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Ch-1

Solutions.

Solvent : Water is a good inorganic solvent.
M_{H_2O} = 18g.

Organic solvent ⇒ Ethanol, Acetone, Ether, methanol.

Note : "Like Dissolves Like"

Concentration : The ratio of amount of solute to amount of solution is called concentration

$$C = \frac{\text{amount of solute}}{\text{amount of solution}}$$

Concentration term

1) Mass % [w/w]

⇒ $\frac{\text{Mass of a component in solution}}{\text{Total mass of solution}} \times 100$

2) Volume % [v/v]

⇒ $\frac{\text{Volume of a component}}{\text{Total volume of solution}} \times 100$

3) Mass by volume % [w/v]

⇒ $\frac{\text{Mass of a component}}{\text{Total Volume}} \times 100$

4) Parts per million [ppm]

⇒ $\frac{\text{No. of parts of a component}}{\text{Total no. of parts of solution}} \times 10^6$

n_1 = no. of moles of solvent

n_2 no. of moles of solute

W_1 = Given mass of solvent

W_2 = Given mass of solute

M_1 = molar mass of solvent

M_2 = molar mass of solute

[1 Litre में 55.5 mol of water]

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5) Mole Fraction

→ No. of moles of a component / Total no. of moles

$x_1 + x_2 \ldots x_n = 1$

6) Molarity

→ Moles of solute / Volume of solution [L]

7) Molality

→ moles of solute / Mass of solvent [Kg]

Eg 1 moles of $C_2H_6O_2$ = $\frac{20}{62}$ (20 → gm, 62 → mm) = 0.322 mol

Moles of H_2O = $\frac{80}{18}$ (80 → 40 g, 18 → g) = 4.44 mol

$X_{C_2H_6O_2} = \frac{n_A}{n_A + n_B}$

$= \frac{0.322}{0.322 + 4.44} = \frac{0.322 \times 1000}{4.766 \times 1000}$

0.322
4.444
4.766

0.06
4766) 32200
28596

$= 0.06$

$X_{glycol} = 0.068$

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Relationship b/w molarity & molality.

We know that,

$$M = \frac{n_2}{V_{(L)\,solution}} \quad \text{— (I)} \qquad m = \frac{n_2}{w_1\,(kg)} \quad \text{— (II)}$$

Divide (I) & (II)

$$\frac{M}{m} = \frac{n_2/V}{n_2/w_1}$$

$$\frac{M}{m} = \frac{\frac{1}{V}}{\frac{1}{w_1}}$$

$$\boxed{\frac{M}{m} = \frac{w_1\,(kg)}{V(\ell)}}$$

Note: $\rho = \frac{m}{V}$ → mass of solution.

Relation b/w molality & mole fraction

We know that,

$$m = \frac{n_2}{w_1(kg)} \quad \text{— (I)} \qquad X_2 = \frac{n_2}{n_1 + n_2} \quad \text{— (II)}$$

Divide (I) & (II)

$$\frac{m}{X_2} = \frac{\frac{n_2}{w_1}}{\frac{n_2}{n_1+n_2}}$$

$$\frac{m}{X_2} = \frac{1}{w_1} \times \frac{n_1+n_2}{1}$$

$$\frac{m}{X_2} = \frac{n_1+n_2}{w_1}$$

$$\boxed{\frac{m}{X_2} = \frac{n_1}{w_1} + \frac{n_2}{w_1}}$$

$$\Rightarrow \frac{m}{x_2} = \frac{n_1}{w_1} + m$$

$$\Rightarrow m\left[\frac{1}{x_2} - 1\right] = \frac{n_1}{w_1}$$

Relation b/w molarity & mole fraction :

$$M = \frac{n_2}{V_{sol}} \quad -(I) \qquad x_2 = \frac{n_2}{n_1 + n_2} \quad -(II)$$

Divide (I) & (II)

$$\frac{M}{x_2} = \frac{\frac{n_2}{V}}{\frac{n_2}{n_1 + n_2}}$$

$$\frac{M}{x_2} = \frac{n_1 + n_2}{V}$$

$$\frac{M}{x_2} = \frac{n_1}{V} + \frac{n_2}{V} \rightarrow \text{molarity}$$

$$\boxed{\frac{M}{x_2} = \frac{n}{V} + M}$$

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* **Solubility :**

Solubility of a substance is its max^m amount that can be dissolved in a specified amount of solvent at specific temperature is k/a solubility

* solubility depends upon the amount of solute or solvent.
  - It also depends upon the nature of solute & solvent.

* **Solubility of a solid in a liquid :**

- "like dissolves like"
- when a solid solute is added to the solvent, some solute dissolves & its concentration increases in solution.
- Some solute particles in solution collide with the solid solution particles & get separated out of solution this process is k/a crystallisation
- eg: salt + water → sugar + water

⇒ Saturated solution & unsaturated solution

- such a solution in which no more solute can be dissolved at the same temp. & pressure is k/a saturated
- such a solution in which more solute can be dissolved at the same temp. & pressure is k/a unsaturated

⇒ Effect of temperature

- Solubility ∝ Temperature [For endothermic rxn]
- Solubility ∝ $\frac{1}{T}$ [For exothermic rxn]

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* Effect of pressure (: No effect)

- Pressure does not have significant effect on solubility.
- It is because solids & liquids are highly incompressibility.

* **Solubility of a gas in liquid :**

- Many gases dissolve in water
- The solubility of gases increases with increase in pressure.
- The solubility of the gas will increase until a new equilibrium is reached resulting in an increase in the pressure of a gas above the solution & thus its solubility increases.

HENRY'S LAW

- The law states that the solubility of a gas in liquid is directly proportional to the partial pressure of the gas present above the surface of liquid or solution

OR

- The mole fraction of gas in the solution is proportional to the partial pressure of the gas over the solution.

OR

- The partial pressure of the gas in vapour phase (p) is proportional to the mole fraction of the gas (x) in the solution.

$$p \propto x$$

$$p = K_H x$$

↳ valid for gases

K_H SI unit = Bar

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⇒ **Effect of Temperature.**

Solubility $\propto \frac{1}{\text{Temp.}}$

Input 1.6

Given : $n_{H_2S} = 0.195$ mol

mass of $H_2O = 1000$ g.

[0.195 m means 0.195 moles of H_2S in 1000 g H_2O]

SP = 0.987 bar.

moles of H_2S (n_A) = 0.195 mol

moles of H_2O (n_B) = $\frac{GM}{mm} = \frac{1000}{18}$ = 55.5 mol

$X_{H_2S} = \frac{n_A}{n_A + n_B} = \frac{0.195}{0.195 + 55.5}$

$= \frac{0.195 \times 1000}{55.695 \times 1000} = \frac{195}{55695}$

$X_{H_2S} = 0.0035$

$p = K_H x$

$0.987 = K_H \times 0.0035$

$\frac{987 \times 1000}{0.0035 \times 1000} = K_H$

$\frac{9870}{35} = K_H$

$282 = K_H$

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Vapour pressure of liquid solution :

- Liquid solution are found when solvent is a liquid.

Raoult's law

⇒ Acc. to Raoult's law for solution of volatile liquids, partial pressure of each component is directly proportional to its mole fraction at constant temperature.

$$p_1 \propto x_1$$

$$p_1 = p_1^0 x_1 \quad -(i)$$

Similarly, $p_2 \propto x_2$

$$p_2 = p_2^0 x_2 \quad -(ii)$$

Adding (i) & (ii)

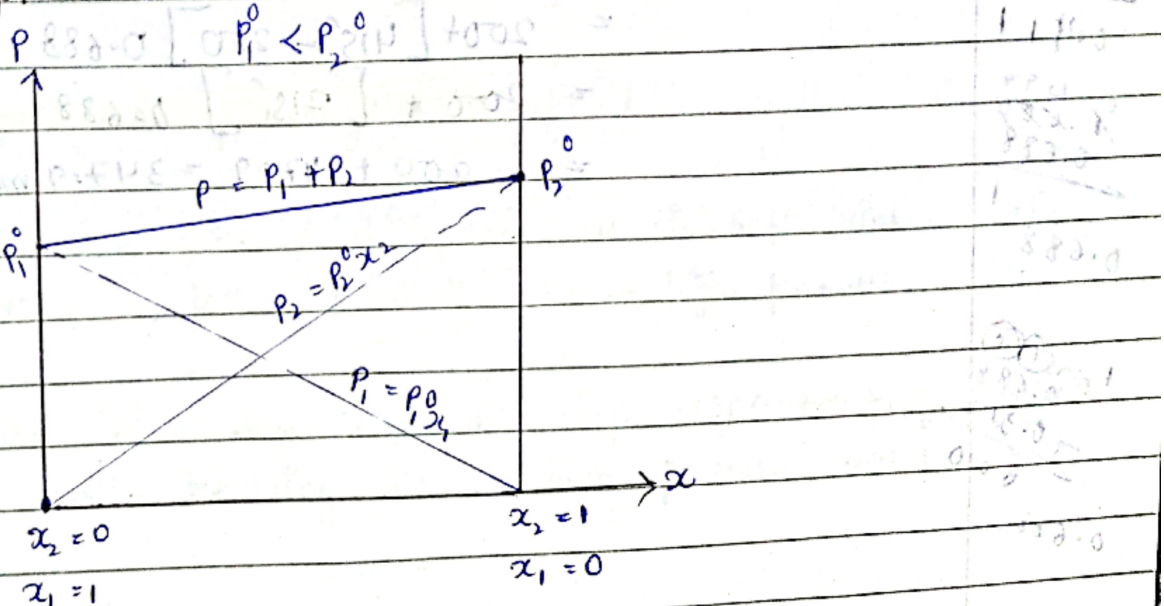
$$p_{total} = p_1 + p_2$$

$$p_{total} = x_1 p_1^0 + x_2 p_2^0$$

$$p_{total} = (1 - x_2) p_1^0 + x_2 p_2^0$$

$$\boxed{p_{total} = p_1^0 + [p_2^0 - p_1^0] x_2}$$

Graph.



(i) Total vapour pressure over the solution can be related to the mole fraction of any one component.

(ii) Total vapour pressure over the solution varies linearly with mole fraction of component 2.

Ans 1.5 $P_1^\circ = 200$ mm Hg $P_2^\circ = 415$ mm Hg.

(i) Moles of $CH_2Cl_2 = \frac{40.8}{85} = 0.47$ mole

Moles of $CHCl_3 = \frac{25.5 \times 10}{119.5 \times 10} = \frac{255}{1195} = 0.213$ mol

Mole Fraction of

$X_{CH_2Cl_2} = \frac{0.47}{0.47 + 0.213} = \frac{0.47}{0.683} = 0.688$

$1 - 0.688 = 0.312$

$P_{total} = P_1^\circ + [P_2^\circ - P_1^\circ] x_2$

$= 200 + [415 - 200] 0.688$

$= 200 + [215] 0.688.$

$= 200 + 147.9 = 347.9$ mm Hg.

(ii) $P_1 = y_1 P_{total}$ $P_2 = y_2 P_{total}$

$P_1 = P_1^0 x_1$

* **Raoult's law is a special case of Henry's law:**

⇒ If we compare the eqn for Raoult's law & Henry's law, the partial pressure of the volatile components of gas is directly proportional to its mole fraction. Only the proportionality constant K_H differs from P_1^0

$P = K_H x \rightarrow$ Henry

$P_1 = P_1^0 x \rightarrow$ Raoult's

* **Applications of Henry's law:**

- To increase the solubility of CO_2 in soft drink & soda water, the bottle is sealed under high pressure.
- Scuba divers must cope with high concentration of dissolved gas while breathing air at high pressure under water.

(iii) At high altitudes the partial pressure of oxygen is less than that at the ground level. This leads to low concentrations of oxygen in the blood and tissues of people living at high altitudes or climbers.

Q The vapour pressure of water is 12.3 kPa at 300 K. Calculate vapour pressure of 1 molal solution of a non-volatile solute in it.

Ans $p_1^0 = 12.3\,kPa = 12.3 \times 10^3\,Pa$ $\quad m = 1\ mol/kg$.

$$x_{solute} = \frac{n_A}{n_A + n_B} = \frac{55.5}{1 + 55.5}$$

$$x_{solute} = \frac{55.5}{56.5}$$

$$p_1 = p_1^0 x_1$$

$$p = 12.3 \times 10^3 \times \frac{55.5}{56.5}$$

$$p = 0.982 \times 12.3 \times 10^3$$

$$\boxed{p = 12.07\ kPa \times 10^3\ Pa}$$

* **Ideal solution :**

* The solutions which obey Raoult's law over the entire range of concentration is like ideal solution.

where, $\Delta_{mix} H = 0$ $\quad \Delta_{mix} V = 0$

eg : n-Hexane & n-heptane, Benzene & toluene, Chloroethane & Bromoethane.

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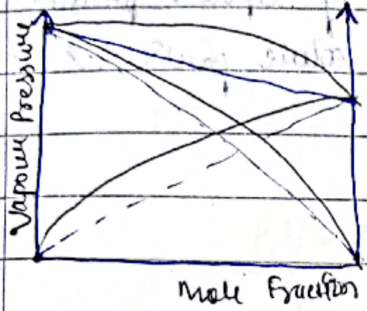
* Non-Ideal solution

→ The solution which doesn't obey Raoult's law over the entire range of concentration.

where, $\Delta_{mix} H \neq 0$ $\Delta_{mix} V \neq 0$

→ Positive deviation

Graph :



→ $\Delta H_{mix} > 0$

$\Delta V_{mix} > 0$

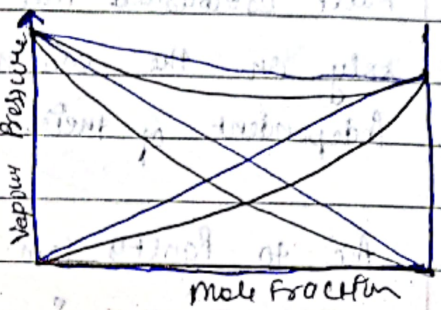
→ eg : ethanol & acetone

carbon-disulphide & acetone

→ (A-A) + (B-B) > (A-B)

→ Negative deviation

Graph :



→ $\Delta H_{mix} < 0$

$\Delta H_{mix} < 0$

→ eg : Phenol & aniline

chloroform & acetone

→ (A-A) + (B-B) < (A-B)

* Azeotropes :

→ Azeotropes are the binary mixture of two or more liquids having same composition in liquid & vapour phase & boils at a constant temp.

Minimum boiling azeotropes

→ The solutions which show a large +ve deviation

→ eg : ethanol + water

maximum boiling azeotropes.

→ The solutions which show a large -ve deviation

→ eg : nitric acid + water

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Imp

Colligative properties [4 marks]

⇒ Colligative properties are those property of solute which depends on the no. of solute & not on the nature of solute.

① **Relative lowering of vapour pressure :**

⇒ Raoult established that the lowering of vapour pressure depends only on the concentration of the solute particles & it is independent of their identity.

Acc. to Raoult's law.

$$p_1 = x_1 p^\circ \quad -①$$

The reduction in the vapour pressure of solvent.

$$\Delta p_1 = p_1^\circ - p_1 = p_1^\circ - p_1^\circ x_1$$
$$= p_1^\circ [1 - x_1]$$

$$\Delta p_1 = p_1^\circ (1 - x_1) \quad -②$$

Now, $\Delta p_1 = x_2 p_1^\circ$

$$\boxed{x_2 = \frac{\Delta p_1}{p_1^\circ}} \qquad [x_1 + x_2 = 1]$$

$$\frac{\Delta p_1}{p_1^\circ} = \frac{p_1^\circ - p_1}{p_1^\circ}$$

$$x_2 = \frac{p_1^\circ - p_1}{p_1^\circ}$$

$$\boxed{\frac{n_2}{n_1 + n_2} = \frac{p_1^\circ - p_1}{p_1^\circ}}$$

If $n_2 \gg n_1$

n_2 is negligible in denominator.

$$\therefore \quad \frac{p_1^0 - p_1}{p_1^0} = \frac{n_2}{n_1}$$

or

$$\frac{p_1^0 - p_1}{p_1^0} = \frac{w_2 \times M_1}{M_2 \times w_1}$$

Q $w_1 = 39$ $w_2 = 0.5$ $M_1 = 78$, $M_2 = ?$

$p_1^0 = 0.850$ bar. $p = 0.845$ bar.

$$\frac{p_1^0 - p_1}{p_1^0} = \frac{w_2 \times M_1}{M_2 \times w_1}$$

$$\frac{0.850 - 0.845}{0.850} = \frac{0.5 \times 78}{M_2 \times 39}$$

$$\frac{0.005}{0.850} = \frac{390}{M_2 \times 39}$$

$$\frac{0.005}{0.850} = \frac{10}{M_2}$$

$$M_2 = \frac{10 \times 0.850 \times 1000}{0.005 \times 1000}$$

$$M_2 = \frac{10 \times 850}{5}$$

$$M_2 = 170 \text{ g/mol}$$

② Elevation of boiling point.

* Elevation of boiling point is directly proportional to molality.

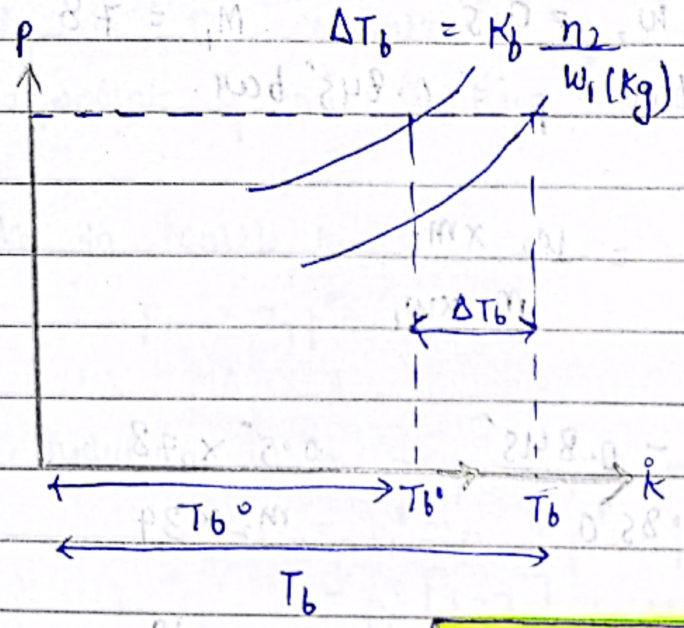
$$\Delta T_b \propto m$$

$$\Delta T_b = K_b m$$

K_b = Boiling point elevation constant.

↳ SI unit = K kg / mol

$$\Delta T_b = K_b \frac{n_2}{w_1 (kg)}$$



$$\Delta T_b = T_b - T_b°$$

$$\Delta T_b = \frac{K_b \times 1000 \times w_2}{m_2 \times w_1}$$

$$m_2 = \frac{1000 \times w_2 \times K_b}{\Delta T_b \times w_1}$$

water: $T_b^\circ = 373k$ [100°C] → boiling

$T_f^\circ = 273k$ [0°C] → Freezing

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③ Depression of Freezing point.

$$\Delta T_f = T_f^\circ - T_f$$

(T_f° → Freezing point of pure solvent; T_f → non-volatile solute.)

$\Delta T_f \propto m$

$$\Delta T_f = K_f m$$

$$m_2 = \frac{K_f \times w_2 \times 1000}{\Delta T_f \times w_1}$$

Eg 1.7 $K_b = 0.52\ K\ Kg/mol$ · $w_2 = 18g$ $w_1 = 1000g$.

$$m = \frac{\text{moles}}{\text{Mass}} = \frac{0.1}{10} = 0.1\ mol/kg.$$

moles of glucose $= \frac{18}{180} = \frac{1}{10} = 0.1\ mol$

$$\Delta T_b = K_b \times m = \frac{0.52}{100} \times \frac{0.1}{10} = \frac{52}{1000}$$

$$\Delta T_b = 0.052\ K$$

373.00
0.052
373.052

water's boiling point = 373k K

∴ boiling point of solution = 373 + 0.052

= 373.052 K

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Eg 1.8

$w_2 = 1.80\ g$ $\quad w_1 = 90$ $\quad \Delta T_b = 354.11\ K$

$M_2 = ?$ $\quad K_b = 2.53\ K\ kg/mol$

$$M_2 = \frac{1000 \times w_2 \times K_b}{\Delta T_b \times w_1}$$

$$M_2 = \frac{1000 \times 1.80 \times 2.53}{354.11 \times 90 \times 100}$$

$$M_2 = \frac{1.80 \times 25.3}{354.11 \times 9}$$

$$M_2 = 58\ g/mol$$

Osmotic Pressure (π)

a) π is proportional to molarity (C) of the solution at given temperature. (C → concentration)

$$\pi = CRT$$

$$\pi = \frac{n_2}{V} RT$$

$$\pi = \frac{W_2 RT}{M_2 V}$$

$$M_2 = \frac{W RT}{\pi V}$$

Osmosis

a) The flow of solvent from its high concentration region to low concentration through a semi-permeable membrane is kla Osmosis.

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* Osmotic pressure

⇒ The excess pressure which must be applied to a solution to prevent the passage of solvent into it through a spm is k/a osmotic pressure.

* 3 types of solution

Hypertonic	Hypotonic	Isotonic
⇒ The solution having osmotic pressure higher than that of another solution is k/a Hypertonic.	⇒ A solution having osmotic pressure lower than that of another solution is k/a Hypotonic.	⇒ The solution having same osmotic pressure at a given temp. are k/a Isotonic solution.

Eg 1.11 $V = 200\ cm^3 = 0.200\ l$ $\quad w_2 = 1.26\ g$

$T = 300\ k$ $\quad \pi = 2.57 \times 10^{-3}$

$R = 0.083$

$$M_2 = \frac{w_2 RT}{\pi V} = \frac{1.26 \times 0.083 \times 300 \times 1000 \times 100}{2.57 \times 10^{-3} \times 0.200 \times 1000 \times 100}$$

$$= \frac{126 \times 83 \times 300}{2.57 \times 10^{-3} \times 200} = \frac{1045.8 \times 10^3 \times 3}{2.57 \times 2}$$

$$= \frac{9437 \times 10^3}{514} = 61,022\ g/mol$$

$$207$$

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Application of osmosis

- A raw mango placed in concentrated salt solution loses water via osmosis & shrivel into pickle.
- Wilted flowers revive when placed in fresh water.
- When placed in water containing less than 0.9% salt, blood cells swell due to flow of water in them by osmosis.
- People taking a lot of salty food experience water retention in tissue cells & intercellular spaces because of osmosis. The resulting puffiness or swelling is called edema.

Reverse osmosis.

⇒ The direction of osmosis can be reversed if a pressure larger than the osmotic pressure is applied to the solution side. That is, now the pure solvent flows out of the solution through the semi-permeable membrane. This is k/a reverse osmosis.

Abnormal Molar Mass.

⇒ Such a molar mass that is either lower or higher than the expected or normal value is called as Abnormal molar mass

Van't Hoff factor.

⇒ The ratio of normal molar mass to the abnormal molar mass.

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i = Normal molar mass / Abnormal molar mass = Observed colligative property / Calculated colligative property

i = Total no. of moles of particles after association / dissociation / no. of moles of particles before association / dissociation.

⇒ Inclusion of van't Hoff factor modifies the eqn of colligative property.

① RLVP :

$$\frac{p_1^\circ - p_1}{p_1^\circ} = i \cdot \frac{n_2}{n_1}$$

② Elevation of Boiling point :

$$\Delta T_b = i K_b m$$

③ Depression of Boiling point :

$$\Delta T_f = i K_f m$$

④ Osmotic pressure

$$\pi = i \frac{n_2 RT}{V}$$

freezing point depression constant of benzene is 5.12 K kg/mol find the molar mass

$$M_2 = \frac{1000 \times K_f \times w_2}{w_1 \times \Delta T_f}$$

$$= \frac{1000 \times 5.12 \times 1 \times 100}{50 \times 0.40 \times 100}$$

$$= \frac{100 \times 5.12}{20} \quad 25°C$$

$$M_2 = 256 \text{ g/mol}$$

* **Degree of Dissociation (α)**

⇒ It is defined as the fraction of the total substance that undergoes dissociation is called DOD.

$$\frac{\text{No. of moles dissociates}}{\text{Total no. of moles taken}} = \alpha$$

$$\alpha = \frac{m_c - m_o}{m_o (n-1)}$$

where, m_c = calculated molar mass

m_o = observed molecular mass.

$$\alpha_a = \left(\frac{m_c - m_o}{m_o}\right)\left(\frac{n}{n-1}\right)$$

$$\alpha_a = n\alpha_d$$

$$n = \frac{\alpha_a}{\alpha_d}$$

* **Degree of Association (α)**

⇒ It is defined as the fraction of the total substance that undergoes association is called DOA.

$$\frac{\text{No. of moles Associate}}{\text{Total moles taken}} = \alpha$$