6.63 \times 10^{-34} \times 3 \times 10^{8}} \times 100 \times 10^{-9}$$

$$= \frac{1}{6.63 \times 3} \times 10^{+22} = \boxed{5 \times 10^{20}}$$

Q. Find the no. of quantum of radiations of frequency $10^{13} s^{-}$ that must be absorbed in order to melt 6.626 g of ice. The energy required to melt 2g of ice is 666 J.

Sol.

$$2 — 666$$

$$6.626 — \frac{666}{2} \times 6.626$$

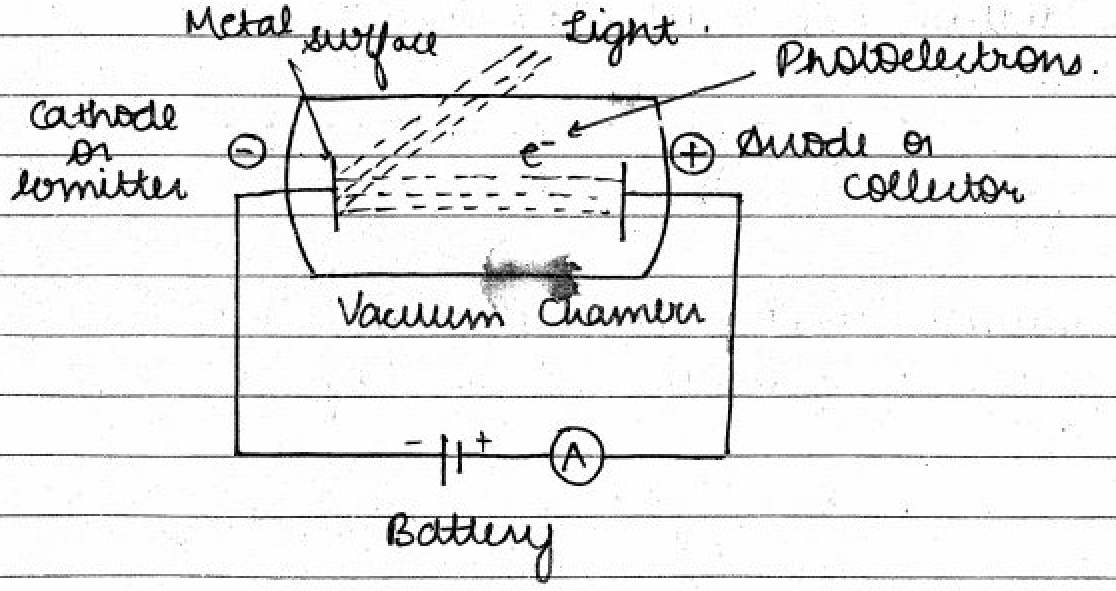
$$E = n(h\nu)$$

$$6.626 \times 333 = n(6.663 \times \times 10^{-34} \times 10^{13})$$

$$\boxed{3.33 \times 10^{23} = n}$$

★ PHOTOELECTRIC EFFECT

Photons ke kaaran electric effect!!



- ★ Einstien suggested a beam of light is stream of particles. Now these particles are called photons.

- ★ When photons of (appropiate frequency) [$\nu > \nu_0$] strike on metal surface, electrons are ejected from metal surface. These emitted electrons are called photo-electrons.

- ★ Phenomena is called :- 'Photoelectric Effect.'

- ★ Threshold Energy (E_0): Minimum energy required for ejection of electron from metal surface.
{Binding energy, Work function, Ionisation energy}
- ★ Threshold frequency (ν_0):
corresponding minimum frequency required for ejection of electron.

- ★ Energy of incident photons on metal surface = (energy required for ejection of e- from m. surface) + (Energy - Max K.E. of ejected photoelectron)

(we talk about max. because of the electrons present at outermost shells)

$$\boxed{E = E_0 + \frac{1}{2}mv^2}$$

∴ Energy of photon = $h\nu$

$$h\nu = h\nu_0 + \frac{1}{2}mv^2$$

∴ Max. K.E. of photoelectrons

$$\boxed{\frac{1}{2}mv^2 = h\nu - h\nu_0}$$

$$\boxed{\frac{1}{2}mv^2 = \frac{hc}{\lambda} - \frac{hc}{\lambda_0}}$$

λ_0 : Threshold wavelength.

★ $\boxed{\frac{1}{2}mv^2 = h\nu - \phi}$ $\boxed{\phi \text{ or } W = \text{Work function}}$.

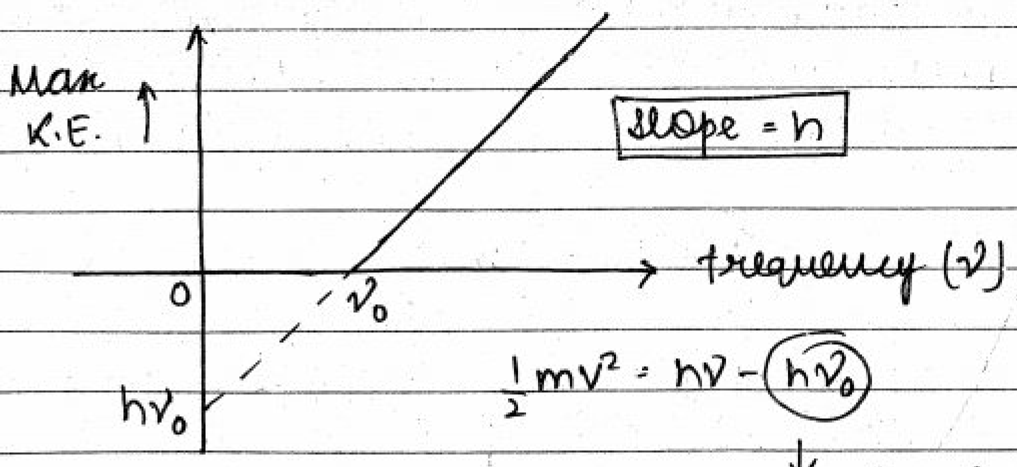
Conclusions :

★ If $\nu < \nu_0$, then no ejection of e^- from metal surface.

★ If $\nu = \nu_0$, then ejection of e^- takes place with zero K.E.

★ If $\nu > \nu_0$ then ejection of e^- takes place with specific K.E.

★



$$\frac{1}{2}mv^2 = h\nu - \textcircled{h\nu_0}$$

↓
negative intercept

$y = mx + c$

★ Intensity of light :-

Energy incident per unit area per unit time.

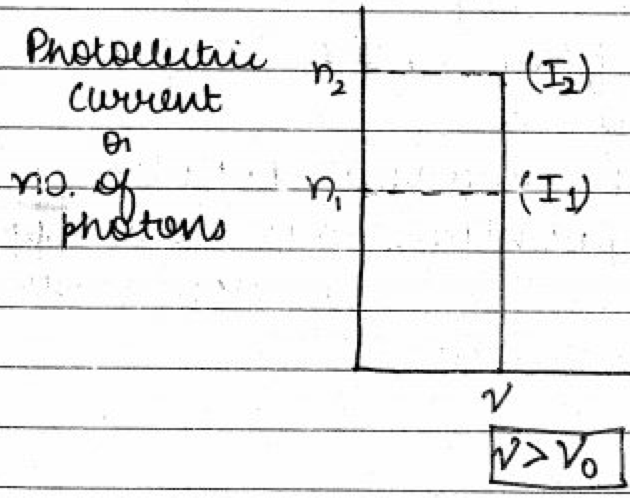
$$\boxed{I = \frac{n(h\nu)}{A.t}}$$

★ For a particular light source :
constant frequency.

★ n = no. of photons, can be varied by bringing the source away and towards the metal surface.

$n\uparrow$: $I\uparrow$: more photoelectrons ejected : Photoelectric current ↑

★ Kinetic energy is same.



When $n_2 > n_1$

★ Work function is dependent on nature of metal surface.

★ If wavelength of photons is greater than threshold wavelength then no ejection of electrons from metal surface.

$$\boxed{\nu\downarrow : \lambda\uparrow : E\downarrow}$$

★ $$\text{Power} = \frac{\text{energy}}{\text{time}}$$

$VI = \frac{E}{t}$

$$VI = \frac{E}{t}$$

$$\boxed{\frac{\text{Charge}}{\text{time}} = \text{current}}$$

$\boxed{Vq = E}$

$$E = VIt$$

★ $$\boxed{E = Vq}$$

$$\boxed{q = IT}$$

q = charge

V = voltage

★ Stopping Energy & Stopping Potential :

Min. Energy required to stop the photo current or ejection of e^- from metal surface is stopping energy and corresponding required potential is called stopping potential. (minimum app. volt. to stop photoelection)

★ for photoelectron

$$\boxed{q = e}$$

$$\therefore E = e.V.$$

$$\boxed{1\,eV = 1.6 \times 10^{-19}\,J}$$

1eV : energy of electron being accelerated by potential of 1 Volt.

$Q = It$

$$\boxed{I = \frac{Q}{t} = \frac{n.e}{t}}$$

no. of e^-

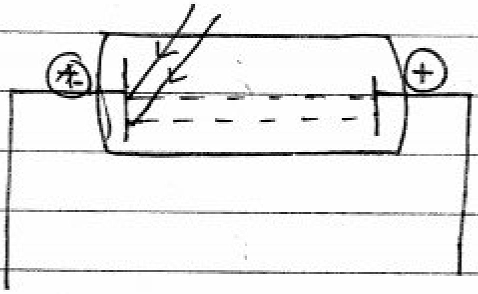
charge of e $= 1.6 \times 10^{-19}$ C

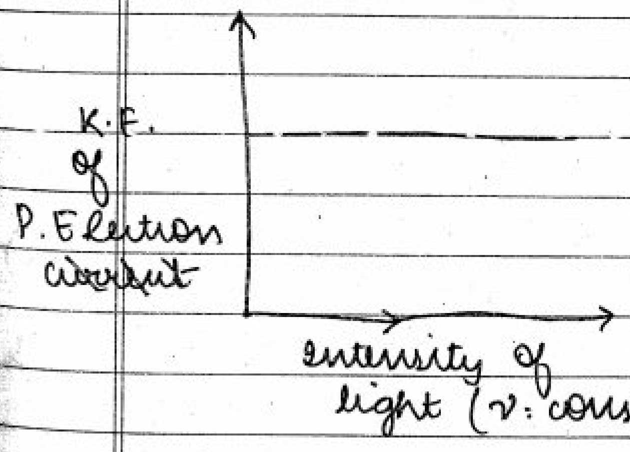
time (sec)

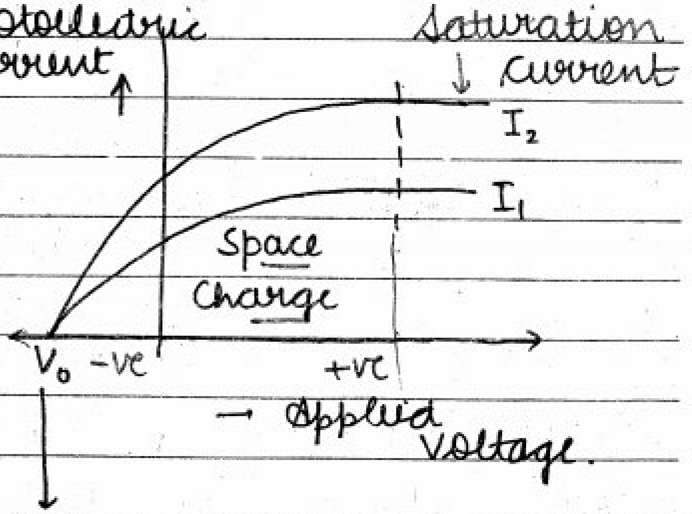
energy = $E = qV$

for photo electrons

$$\boxed{E = e.V}$$







stopping voltage: (comes into picture when battery connected reversely)

★ If intensity of light increases, photoelectric current ↑. (source of light remains same) But K.E. remains same.

★ **Stopping Energy** :

$$\boxed{E = q.V_0 = e.V_0}$$

★ $\boxed{V_0 \text{ or } V_s : \text{ stopping potential}}$

★ Max. K.E. of photoelectrons = $h\nu - h\nu_0 = h\nu - \phi$

$$\frac{1}{2}mv^2 = h\nu - h\nu_0$$

$$= eV_0$$

Work Function

$$\boxed{V_0 = \frac{h\nu}{e} - \frac{h\nu_0}{e}}$$

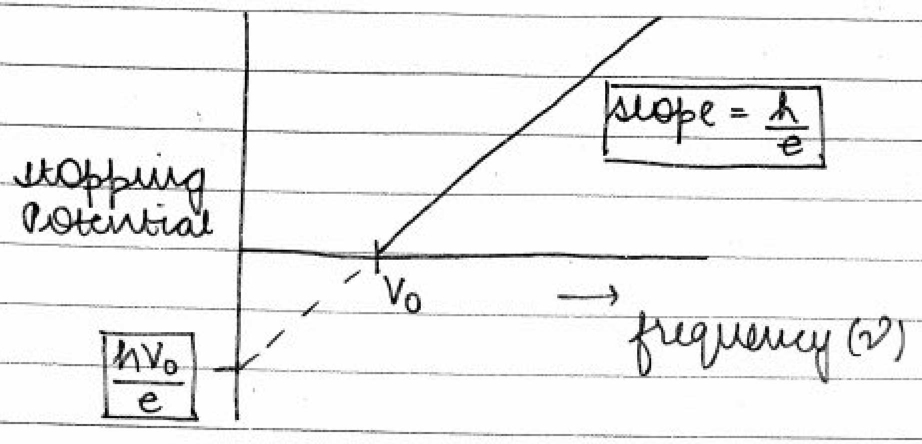
$$\boxed{V_0 = \left(\frac{h\cancel{\nu}c}{\lambda}\right)\frac{1}{e} - \left(\frac{hc}{\lambda_0}\right)\frac{1}{e}}$$

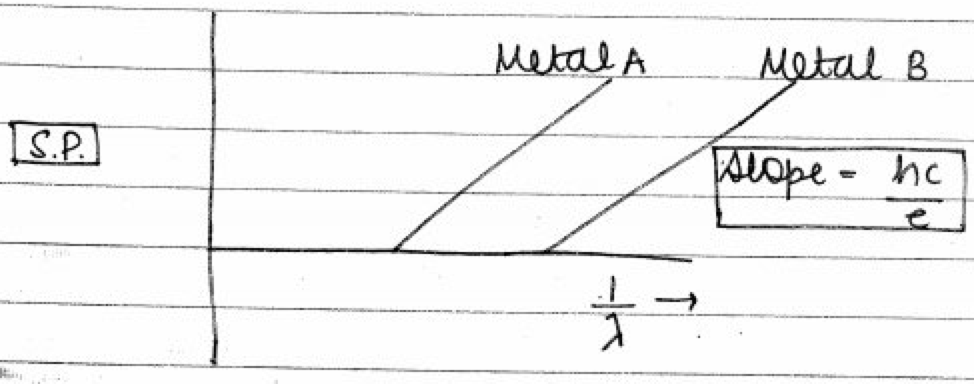
★★ Work Function depends on metal surface.

★★ Stopping potential depends on frequency of incident photons and work ~~done~~ function.

★★ $\boxed{V_0 \propto \text{K.E. of photo } e^-}$

directly proportional.





A Photoelectric effect supported particle nature of electromagnetic radiation.

★ (1) calculate K.E. of a photoelectron (in eV) emitted by a sodium metal when light of wavelength 400nm strikes on it. The work function is 2.3 eV. Calculate stopping potential (in V).

(.8)V

Sol.

KE. =

$$V_0 = \frac{1240}{400} - 2.3$$

$$= 3.1 - 2.3$$

$$= \boxed{.8V}$$

★ Energy of photon :

★ In Joules :-

$$E = \frac{\overset{\nearrow (J.s)}{h}\,c \rightarrow (m/s)}{\lambda \rightarrow (m)}$$

$$\boxed{1\,eV = 1.6 \times 10^{-19}\,J}$$

$h = 6.626 \times 10^{-34}$ Js.

$c = 3 \times 10^{+8}$ m/s

★ In eV :

$$\boxed{E = \frac{1240}{\lambda\,(\text{in nm})}}$$

↓ nanometer.

Q. If the calculate threshold wavelength for producing photoelectric current is 200 nm. Calculate the wavelength of radiation to produce photoe⁻ having half the K.E. of those produced by radiation of wavelength 120 nm.

Sol.

$$(K.E.)_{max} \text{ of photo } e^- = \frac{hc}{\lambda} - \frac{hc}{\lambda_0}$$

$\lambda_2 = 150$ nm.

$$(KE)_{max} = \frac{1}{2}(KE)_{max}.$$

$$\left(\frac{1}{\lambda_2} - \frac{1}{\lambda_0}\right) = \frac{1}{2}\left(\frac{1}{\underset{\underset{120}{\uparrow}}{\lambda_1}} - \frac{1}{\underset{\underset{200}{\uparrow}}{\lambda_0}}\right)$$

$$\frac{200-\lambda_2}{200\lambda_2} = \frac{1}{2}\left(\frac{200-120}{200\times 120}\right)$$

$$120(200-\lambda_2) = 40\lambda_2$$

$$120\times 200 = 160\,\lambda_2$$

$$\lambda_2 = \frac{12}{16}\times 200 = 3\times 50$$

$$= \boxed{150\ nm}$$

* when surface of metal is exposed to radiation of 200 nm the electrons emitted have K.E. 10.52×10^4 J/mol. Calculate energy required to remove electrons from the metal surface (ionization energy) in J/mol.

492 kJ/mol.

Sol

$$(K.E)_{max} = \frac{hc}{\lambda} - I.E.(\phi)$$

$$10.52\times 10^4 = \frac{1240}{200}\times 1.6\times 10^{-19}\times 6.022\times 10^{23} - \phi$$

$$= \frac{1240}{200}\times 10^5 - \phi\ (J/mol)$$

$$\phi = 10^4(62-10.52)$$

* (1) Calculate no. of photons emitted by a 100 W bulb in 1 sec and in 1 hr that emit light of wavelength 620 nm also calculate max photo current (in amps) that can be generated by this bulb.

Sol. $100W = 100J/s$

$$E(\text{in eV})_{\text{1 photon}} = \frac{1240}{620} = 2\text{ eV.}$$

No. of photoelectrons ejected $= \dfrac{100}{2\times1.6\times10^{-19}}$

$$= 3.125\times10^{\cancel{19}\,20}$$

No. of photoelectrons emitted in 1 hr

$$= 3.125\times10^{20}\times60\times60$$

$$= \boxed{11.25\times10^{23}}$$

(I) current (in amp.)

$$= \frac{Q}{t}$$

$$= \frac{ne}{t} = \frac{\cancel{11.25\times10^{23}}}{3.125\times10^{20}\times1.6\times10^{-19}}$$

$$= \boxed{5.0\text{ amp}}$$

(ii) If emitted photons strikes a metal plate and only 25% electrons emitted, then calculate (I)

$$50\times.25 = 12.5\text{ amp.}$$

Imp.

(Q) A light source is emitting 3 radiations of wavelength 310 nm, 620 nm and 1000 nm. 3 watt power equally ~~an~~ distributed among the three wavelengths. If light beam strikes on a metallic surface having work function of 1.75 eV then calculate total no. of photons emitted per second and charge emitted per second. (current in amp.)

sol.

	1 watt	1 watt	1 watt
	310 nm	620 nm	1000 nm
E → (in eV)	4 eV	2 eV.	1.24 eV × No photoelectrons ejected.

Total no. of photons e⁻ emitted = $\frac{\text{Total energy}}{\text{energy from 1 photon}}$

$$= \frac{1}{4 \times 1.6 \times 10^{-19}} + \frac{1}{2 \times 1.6 \times 10^{-19}}$$

$$= 10^{19}(\cancel{6.4} + \cancel{3.2}) \quad \left(\frac{1}{6.4} + \frac{1}{3.2}\right)$$

$$= \cancel{9.6 \times 10^{19}}$$

$$= 4.6875 \times 10^{18}$$

$$I = \frac{ne}{t} = \frac{4.6875 \times 10^{18} \times 1.6 \times 10^{-19}}{1}$$

$$= .75 \text{ amp or } \textcircled{C/s}$$

* When light falls of wavelength 470 nm falls on surface of potassium metal, electrons emitted with velocity of 6.4×10^4 m/s

1.86×10^{-21} (a) calculate K.E of emitted electrons

4.22×10^{-19} (b) energy of photons 2.6 eV

4.2×10^{-19} (c) min. energy required to remove e^-

(a) $\frac{1}{2} \times 9.1 \times 10^{-31} \times (6.4 \times 10^4)^2 = $ K.E.

$1.86 \times 10^{-21} =$ KE.

(b) $\frac{1240}{470} = 2.6$ eV

(c) $2.6 \times 1.6 \times 10^{-19} - 1.86 \times 10^{-21}$

$= 4.2 \times 10^{-19}$ J

* BOHR'S ATOMIC MODEL:

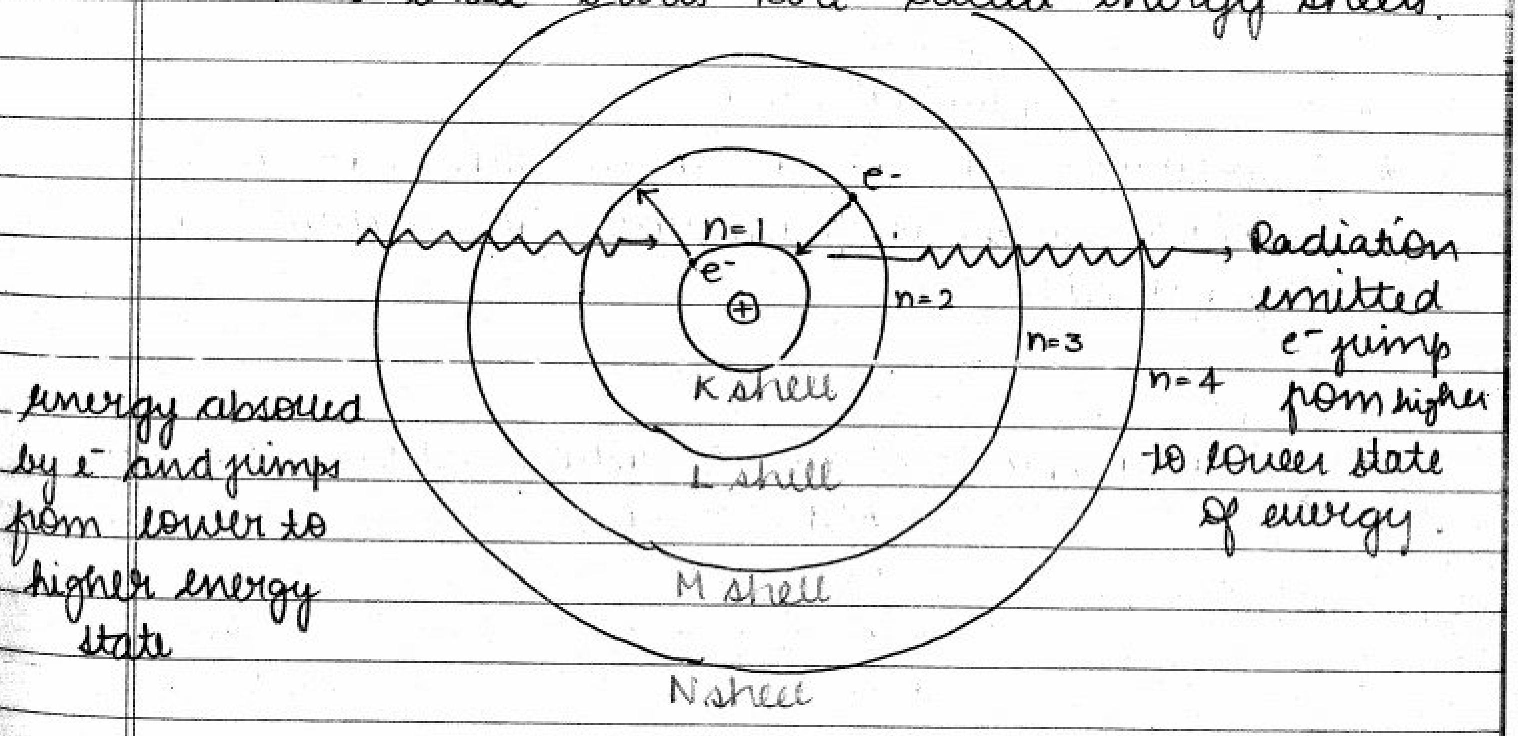
- Positively charged nucleus present at centre of atom.

- electrons revolve outside the nucleus in definite circular paths called orbits.

↓

called stationary orbits when electrons revolve in definite orbits. No energy emission takes place by electrons.

Now these orbits are called energy shells.



→ energy change when e⁻ jump first orbit to second orbit.

$$\boxed{\Delta E = E_2 - E_1 = \frac{hc}{\lambda}}$$

★ K-shell, $n=1$, ground state
★ L-shell, $n=2$, First excited state.

★ when electrons revolve in an orbit, angular momentum of electron is quantized and simple multiple of

$$\frac{h}{2\pi}$$

★ ANGULAR MOMENTUM: of e^-:

$$\boxed{M v r = n\left(\frac{h}{2\pi}\right)} \qquad n = 1,2,3\ldots\ldots \qquad \boxed{(iii)}$$

M_e = mass of electron
v = Velocity of e^-
r = radius of orbit

★ electrons revolve only in those orbits in which angular momentum of electrons is quantized.

★ Coulombic Force Of Attraction between Nucleus And Electron :-

$$= \frac{K\overset{\text{Nucleus charge}}{q_1}\,q_2}{r^2} = \frac{K(z.e).e}{r^2} \qquad \boxed{(i)}$$

z = atomic no.

and $F = ma = m\left(\frac{V^2}{R}\right)$ $\boxed{(ii)}$

where $\frac{V^2}{R}$ is centripetal acceleration

★ equating (i), (ii)

⌖ ★ $\boxed{\frac{mV^2}{R} = \frac{kze^2}{R^2}}$

$K = 9 \times 10^9\ Nm^2C^{-2}$

$= \frac{1}{4\pi E_0}$

↓

Permittivity of vacuum.

$v^2 = \frac{kze^2}{r.m}$

←★ equating (ii) and (iii)

$$\frac{n^2h^2}{4\pi^2} \cdot \frac{1}{m^2r^2} = \frac{kze^2}{r.m}.$$

$$r = \left(\frac{n^2h^2}{4\pi^2 m}\right)\left(\frac{1}{kze^2}\right)$$

$$\boxed{r = \left(\frac{n^2}{z}\right)\left(\frac{h^2}{4\pi r^2 M_e ke^2}\right)}$$ $\boxed{(iv)}$

⌖ $\boxed{r = 0.529 \times \frac{n^2}{z}\ A^\circ}$ n = no. of orbit.

$\boxed{r \propto \frac{n^2}{z}}$

★ equating (iii), (iv),

$$mvr = \frac{nh}{2\pi}$$

$$V = \frac{nh}{2\pi} \times \frac{1}{mr} = \frac{nh}{2\pi M} \times \frac{4\pi^2 Mk.e^2.z}{n^2h^2}$$

$$V = \frac{2\pi Ke^2}{h}\left(\frac{z}{n}\right)$$

$$\boxed{V = 2.18 \times 10^6 \left(\frac{z}{n}\right) ms^{-1}}$$

Very Important

$$\boxed{V \propto \frac{z}{n}}$$

* Frequency of Revolution of e^- (f):
No. of revolution of electrons per SECOND.

$$\boxed{f = \frac{v}{2\pi r}}$$

$V \propto (z/n)$
$r \propto (n^2/z)$

$$\boxed{f \propto \frac{z^2}{n^3}} \qquad \boxed{\frac{f_1}{f_2} = \left(\frac{z_1}{z_2}\right)^2\left(\frac{n_2}{n_1}\right)^3}$$

* Time Required (T): Time Period
for complete revolution.

$$\boxed{T = \frac{1}{f} = \frac{2\pi r}{v}} \text{ seconds}$$

★ K.E. of electron = $\frac{1}{2}mv^2$:

⊕ ∴ $\boxed{K.E. \text{ of } e^- = \frac{1}{2}\left(\frac{kze^2}{r}\right)}$

using:
$$\frac{mv^2}{r} = \frac{kze^2}{r^2}$$

⊕ $\boxed{P.E. \text{ of } e^- \text{ and nucleus } = -\frac{kq_1q_2}{r} = -\left(\frac{kze^2}{r}\right)}$

★ Change in Potential energy = -Workdone

$$E_{P.E.(r)} - \underset{\boxed{0}}{\underset{\downarrow}{E_{P.E.(\infty)}}} = -W$$

$$= -\frac{kze^2}{r^2}$$

★ At infinite distance, no attraction between e⁻ and nucleus.

★ P.E of system has been assigned zero values (man.)

★ Total energy of atom (electron) = KE + PE

$$= \frac{1}{2}\cancel{m}\left(\frac{kze^2}{r}\right) - \frac{kze^2}{r}$$

⊕ $\longrightarrow \quad = \boxed{-\frac{1}{2}\left(\frac{kze^2}{r}\right)}$ $\boxed{(V)}$

conclusions :

★(1) 2x Total energy of atom = Potentail energy

★(2) K.E = - T.E.

★(3) 2x K.E. = - P.E.

equating (iv) , (v)

$$\text{Total energy} = -\frac{1}{2} kze^2 \left(\frac{4\pi^2 m k e^2 z}{n^2 h^2} \right)$$

$$= \frac{-2\pi^2 k^2 e^4 . Me}{h^2} \left(\frac{z^2}{n^2} \right)$$

imp

⊕ $\boxed{\text{Total energy} = E_T = -13.6 \left(\frac{z^2}{n^2} \right) \text{ eV/atom}}$

$$1 eV = 1.6 \times 10^{-19} J$$

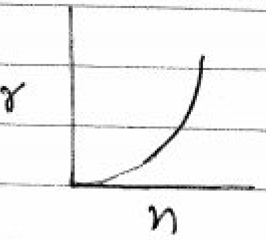
$$E_T = -13.6 \times 1.6 \times 10^{-19} \left(\frac{z^2}{n^2} \right) \text{J/atom}$$

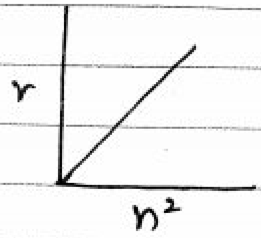
⊕ $\boxed{E_T \simeq -2.18 \times 10^{-18} \left(\frac{z^2}{n^2} \right) \text{J/atom}}$

* GRAPHS :-

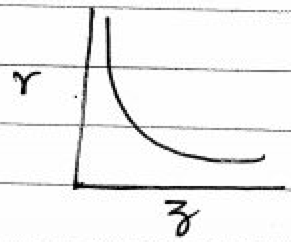
* Bohr model is applicable for single electronic species.
(H, He$^+$, Li^{+2}, etc.)

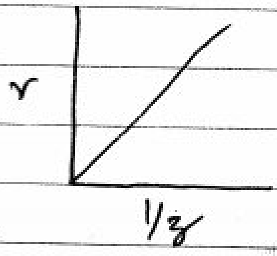
* At constant Z :





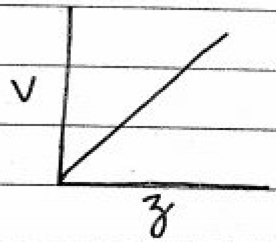
* At constant n :



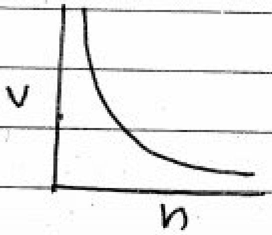


$\boxed{yx = k}$

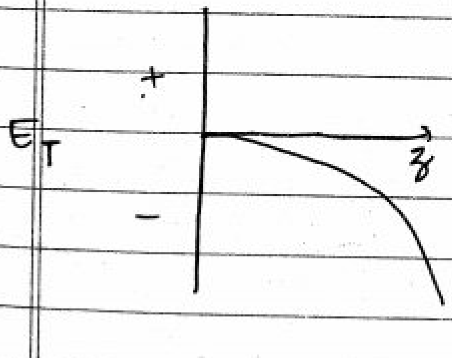
* At constant n :

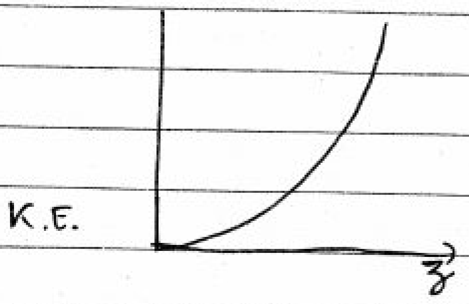


At constant Z :



* At constant n :





Q. calculate radius ratio $\left(\frac{r_1}{r_2}\right)$ where r_1 = radius of second orbit of H-atom.
r_2 = radius of third orbit of He^+

sol.

ans. → $\boxed{8/9}$ $\frac{r_1}{r_2} = \left(\frac{n_1}{n_2}\right)^2\left(\frac{z_2}{z_1}\right) = \frac{8}{9}$

Q. calculate velocity of e^- of He^+ present in K sheel (first orbit).

sol.

$2.18 \times 10^6 \left(\frac{2}{1}\right)\ ms^{-1}$

Q. Calculate the value of $\frac{f_1}{f_2}$ where f_1 = frequency of resolution of e^- in IInd orbit of H atom
3/8
f_2 = ______ III orbit Li^{2+} ion.

$$\frac{f_1}{f_2} = \left(\frac{z_1}{z_2}\right)^2\left(\frac{n_2}{n_1}\right)^3$$

ans → $\frac{3}{8}$.

☆ $\boxed{T.E. = -13.6\left(\frac{z^2}{n^2}\right)}$ eV/at

For H-atom : z=1

★ energy associated with e^- present in K-shell or energy of K-shell :

$$\boxed{E_1 = -13.6 \text{ eV/atom}}$$ (Minimum)

★ energy of L-shell

$$\boxed{E_2 = \frac{-13.6}{4} \text{ eV/atom}}$$

$= -3.4$ eV/atom

★ energy of M-shell n = 3.

$$\boxed{E_3 = -1.51 \text{ eV}}$$

★ energy of N-shell , n = 4

$$\boxed{E_4 = -0.85 \text{ eV/atom}}$$

(Max

energy Change :-

$$\Delta E = E_f - E_i$$

$$= E_2 - E_1 = -3.4 - (-13.6)$$

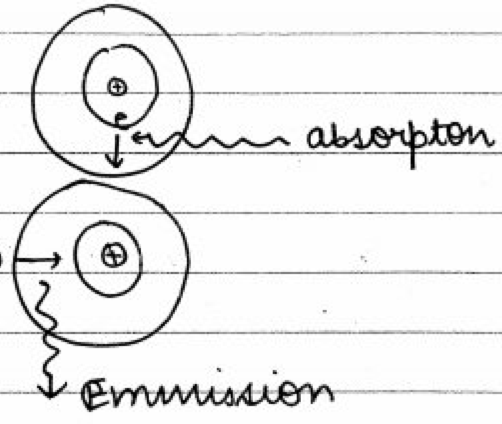
$$= 10.2 \text{ eV/atom}$$

★ $\boxed{\Delta E = E_2 - E_1 = 10.2 \text{ eV/atom}}$

Number

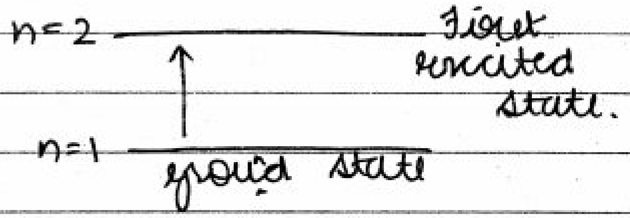
$-E$ → energy emission
E → energy absorption.

★ $\Delta E = E_1 - E_2$
$= -10.2$ eV/atom



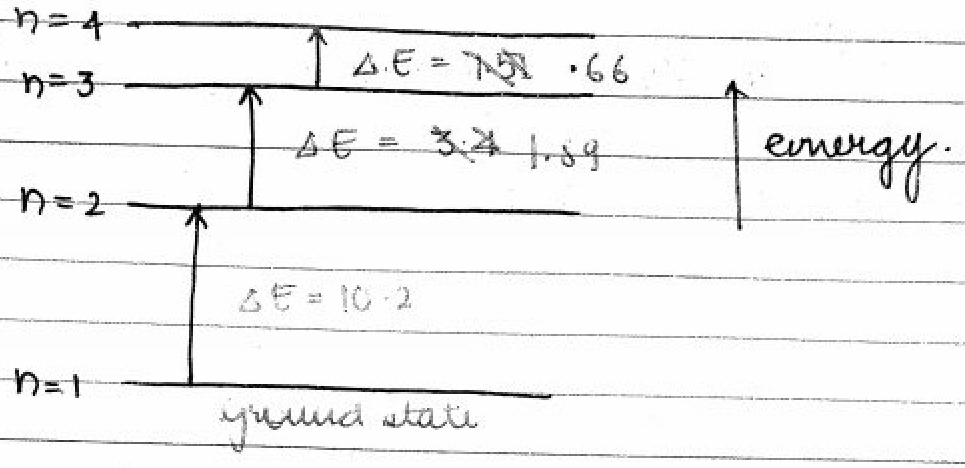
★ $\Delta E_{1\to3} = E_2 - E_3$
$= -1.51 - (-3.4)$
$= 1.89$ eV/atom.

★ $\Delta E_{3\to4} = E_3 - E_4$
$= -.85 - (-1.51)$
$= 0.66$ eV/atom.



★ energy difference between energy shells decreases with increase in n

$\boxed{n\uparrow : E\downarrow}$



—x—

* Calculate energy required to separate e⁻ from gaseous state from gaseous atom of hydrogen

$$(H(g) \rightarrow H_g^+(g) + e^-)$$

$$\Delta E = E_\infty \rightarrow E_1$$
$$= 0 - (-13.6)$$
$$= 13.6 \text{ eV/atom}$$

* Rydberg equation for H-atom

[* Using experiment. value of R.]

$$\frac{1}{\lambda} = R_H z^2\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$ where $z=1$

* $$\boxed{\frac{1}{\lambda} = R_H\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)}$$ R_H ~~α R~~

where R_H = Rydberg's constant

= 109678 or

$\boxed{109700 \text{ cm}^{-1}}$

Value matched by using equations of Bohr.

* ACCORDING TO BOHR:

$$\Delta E = E_f - E_i = \frac{2\pi^2 K^2 e^4 m_e z^2}{h^2}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

$$\Delta E = \frac{hc}{\lambda} = \boxed{\frac{2\pi^2 K^2 e^4 m_e}{h^2}}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

$\downarrow$

Wave no. = $$\bar{\nu} = \frac{1}{\lambda} = R . z^2\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

$\downarrow$

where $$R = \frac{2\pi^2 k^2 e^4 . m_e}{h^3 c}$$

$= 109678 \text{ cm}^{-1}$

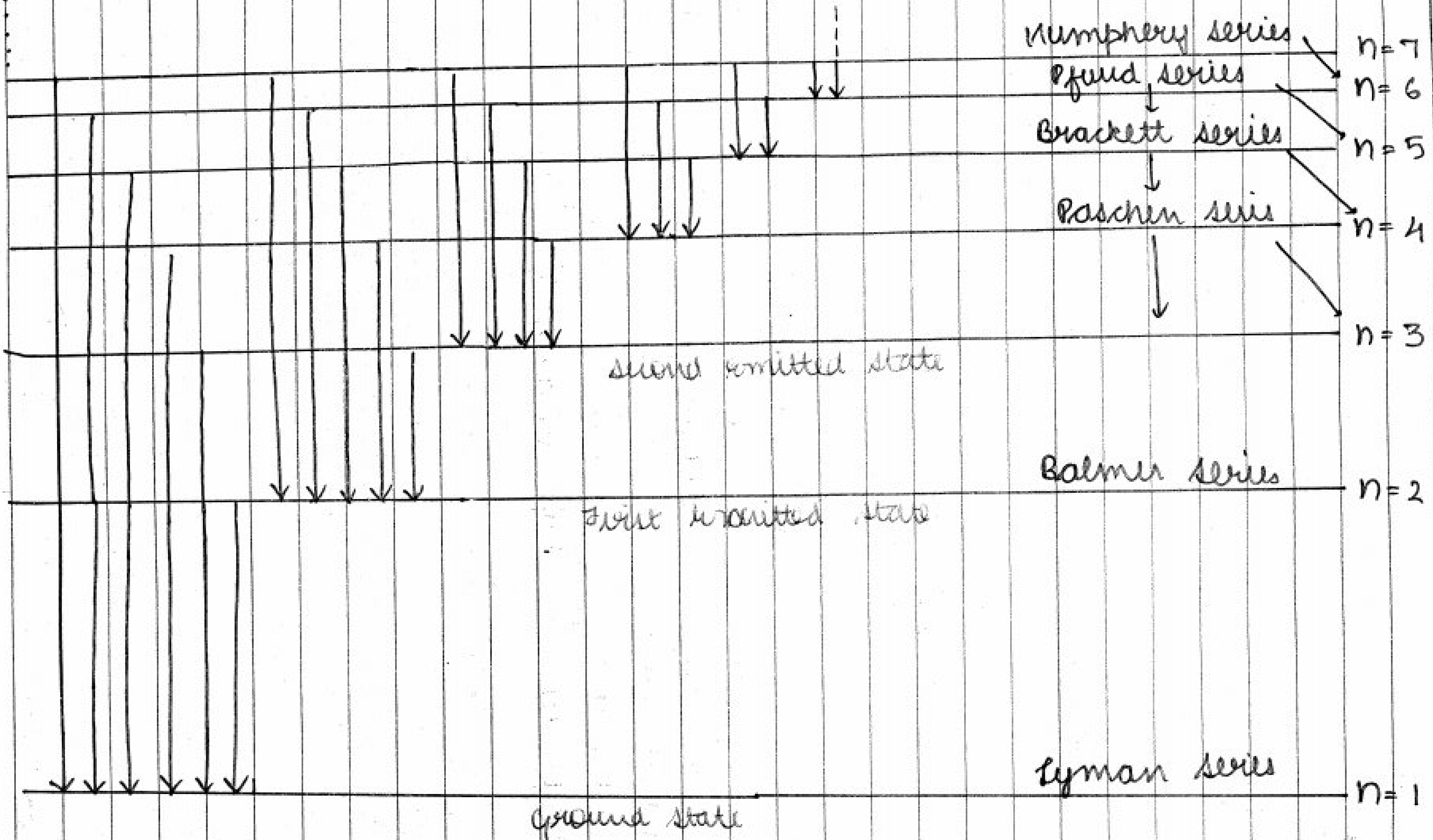
$\simeq 109700 \text{ cm}^{-1}$ * Same as R_H

★ $\boxed{\frac{1}{\lambda} = R_H \cdot Z^2 \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)}$

↳ Wavelength of emmission/ radiation (photon)

★ HYDROGEN EMISSION SPECTRUM (z=1)

★ series	Spectral Region (or spectrum)	n_f	n_i
1) Lyman series	ultraviolet	1	2,3,4. . .
2. Balmer series	Visible	2	3,4. . .
3. Paschen series	Infrared	3	4,5. . . .
4. Brackett series	— " —	4	5,6. . .
5. Pfund series	— " —	5	6,7. . . .
6. Humphery	Far infrared.	6	7,8. . . . .


Humphery series
Pfund series
Brackett series
Paschen series
Balmer series
Lyman series
n=7
n=6
n=5
n=4
n=3
n=2
n=1
second excited state
First excited state
Ground state

* CACCULATION OF MIN. & MAX. WAVELENGTH OF PHOTON (Radiation) EMISSION :- H atom.

* For **Lyman series**

$$\boxed{n_f = 1}$$

$$\Delta E_{min} = \frac{hc}{\lambda_{max.}}$$

$$\lambda_{max} = \frac{4}{3R_H} \qquad \lambda_{min} = \frac{1}{R_H}$$

* **Balmer series**

$$\lambda_{max} = \frac{36}{5R_H} \qquad \lambda_{min.} = \frac{4}{R_H}$$

* **Paschen series**

$$\lambda_{max} = \frac{144}{7R_H} \qquad \lambda_{min.} = \frac{9}{R_H}$$

* **Brackett series**

$$\lambda_{max.} = \frac{400}{9R_H} \qquad \lambda_{min.} = \frac{16}{R_H}$$

★ **PFund series**.

$\lambda_{max.} = \frac{900}{11 R_H}$ $\lambda_{min.} = \frac{325}{R_H}$

★ **Humphery series**.

$\lambda_{max} = \frac{1764}{13 R_H}$ $\lambda_{min} = \frac{36}{R_H}$

★ Lyman series :-
lower energy level
$\boxed{n_f = 1}$

★ For first line transition (H_α line transmition)
$\boxed{n_i = 2}$

★ For second line transition (β line transition)
$\boxed{n_i = 3}$

★ For third line transition (γ line ")
$\boxed{n_i = 4}$

★ For fourth " " (δ line transmision)
$\boxed{n_i = 5}$

☆ Balmer series :-

lower energy level :- $\boxed{n_f = 2}$

★ First line transition : $\boxed{n_i = 3}$ (α)

★ second (β) line transition $\boxed{n_i = 4}$

★ Third (γ) line transition $\boxed{n_i = 5}$

Q. Calculate wavelength (in R) of radiation emitted when e^- transition takes place from third excited state to ground state for He^{+2}.

sol.

$$\frac{1}{\lambda} = R.z^2\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

$$= R\times4\left(\frac{1}{16} - 1\right)$$

$$\frac{1}{\lambda} = \frac{15}{4}R$$

$$\boxed{\lambda = \frac{4}{15}R}$$

Q. Calculate wave no. of radiation emitted for β line in case of Balmer series of H atom.

sol.

$$\frac{1}{\lambda} = R\left(\frac{1}{4} - \frac{1}{16}\right)$$

$\bar{\nu}$ = Wave no. $\boxed{= \frac{3R}{16}}$ Ans.

Q. calculate total energy, P.E and K.E. of electron for L shell, M-shell of Li^{2+}

sol. **L-shell**

$$T.E. = -13.6 \times \frac{Z^2}{n^2} \text{ eV/atom}$$

$$= -13.6 \times \frac{9}{4} = -3.4 \times 9$$

$$= -30.6 \text{ eV/atom}$$

$E_T = -30.6$ eV/atom

$E_p = 2 \times -30.6 = 61.2$ eV/atom

$E_k = 30.6$ eV/atom.

M-shell:

$$T.E. = -13.6 \times \frac{9}{9}$$

$$= -13.6 \text{ eV/atom.}$$

$E_T = -13.6$ eV/atom

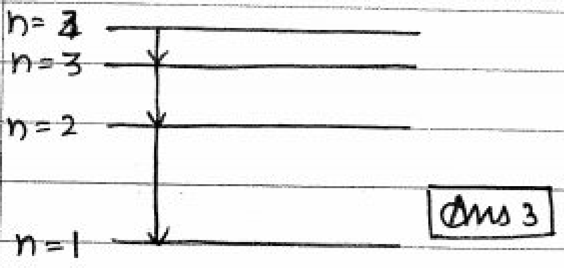
$E_K = 13.6$ eV/atom

$E_p = 27.2$ eV/atom.

imp for JEE :

★★ $\boxed{\text{MAXIMUM NO. OF RADIATIONS}}$ of different wavelength emitted when $\boxed{\text{an electron}}$ transition takes place from higher energy state to ground state ($n_f = 1$)

$$\boxed{= n_i - 1}$$



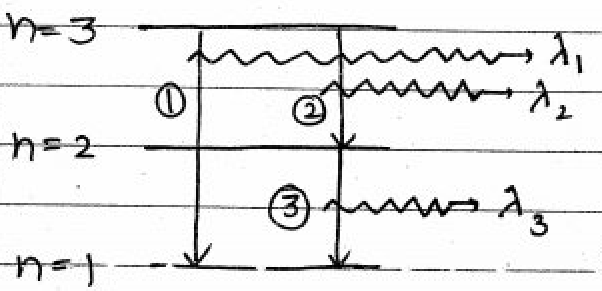
$\boxed{\text{ans 3}}$

{ sample of H atom / He^+ ion / Li^{2+} ion }

★★ $\boxed{\text{MAXIMUM NO. OF RADIATIONS}}$ of different wavelength emitted when $\boxed{\text{electrons}}$ transition takes place from higher energy state to ground state

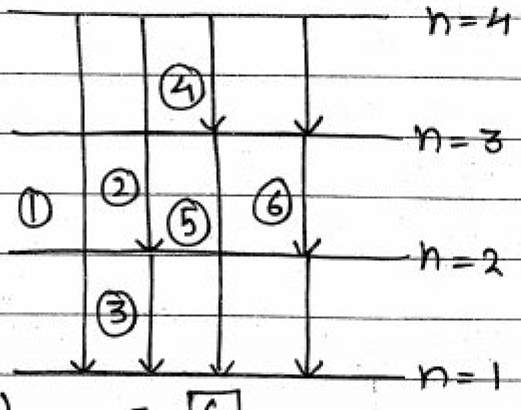
$$= \boxed{\frac{n(n-1)}{2}}$$

{sample of sufficient H atoms / He^+ ions / Li^{2+} ions)



ans ③

★ When n=4.



ans $\frac{4(4-1)}{2} = \boxed{6}$

★ To Prove the formula:

= Transition related to Lyman series
+ " " " Balmer series
+ " " " Paschen series

. . . . .

. . . . .

$= (n-1) + (n-2) + (n-3) \ldots\ldots 3, 2, 1$

$= \boxed{\frac{n}{2}(n-1)}$

★★ MAXIMUM NO. OF RADIATIONS of different λ when electrons transition take place from one excited state to another excited state

$= \frac{(n_2-n_1)(n_2-n_1+1)}{2}$

$= \boxed{\frac{N(N+1)}{2}}$ $\boxed{N = n_2 - n_1}$

Proof:

$= (n_2-n_1) + (n_2-n_1-1) + (n_2-n_1-2) \ldots\ldots 3, 2, 1$

$= N + (N-1) + (N-2) \ldots\ldots 3, 2, 1$

$= N + \frac{N(N-1)}{2}$

$= \boxed{\frac{N(N+1)}{2}}$

** Maximum ________ when one electron ________.

$\boxed{n_2 - n_1}$

*Q. Due to heating of a substance (unielectronic) excited electrons of atoms present in 8 th excited state, then

(i) calculate max. no. of radiation which can be observed when falls on ground state.

(ii) max. no. of radiations when electrons fall to 4th excited s. form 8th excited state.

(iii) max. no. of radiations in both cases in case of one electron.

Sol.

(i) $\frac{8\times 9}{2}$ = ~~28~~ 36

(ii) $\frac{4\times 5}{2}$ = 10

(iii) 8

(iv) 4

*Q. A gas emits radiations of wavelength 45.9 nm., 82.8 nm, 103.33 nm. Assume that atoms has only 2 excited and a ground state and the difference between consecutive energy level decreases as energy is increased. Consider energy of highest state as zero. calculate energy of Ist level and Ist excited state.

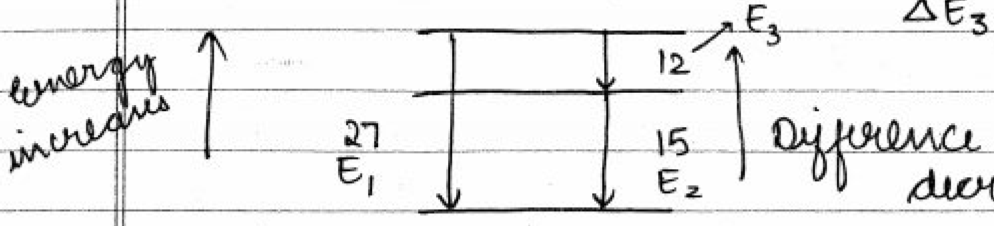
Sol.

$\lambda \propto \frac{1}{E}$

(i) $\underline{1240}$

$\Delta E_1 = \frac{1240}{45.9} = 27 \text{ eV/atom}$ $\Delta E_2 = \frac{1240}{82.8} = 15 \text{ eV/atom}$

$\Delta E_3 = \frac{1240}{103.33} = 12 \text{ eV/atom}$



$\lambda_1 = 45.9$ nm
$\lambda_2 = 82.8$ nm
$\lambda_3 = 103.33$ nm

$\Delta E_1 = E_3 - E_1 = 27 \text{ eV/atom}$
$0 - E_1 = 27 \text{ eV/atom}$
$\boxed{E_1 = -27 \text{ eV/atom}}$ ans.

$\Delta E_3 = E_3 - E_2 = 12 \text{ eV/atom}$
$\boxed{E_2 = -12 \text{ eV/atom}}$ ans.

—x—

Q. A hydrogen like ion sample emit six radiations of diff. λ. find the possible transition between group of levels. these levels have energy from $\boxed{-.85 \text{ to } -.544}$ eV/atom

Calculate :
(i) the quantum no. of level between which transition is taking place
(ii) find the atomic no of ions of substance.

sol $\frac{(n_2-n_1)(n_2-n_1+1)}{2} = 6$

$(n_2-n_1) = 3$

E at higher level $= -.544 = -13.6\left(\frac{z^2}{n_2^2}\right)$

$E_2 = -.85 = -13.6\left(\frac{z^2}{n_1^2}\right)$

$\frac{544}{850} = \left(\frac{n_1}{n_2}\right)^2$

$= \frac{272}{425}$ $.64 = \left(\frac{n_1}{n_2}\right)^2$

$\boxed{\frac{4}{5} = \frac{n_1}{n_2}}$

$9 = \frac{n_1+n_2}{3}$

$27 = n_1+n_2$

$3 = n_2-n_1$

$\boxed{\begin{matrix}15 = n_2\\12 = n_1\end{matrix}}$

$-.85 = -13.6\left(\frac{z^2}{144}\right)$

Ans $\boxed{z=3}$

★ IONISATION ENERGY (I.E.)

energy necessary to remove (Binding energy)
electron present in isolated gaseous atom.

$$\underset{\text{gaseous atom}}{H(g)} \rightarrow \underset{\text{Cation}}{H^+(g)} + e^-$$

★ for H atom :- e^- present in $\boxed{\text{ground state}}$

★ for other atoms :- e^- present in $\boxed{\text{valence shell}}$

$$X(g) \longrightarrow X^+(g) + \underset{\substack{\text{(from valence shell)} \\ (\text{Max. } e^-)}}{e^-}$$

★ electron must be removed from either ground state (for H atom) or from valence shell.

★★ ionization energy of H-atom :-

$$\Delta E = \overset{0}{\cancel{E_\infty}} - E_2 = -13.6\left(\frac{1}{1}\right)^2$$

$$= \text{I.E.}$$

for H-atom

$$\therefore \text{I.E.} = 13.6 \text{ eV/atom}$$
$$= 13.6 \times 1.6 \times 10^{-19} \text{ J/atom}$$
$$= 13.6 \times 1.6 \times 10^{-19} \times 6.022 \times 10^{23} \text{ J/mole}$$
$$= \boxed{1310 \text{ KJ/mole}}$$

$\boxed{\text{IE}}$

** experimental value of H atom for 1 mole
$\simeq$ 1312 KJ/mole.

* Ionization energy :-

*
$$\underset{1\,\text{mole}}{H(g)} \longrightarrow \underset{1\,\text{mole}}{H^+(g)} + \underset{1\,\text{mole}}{e^-} \qquad \Delta_r H \text{ or}$$

$$\boxed{\Delta H = 1312 \text{ KJ/mole}}$$

H : enthalpy
ΔH : change in enthalpy

* Binding energy :-

$$H^+(g) + e^- \longrightarrow H(g)$$

$$\boxed{\begin{aligned}\Delta H &= \text{Binding energy} \\ &= -1312 \text{ KJ/mole}\end{aligned}}$$

* BOND ENERGY :-

Energy required to break 1 mole of specific bonds present in substance.

$H_2(g) \longrightarrow 2H(g) \qquad \Delta_r H = \text{Bond E.}$
$\qquad\qquad\qquad\qquad\qquad\quad = x \text{ KJ/mole}$

$Cl_2(g) \longrightarrow 2Cl(g) \qquad \Delta_r H = y \text{ KJ/mol}$
Bond dissociation

$2Cl(g) \longrightarrow Cl_2(g) \qquad \Delta_r H = -y \text{ KJ/mol}$
Bond formation

$\boxed{\Delta H = q_p}$ = Heat absorbed by system at const. pressure.

$$A(g) + B(g) \longrightarrow 2C(g)$$

1mol 1mol 2mol

$\boxed{\Delta H = -y\,kJ}$ y KJ heat evolved at const. pressure in reaction.

* in bond dissociation → energy required +ve

* in bond formation → energy released -ve

* PHOTOCHEMICAL REACTION :

Reaction takes place in light.

$$HBr(g) \xrightarrow{h\nu} H(g) + Br(g)$$

$\boxed{E = h\nu}$ = energy of photon

* photons absorbed by HBr molecule and converted into respective atoms.

$$H_2(g) \xrightarrow{h\nu} H(g) + H(g) \qquad \Delta_r H = B.E.*(H-H)$$

$$\boxed{E = \frac{hc}{\lambda}}$$

* QUANTUM YIELD OR EFFICIENCY (ϕ) :-

$$\phi = \frac{\text{No. of molecule of reactant reacted}}{\text{No. of photons absorbed.}}$$

or

$$\phi = \frac{\text{no. of moles of reactant reacted}}{\text{no. of moles of photon absorbed.}}$$

Q. The electromagnetic radiation of wavelength 242 nm. is just sufficient to ionise the sodium atom. calculate ionization energy of sodium in KJ/mol.

Sol.

$$\frac{1240}{242} = 5 \text{ eV/atom}$$

$$= 5 \times 1.6 \times 10^{-19} \text{ J/atom}$$

$$= \frac{5 \times 1.6 \times 10^{-19} \times 6.02 \times 10^{23}}{1000}$$

$$= \frac{8 \times 6.02 \times 10^{4}}{1000}$$

$$= 48.16 \times 10$$

$$= 481.6 \ldots$$

494. KJ/mol.

Q. calculate the maximum wavelength (in Å) of a photon that can dissociate a $X_2(g)$ molecule into gaseous atom.

given : $\Delta_r H = B.E\,(X-X) = 396\ kJ/mol$

$N_A = 6 \times 10^{23}$, $h = 6.6 \times 10^{-34}\ J.s.$

sol. For dissociation of 1 molecule of X_2 energy required $= \dfrac{396 \times 1000}{6 \times 10^{23}}\ J$

$$E = \frac{hc}{\lambda}$$

$$66 \times 10^{-20} = \frac{6.6 \times 10^{-34} \times 3 \times 10^{8}}{\lambda}$$

$$\lambda = \frac{10^{-26} \times 3}{10^{-19}}$$

$$= 10^{-7} \times 3 \times 10^{10}$$

$$= \boxed{3000\ Å}$$ Ans.

★ FAILURE OF BOHR MODEL :

(1) Bohr model is applicable for only H atom or H-like ions (unielectronic species)

eg:- He^+, Li^{2+}, Be^{3+}, etc.

(2) Could not explain spectrum of multi-electronic species.

(3) Could not explain very fine line spectrum of H-atom. (further splitting of line spectrum of H-atom in presence of electric or magnetic field)

{ Zeeman effect : Magnetic field }
{ Stark effect : electric field }

(4) Was unable to explain why angular momentum of electron is quantized.

(5) Was not according to De Broglie concept and Heizenberg uncertainity principle.

(6) Could not explain 3-dimensional model of H-atom.

(7) Not valid isotopes. (z not whole no.)

∴ $\boxed{\lambda = \frac{h}{mc}}$ radiations.

★ DE-BROGLIE CONCEPT ★

★ concept is related to dual nature of moving particles (like e^-, even car or anything moving)

★ according to De broglie, like electromagnetic radiations electrons / moving particles also have dual nature.

$$\boxed{\lambda = \frac{h}{m.v} = \frac{h}{P}}$$

λ : wavelength associated with e^-/moving particle

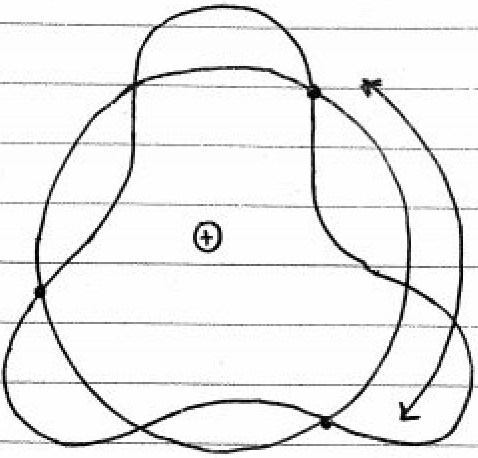
$\boxed{\lambda : \text{Wave Property}}$

m : Mass of moving e^-/particle ~~λ~~ $\boxed{\text{m: Particle property}}$

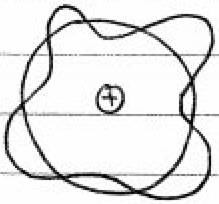
v : velocity of e^-/particle

★ according to him, electron bound with nucleus is like standing wave.

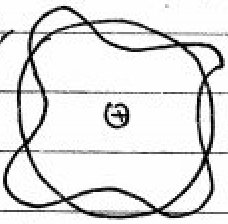
★



Imp:-
No. of waves = n



★



e^- not stable in this type of orbit.

★ CIRCUMFERENCE OF ORBIT :

$\boxed{2\pi r = n\lambda}$ → wavelength associated with e^-

where n = 1, 2, 3, 4 . . . .

★ fractional value is not possible.

★ ★ electrons revolve only on those orbits where circumference of orbit is a simple multiple of λ.

★ $\lambda = \frac{h}{P}$ ∴ $\boxed{\lambda \propto \frac{1}{P}}$

$\boxed{\lambda = \frac{h}{mv}}$ . . . . (i)

and, $\boxed{2\pi r = n\lambda}$ . . . . (ii)

Using (i) and (ii),

$$2\pi r = n\left(\frac{h}{mv}\right)$$

gives ↓ ~~2πr~~

$\boxed{mvr = n\left(\frac{h}{2\pi}\right)}$ ★ m: mass of moving particle.

ϕ ★★★ De Broglie has great significance for microscopic particle.

★ explains Bohr's formula also of angular momentum.

★ Calculation of wavelength associated with moving particle under influence of electric field :-

★ K.E. of moving particle = $\frac{1}{2}mv^2$

$$2 \times KE = mv^2$$

$$2mKE = m^2v^2 = p^2$$

$$p = \sqrt{2mKE}$$

★ $\boxed{p = \sqrt{2mK.E}}$ p = momentum of moving particle.

K.E. of particle = qV

$$\lambda = \frac{h}{p} = \frac{h}{\sqrt{2mKE}}$$

$$\therefore \boxed{\lambda = \frac{h}{\sqrt{2mqV}}}$$

★ ★ V : Potential difference or applied voltage needed.

★ For electron : $q = 1.6 \times 10^{-19}$ C $m = 9.1 \times 10^{-31}$

$$\therefore \boxed{\lambda = \sqrt{\frac{150}{V}}} \text{ in Å}$$

$\lambda \downarrow$

★ Wavelength associated with moving e^-.

Q. calculate De-Broglie wavelength of an electron accelerated by a V of 9.375 V.

Sol.

$$\lambda = \sqrt{\frac{150}{9.375}}$$

$$\boxed{\lambda = 4\text{Å}}$$

Q. calculate the wavelength associated with a 663 gram tennis ball moving with 50 m/s.

Sol.

$$\lambda = \frac{\cancel{f}\,h}{v}$$

$$\lambda = \frac{6.63 \times 10^{-34}}{.663 \times 50}$$

$$= \boxed{2 \times 10^{-3\cancel{5}4}\ \text{m}}$$

Q. calculate the wavelength associated with an electron moving with 66.3 m/s

1. 10^{-5}

Sol.

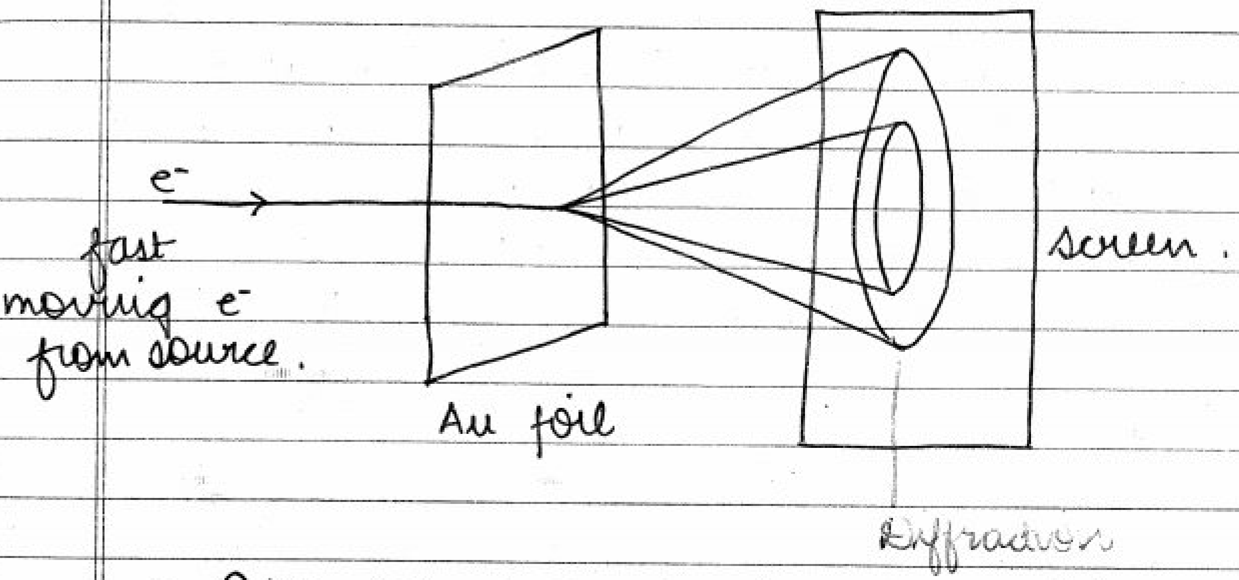
$$\lambda = \frac{6.63 \times 10^{-34}}{66.3 \times 9.1 \times 10^{-31}}$$

$$= \frac{1}{91} \times 10^{-3}$$

★★ WAVE NATURE of e^- confirmed by THOMSON, DAVISSON & GERMER

Thomson used : Au foil

Davisson & Germer used : Ni/Al foil



★ Diffraction pattern of e^- was similar to diffraction of X-rays.

★ Dual nature of electromagnetic radiation

★ Dual nature of matter (e^-)

★ Electron has wave dominating nature because of its small mass.

Heisenberg Uncertainity Principle

Uncertainity means 'change' or 'error' may be +ve or -ve.

★ It is impossible to measure simulataneously both position and momentum of moving particle with certainty.

can measure one thing (Δx or Δp) at one time. $\because$ $\boxed{\Delta x . \Delta p = k(\text{constant})}$

★ $$\boxed{\Delta x \, \Delta p \geq \frac{h}{4\pi}} \quad \text{———} \quad (i)$$

Δx : Uncertainity of position of e^-

Δp : " " momentum of e^-

★ If one uncertainity is small, then other is larger $\because \Delta x \Delta p$ = constant.

★ Holds good for microscopic objects like De broglie concept.

$$\boxed{\Delta p = m\Delta v}$$

$$\Delta x . \Delta m\Delta v \geq \frac{h}{4\pi}$$

$$\boxed{\Delta x . \Delta \underset{v}{\cancel{p}} \geq \frac{h}{4\pi m}}$$ ——— (ii)

Δv : uncertainity in velocity.

★ $$\lambda = \frac{h}{p}$$

$$p = \frac{h}{\lambda}$$

$$\frac{dp}{d\lambda} = \frac{h}{\lambda^2}$$ (no sign consideration)

$$\boxed{\frac{\Delta p}{\Delta \lambda} = \frac{h}{\lambda^2}}$$

from eq.(i),

$$\Delta x \frac{h}{\lambda^2} \Delta \lambda \geq \frac{h}{4\pi}$$

$$\boxed{\Delta x . \Delta \lambda \geq \frac{\lambda^2}{4\pi}}$$

$\Delta \lambda$: uncertainty in wavelength.

Q. Prove that.

$$\boxed{\Delta x\,\Delta E = \frac{hv}{4\pi}}$$

where E = Kinetic energy of particle.

K.E. of particle $= \frac{1}{2}mv^2$

$$2E = mv^2$$

$$2\,dE = 2mv\,dv$$

$$\frac{dE}{dv} = mv$$

$$\boxed{\frac{\Delta E}{\Delta v} = mv}$$

$$\Delta x \frac{\Delta E}{mv} = \frac{h}{4\pi m}$$

$$\boxed{\Delta x . \Delta E = \frac{hv}{4\pi}}$$

ΔE : uncertainity in K.E.

Q. calculate uncertainity in position of an electron moving with $(6.62 \times 10^6 \pm 1\%)$ m/s

8.7 Å

Sol. $\Delta v = 6.62 \times 10^4$ $\frac{1}{100} \times 6.62 \times 10^6$

$\Delta x . \Delta v = \frac{h}{4\pi m}$ $\Delta x =$

Q. If the uncertainity in the position of an electron is $\frac{1}{3}$ PM (picometre) What is ΔV?

Sol. $\frac{1}{3}\times10^{-12}$ m

$$\Delta x = \frac{6.63\times10^{-34}}{4\times3.14\times9.1\times10^{-31}\times\frac{1}{3}\times10^{-12}}$$

$$= 1.75\times10^{8}\ \text{m/s}$$

① $\Delta x\Delta p \geq \frac{h}{4\pi}$ ② $\Delta x\Delta v \geq \frac{h}{4\pi m}$

③ $\Delta x\Delta\lambda \geq \frac{\lambda^2}{4\pi}$ ④ $\Delta x\Delta E \geq \frac{h\nu}{4\pi}$

$\Delta x\Delta p \geq h/4\pi$ $\Delta x\Delta v \geq \frac{h}{4\pi m}$

$\Delta x\Delta\lambda \geq \frac{\lambda^2}{4\pi}$ $\Delta x\Delta E \geq \frac{h\nu}{4\pi}$

D.P.D. ⑧

* (B)
* (D)
* (A)
* (A)
* (B)
* (B)
* (C)
* (A)
* (A, C)
* (A, B, C)
* (C, D)
* (ii), (iv)

Q.B.

$\boxed{1-15}$

Ex II $\longrightarrow \boxed{1-18}$

16 – frequency of revolution of e^-