

SOLUTION

* Solutions :→

It is homogeneous mixture of two or more substances having same phase and uniform properties.

"Single phase containing two or more components"

→ It is the combination of solute and solvent :→

i) Solute :→

It is the component which is present in lesser amount.

ii) Solvent :→

It is the component which has the same phase as that of solution and is present in larger amount.

* Types of Solution :→

i) Binary Solution :→

It is made up of two components.

Ex:- Sugar solⁿ → Sugar + water

ii) Ternary Solution :→

It is made up of three components.

Ex:- Sugar + Salt + Water

iii) Quaternary Solution :→

It is made up of four components. Ex:- Sugar + Salt + Water + Lemon.

There are three types of solutions according to physical state:

- Gaseous solution
- liquid solution
- solid solution

i) Gaseous Solution: \Rightarrow

The combination of gas with gas, is known as the gaseous solution. All the gaseous mixtures are known as the Homogeneous mixture.

Example:-

Solvent	Solute	Examples
Gas	Gas	Air, $O_2 + H_2 + N_2$
Gas	Liquid	Humidity
Gas	Solid	NH_4Cl , Camphor

ii) Liquid Solution: \Rightarrow

The combination of liquid with liquid, gas or solid is known as liquid soln.

Example:-

Solute	Solvent	Examples
Solid	Liquid	salt + H_2O ; sugar
Liquid	Liquid	Alcohol + H_2O
Gas	Liquid	soda water, cold drink

iii) Solid Solution: \Rightarrow

The combination of solid with gas, liquid or solid is called solid soln.

solid	solid	Alloy $\rightarrow Cu + Zn = Brass$
liquid	solid	Amalgam $\rightarrow Hg + Cu$
gas	solid	$H_2 + Pd$

* Concentration Terms:

The concentration terms are used to express the amount of solute and solvent present in the solution.

→ The amount of solute dissolve in per litre volume of solution or in per kg weight of solvent is called concentration.

Concentration Terms

Temperature dependent (volume)

- Volume percentage
- Mass-volume percentage
- Molarity
- Normality
- PPM v/v
- PPM w/v

Temperature independent (Mass)

- Mass percentage
- Mole fraction
- Molality
- PPM w/w

* Temperature Dependent Concentration terms:

* Volume percentage :

The mass of solute in 100 cm^3 or 100 ml of the solⁿ is known as the volume percentage.

$$\text{Volume \% of A} = \frac{\text{volume of solute (ml)}}{\text{volume of solution (ml)}}$$

If any numerical in the volume of solⁿ is unknown then

Let the vol. of solⁿ = 100 ml .

Solution \rightarrow AB
Solvent - A
Solute - B

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* Mass-volume percentage:

The number of parts by mass of solute per 100 parts by the volume of solution.

\rightarrow It is denoted by $w/v\%$:

$$w/v\% = \frac{\text{Weight of Solute (g.m)}}{\text{Volume of Solution (ml)}} \times 100$$

* Molarity:

The no. of moles of the solute dissolves in the per litre volume of the solution.

\rightarrow It is denoted by M:

\rightarrow S.I unit :- mole/litre.

$$\text{Molarity} = \frac{\text{Moles of Solute}}{\text{Volume of Soln. (litre)}}$$

$$M = \frac{n_B}{V_L} \quad (i)$$

We know that

$$\text{no. of moles} = \frac{\text{Weight}}{\text{Molar Mass}}$$

$$n_B = \frac{W_B}{M_B} \quad (ii)$$

By the eq (i) & (ii)

$$\boxed{M = \frac{W_B}{M_B \times V_L}}$$

Molarity changes with temperature because it depends on the volume.

* Strength: \Rightarrow

The mass of solute in per litre volume of the solution is called strength of soln.

$$\text{Strength} = \frac{\text{Mass of solute (gm)}}{\text{Vol. of soln (litre)}}$$

$$S = \frac{W_B}{V_L} \quad \text{(i)}$$

We know that

$$M = \frac{W_B}{M_B \times V_L} \quad \text{(ii)}$$

By the equation (i) & (ii)

$$M = \frac{S}{M_B}$$

$$S = M \times M_B$$

where

S = Strength of soln.

M = Molarity of soln

M_B = Molar Mass of Solute

* Normality: \Rightarrow

It is the no. of gm equivalent of the solute dissolved in per liter volume of solution.

\rightarrow If is denoted by N .

\rightarrow S.T. Unit := gm eq. / litre.

$$\text{Normality} = \frac{\text{No. of gm. equivalent of Solute}}{\text{volume of soln (litre)}}$$

$$N = \frac{E_B}{V_L} - (i)$$

No. of gm equivalent = $\frac{\text{Weight of Solute}}{\text{Equivalent Mass of Solute}}$

$$E_B = \frac{W_B}{E_m \times W_f} - (ii)$$

By the equation (i) & (ii)

$$N = \frac{W_B}{E_m \times V_L} - (iii)$$

Equivalent Mass = Molar Mass of Solute
 n factor

$$E_m = \frac{M_B}{n_f} - (iv)$$

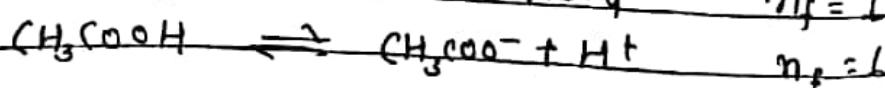
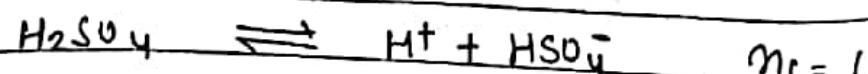
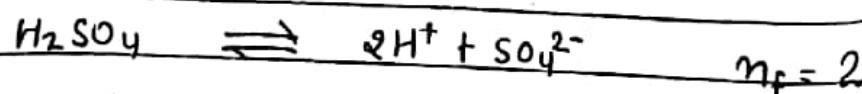
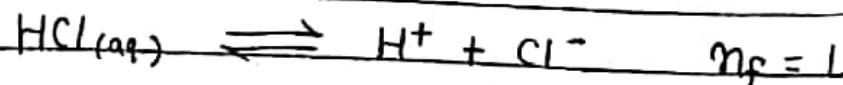
By the equation (iii) & (iv)

$$\boxed{N = \frac{W_B \times n_f}{M_B \times V_L}}$$

Calculation of n -factor: \Rightarrow

i) In the Case of Acids: \Rightarrow

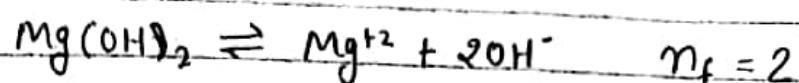
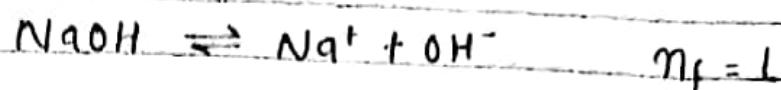
given by acids in aqueous soln. $\quad \text{No. of H}^+$ ions



ii) In the case of Bases :-

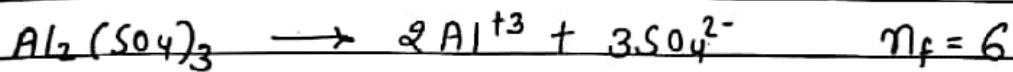
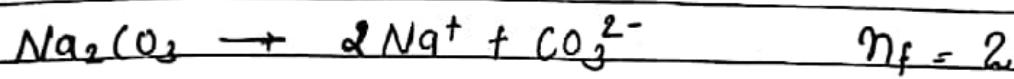
by the Base.

no. of OH^- ion given



iii) In the case of Salts :-

Total magnitude of either positive or negative charge of formula.



* Temperature Independent Concentration terms :-

* Mass percentage :-

The mass of the Solute, per 100 gm. of solution is called Mass percentage.

+ It is represented by w/w %.

$$\% \text{ w/w} = \frac{\text{Weight of Solute (gm)}}{\text{Weight of soln. (gm)}} \times 100$$

If the weight of soln is unknown then

Let the weight of soln = 100 gm.

* Mole Fraction: →

The Ratio of the no. of moles of one component to the total no. of moles present in a Solution.

→ It is denoted by X .

→ It is Unitless quantity.

solvent ↓ n_A	A + B solution	solute b
		n_B

$$\text{no. of moles of solvent} = n_A$$

$$\text{no. of moles of Solute} = n_B$$

$$\text{Total mole} = n_A + n_B$$

Mole fraction of solvent :-

$$X_A = \frac{n_A}{n_A + n_B} \quad (i)$$

Mole fraction of Solute :-

$$X_B = \frac{n_B}{n_A + n_B} \quad (ii)$$

by the adding eq (i) & (ii)

$$X_A + X_B = \frac{n_A}{n_A + n_B} + \frac{n_B}{n_A + n_B}$$

$$x_A + x_B = \frac{m_A + m_B}{m_A + m_B}$$

$$x_A + x_B = 1$$

* Molality: \Rightarrow

The number of moles of the solute present in per kg. weight of solvent.
 → It is denoted by m .
 → S.I. Unit = moles/kg.

$$\text{Molality} = \frac{\text{Moles of Solute}}{\text{Weight of Solvent (kg)}}$$

$$m = \frac{n_B}{W_A} \quad (i)$$

We know that

$$n_B = \frac{W_B}{M_B} \quad (ii)$$

By the eq (i) & (ii)

$$m = \frac{W_B}{M_B \times W_A}$$

If weight of solvent is given in gm.

$$m = \frac{W_B \times 1000}{M_B \times W_A}$$

Molality does not change with the temperature because it depends on mass.

* Relation between molarity and Normality : \Rightarrow

We know that

$$N = \frac{W_B \times n_f}{M_B \times V_L} \quad (i)$$

$$M = \frac{W_B}{M_B \times V_L} \quad (ii)$$

By the eq (i) & (ii)

$$[N = M \times n_f] \quad (iii)$$

Now

$$N = \frac{e_g}{V_L} \quad (iv)$$

$$M = \frac{n_B}{V_L} \quad (v)$$

By the eq (iii), (iv), (v)

$$\frac{e_g}{V_L} = \frac{n_B \times n_f}{V_L}$$

$$e_g = n_B \times n_f$$

* Mole Concept:-

→ It is used to determine the amount of substance.

MOLE		
Quantity	Mass	Volume
1 mole = N_A $N_A = 6.022 \times 10^{23}$	1 mole = Atomic mass	Apply in gas only STP + 273 K / 1 atm
$n = \frac{\text{Quantity}}{N_A} = \frac{6.023 \times 10^{23}}{}$	$n = \frac{\text{Weight}}{\text{Molar Mass}}$	1 mole = 22.4 Ltr. $n = \frac{\text{Vol. of gas}}{22.4}$

$$n = \frac{Q}{N_A} = \frac{W}{M} = \frac{V_{\text{gas}}}{22.4}$$

* Solubility:-

It is a chemical property referring to the ability for a given substance, the solute, to dissolve in a solvent.

→ There are three types of Solubility:-

- i) Solid in liquid.
- ii) Gas in liquid
- iii) Liquid in liquid.

* Solubility of Solids in liquid:

The maximum amount of the solid that can be dissolved in 100 g m liquid solvent at certain temperature is called solubility of solid.

→ The obtained solution is known as the Saturated Solution.

* Factors affecting solubility of Solids:

i) Nature of Solid Solute and Solvent:

→ Ionic and polar solute dissolves in the water.

→ Non-polar solute dissolves in the organic solvent like Benzene, CCl_4 etc.

ii) Temperature:

+ when $\Delta H > 0$, it indicates that the process is endothermic and solubility is increases with increase in temperature.

+ when $\Delta H < 0$, it indicates that the process is exothermic and solubility decreases with increase in temperature.

iii) when hydrated salt is taken as solute:

→ The solubility increases with increase in temperature but after the loss of water of crystallization.

e.g. - Epsom salt ($\text{Na}_2\text{SO}_4 \cdot 10\text{H}_2\text{O}$)

* There is no any effect of the pressure on the solubility of the solid.

* Solubility of gases in liquids: \Rightarrow

\rightarrow gas and liquid both are combine to form a true solution

\rightarrow The volume (ml) of a gas which dissolved in 1 ml water is called absorption coefficient.

* Factors affecting solubility of gases: \Rightarrow

i) Nature of gas: \Rightarrow

The gas which is easily ionised in solution, are soluble in water.

e.g. NH_3 , HCl , SO_2 , etc.

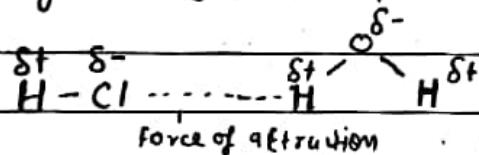
$\rightarrow \text{N}_2$, O_2 and CO_2 are soluble in ethyl Alcohol.

$\rightarrow \text{O}_2$ is readily soluble in blood.

ii) Nature of Solvent: \Rightarrow

\rightarrow Polar gases are soluble in polar solvent.

e.g. HCl gas is soluble in water.



\rightarrow Non-polar gases are soluble in non-polar solvent.

\rightarrow "Like dissolves like".

iii) Temperature: \Rightarrow

Solubility causes release of heat

which indicates that the process is exothermic.

Due to which solubility decreases with increase in temperature.

Exceptions:

H_2 and He dissolves in liquid with the absorption of heat. So, the solubility of these gases increases with increase in temperature.

iv) Effect of Pressure:**Henry's Law:**

The mass of a gas dissolved in a given volume of the liquid, at constant temperature is directly proportional to the pressure of the gas.

$$m \propto P$$

$$m = K_H \cdot P \quad \text{where } K_H = \text{Henry's const.}$$

The partial pressure applied by any gas on liquid surface is directly proportional to its mole fraction present in liquid solvent.

$$P \propto \text{solubility}$$

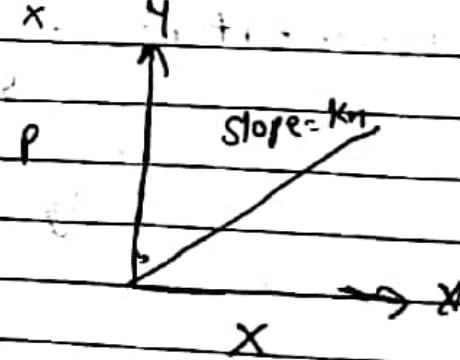
$$X \propto P$$

$$X = K_H \cdot P \quad \text{where } K_H = \text{Henry const.}$$

$$\boxed{P = \frac{X}{K_H}}$$

$$P = K_H \cdot X$$

$$Y \propto M \cdot X$$



at Constant Pressure

$$X \propto \frac{L}{K_H}$$

K_H - P_{es}, X - J_{es}, Solubility - J_{es}.

* Limitations of Henry law: →

- + Pressure should not be too high.
- + Temperature should not be too low.
- + The gas should not be chemically react with liquid.
- The solubility of a gas is lesser in solvent means the solution is dilute.

* Solubility of a Liquid in a Liquid: →

Vapour Pressure: →

The partial pressure exerted by any vapour or gas on the surface of any liquid or solution which is volatile in nature at equilibrium.

* Factors affecting Vapour pressure: →

i) Temperature: →

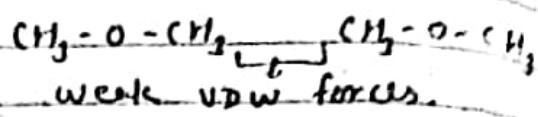
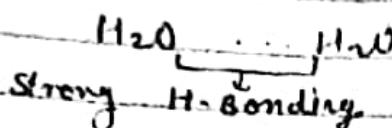
As the temperature increases, rate of vapour formation also increases so, greater will be the vapours, greater will be its pressure.

T:- P_{es} ; V.P.:- P_{es}.

iii) Nature of liquids:

The liquids having strong attraction forces between its molecules will have lower vapour pressure as compare to other liquids.

Cat:-



iii) Boiling point:-

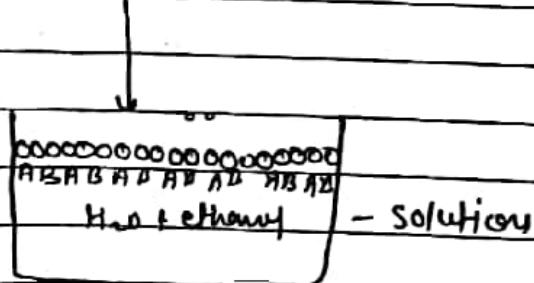
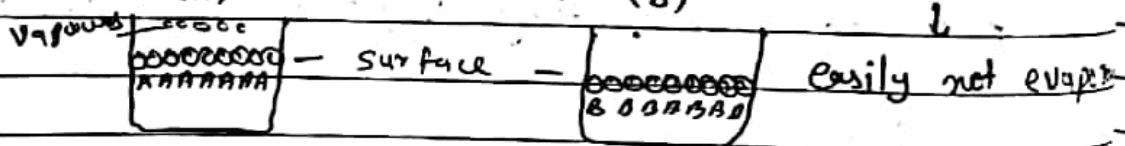
The molecules having lower boiling points can easily convert into vapours due to which they exert high vapour pressure.

low B.P. \rightarrow high Vapour pressure.

iv) Surface area of liquid:-

\rightarrow Vapour pressure of pure liquid, will also be greater than the vapour pressure of its soln. with a non-volatile solute.

(Volatile) H_2O + ethanol (Non-volatile)



V.P. of $\text{H}_2\text{O} >$ V.P. of solution.

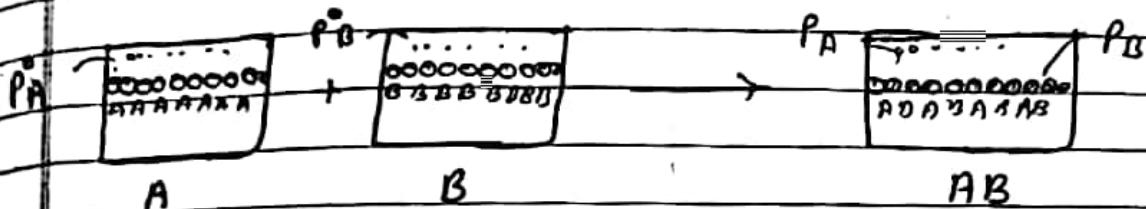
* Raoult's Law: →

for any volatile substance, the partial pressure exerted by vapour or gas on the surface of liquid or solution is directly proportional to its mole fraction in liquid or solution.

CASE-I

Raoult's Law for soln of volatile solute

Solution = Volatile solute + Volatile solvent.



If $P_A > P_B^°$

Let the two volatile components A and B forms a soln. P_A and P_B are the vapour pressure in solution respectively.

liquid 'A'

liquid 'B'

Pure (v.p)

$P_A^°$

$P_B^°$

moles

n_A

n_B

Mole fraction

x_A

x_B

According to Raoult's law

Vapour pressure of A

$$P_A \propto X_A$$

$$P_A = P_A^{\circ} \cdot X_A \quad \text{--- (i)}$$

Vapour pressure of B

$$P_B \propto X_B$$

$$P_B = P_B^{\circ} \cdot X_B \quad \text{--- (ii)}$$

Now, By the Dalton's law of Partial Pressure.

Vapour pressure of solution

$$P = P_A + P_B \quad \text{--- (iii)}$$

By the eq. (i), (ii) & (iii)

$$P = P_A + P_B$$

$$P = P_A^{\circ} X_A + P_B^{\circ} X_B \quad \text{--- (iv)}$$

We know that

$$X_A + X_B = 1$$

$$X_A = 1 - X_B \quad \text{--- (v)}$$

$$X_B = 1 - X_A \quad \text{--- (vi)}$$

Now, By the equation (iv) & (v)

$$P = P_A^o x_A + P_B^o x_B$$

$$P = P_A^o (1-x_B) + P_B^o x_B$$

$$P = P_A^o - P_A^o x_B + P_B^o x_B$$

$$\boxed{P = P_A^o + x_B (P_B^o - P_A^o)}$$

By the equation (iv) & (vi)

$$P = P_A^o x_A + P_B^o x_B$$

$$P = P_A^o x_A + P_B^o (1-x_A)$$

$$P = P_A^o x_A + P_B^o - P_B^o x_A$$

$$\boxed{P = P_B^o + x_A (P_A^o - P_B^o)}$$

Where;

P = Vapour pressure of soln

P_A^o = Vapour pressure of pure liquid A

P_B^o = Vapour pressure of pure liquid B

P_A = Vapour pressure of A in soln.

P_B = Vapour pressure of B in soln.

x_A = Mole fraction of liquid 'A'

x_B = Mole fraction of liquid 'B'.

Graph:

i) $P_A = P_A^o x_A$ and $P_B = P_B^o x_B$

ii) $P_A^o > P_B^o$ then $P = P_B^o + \kappa_A (P_A^o - P_B^o)$

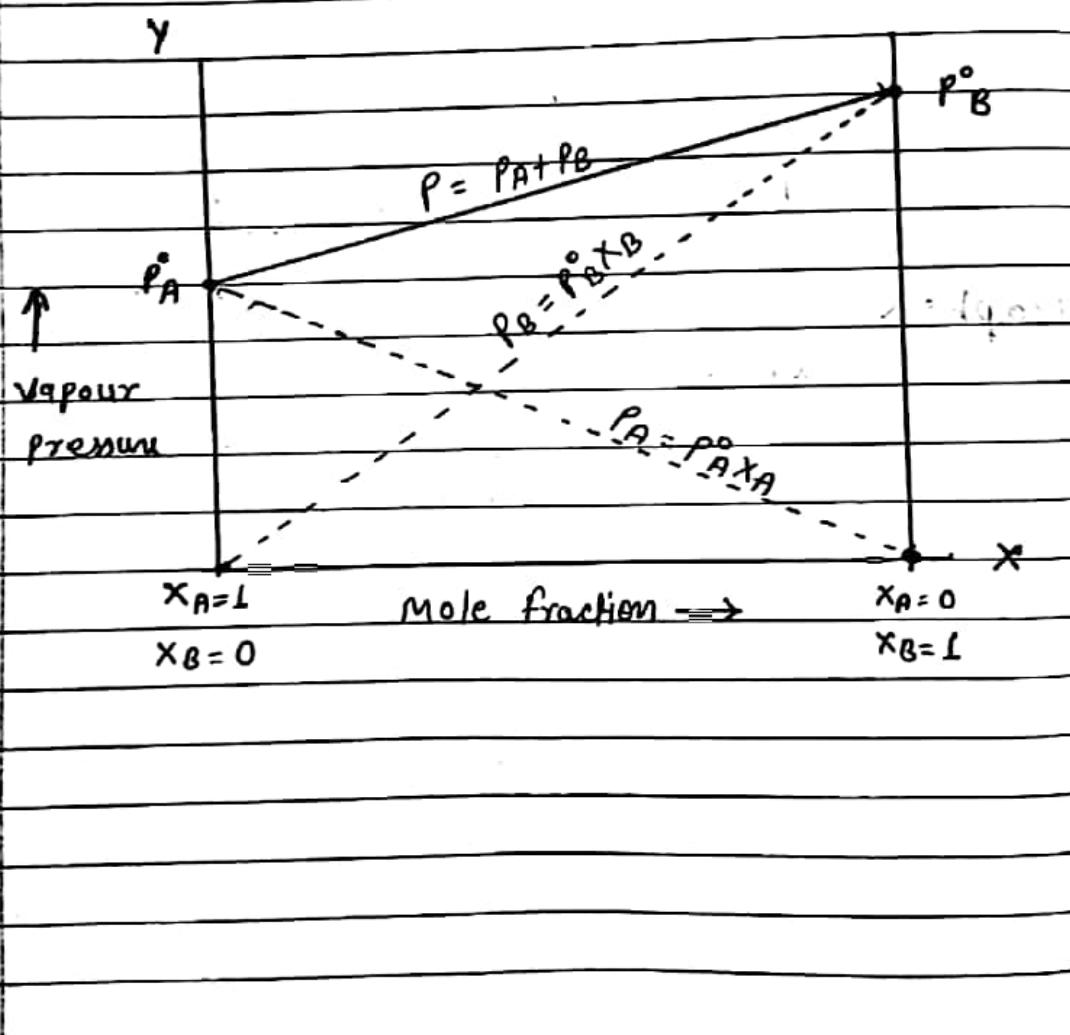
iii) $P_A^o < P_B^o$ then $P = P_A^o + x_B (P_B^o - P_A^o)$

$$P_A = P_A^o x_A \quad \text{and} \quad P_B = P_B^o x_B$$

$$y = mx$$

$$P = P_A^o + x_B (P_B^o - P_A^o)$$

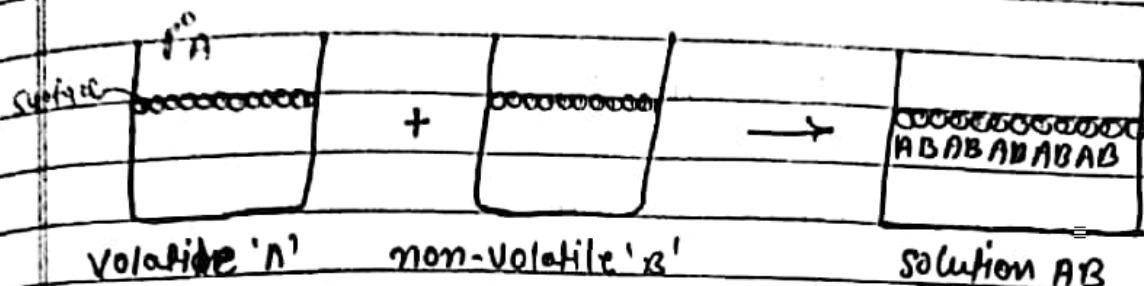
$$y = C + x M \quad [\because y = mx + c]$$



CASE-II

Raoult's Law for the solution of non-volatile solute(s).

solution = volatile solvent + non-volatile solute



The vapour pressure of solution of non-volatile solute is equal to the partial pressure of pure solvent 'A'. Because the partial pressure of non-volatile solute B is negligible.

We know that

$$P_A^{\circ} > P_A$$

According to Raoult's law

$$P_A \propto X_A$$

$$P_A = P_A^{\circ} X_A \quad \text{---(i)}$$

$$\text{Vapour Pressure of Sol}^n \quad P = P_A \quad \text{---(ii)}$$

By the equation (i) & (ii)

$$P = P_A^{\circ} X_A \quad \text{---(iii)}$$

We know that

$$X_A + X_B = 1$$

$$X_A = 1 - X_B \quad \text{---(iv)}$$

By the equation (iii) & (iv)

$$P = P_A^o X_A$$

$$P = P_A^o (1 - X_B)$$

$$P = P_A^o - P_A^o X_B$$

$$P_A^o X_B = P_A^o - P$$

$$\boxed{X_B = \frac{P_A^o - P}{P_A^o}}$$

→ $P_A^o - P$ is called lowering in vapour pressure.

→ $\frac{P_A^o - P}{P_A^o}$ is called Relative lowering in vapour.

→ The relative lowering of vapour pressure is equal to mole fraction of solute in solution having non-volatile solute at a particular temperature.

* There are two types of solutions on the basis of Raoult's law:-

i) Ideal solution.

ii) Non-Ideal solution.

* Ideal Solutions:-

→ The solution which obey Raoult's law at the entire range of concentration and temperature is called ideal solution.

According to Raoult's law

$$P_A = P_A^0 X_A \quad \text{---(i)}$$

$$P_B = P_B^0 X_B \quad \text{---(ii)}$$

By the Dalton's law of Partial Pressure.

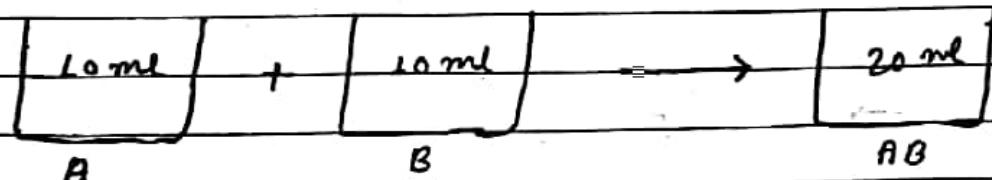
$$P = P_A + P_B$$

$$P = P_A^0 X_A + P_B^0 X_B$$

$$\rightarrow P_m = P_{\text{obs.}}$$

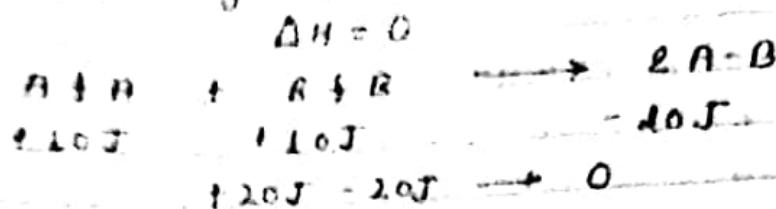
→ change in volume is equal to zero.

$$\Delta V = 0$$



$$\Delta V = V_{\text{soln}} - (V_{\text{solute}} + V_{\text{solvent}})$$

→ The change in enthalpy is zero.



→ The magnitude of attractive force between is equal to the sum of attractive forces present in A-A and B-B in the solution.

→ Liquids having same functional group form ideal solution.

Examples: →

→ n-Hexane + n-Heptane

→ Benzene + Toluene

→ Methanol + Ethanol

→ Ethyl chloride + Ethyl Bromide

* Non-ideal Solutions: →

→ The solutions which do not obey Raoult's law are called non-ideal solutions.

→ $P_A \neq P_A^0 X_A$ and $P_B \neq P_B^0 X_B$

→ $P_A \neq P_A + P_B$

→ $P_m \neq P_{obs}$

→ $\Delta V \neq 0$

→ $\Delta H \neq 0$

⇒ There are two types of non-ideal solutions which are as follows:-

i) Non-ideal solutions with Positive deviation \Rightarrow

\rightarrow A-B interactions are weaker than A-A and B-B interactions.

$\rightarrow P_A > P_A^\circ X_A$ and $P_B > P_B^\circ X_B$

$\rightarrow P_{\text{obs}} > P_m$

$\rightarrow \Delta H > 0$ [endothermic]

$\rightarrow \Delta V > 0$ [positive]

\rightarrow Examples:-

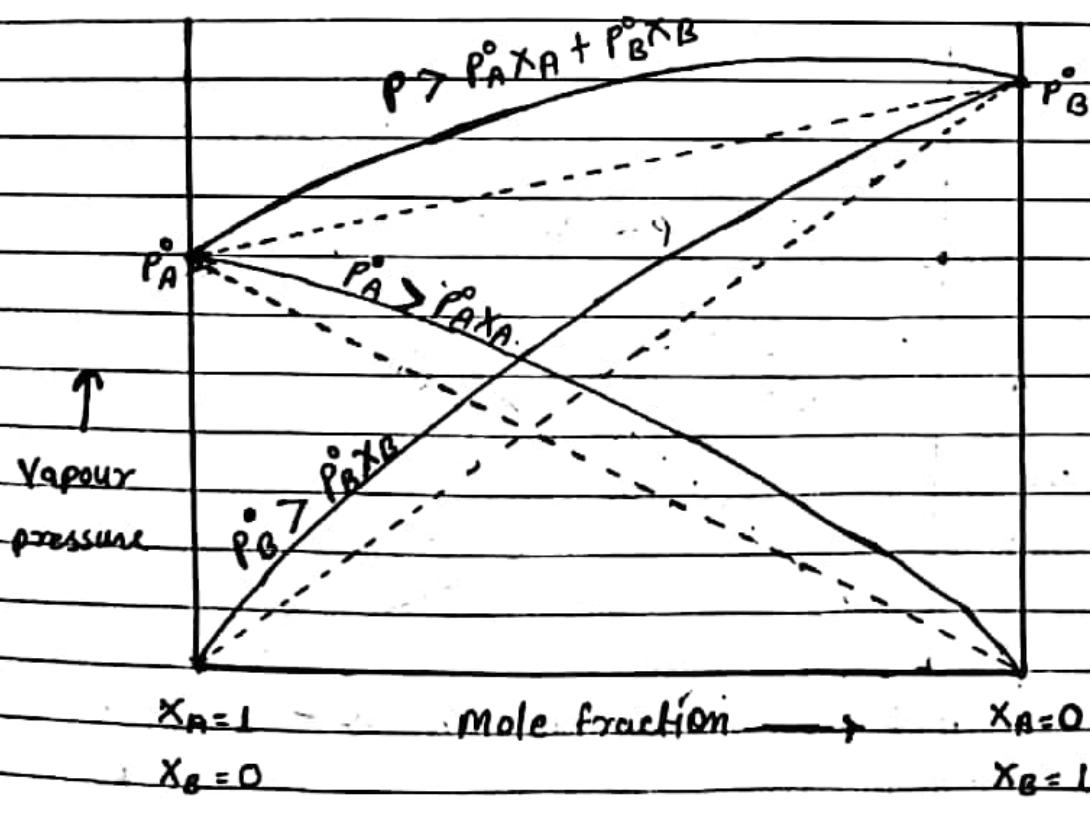
Methanol + water

Methanol + CCl_4

Acetone + Ethanol

Ethanol + Water

Graph:-



ii) Non-ideal solutions with negative deviation.

→ A-B interactions are stronger than A-A, B-B interactions.

→ $P_n < P_A^{\circ} X_A$ and $P_n < P_B^{\circ} X_B$

→ $P < P_A + P_B$

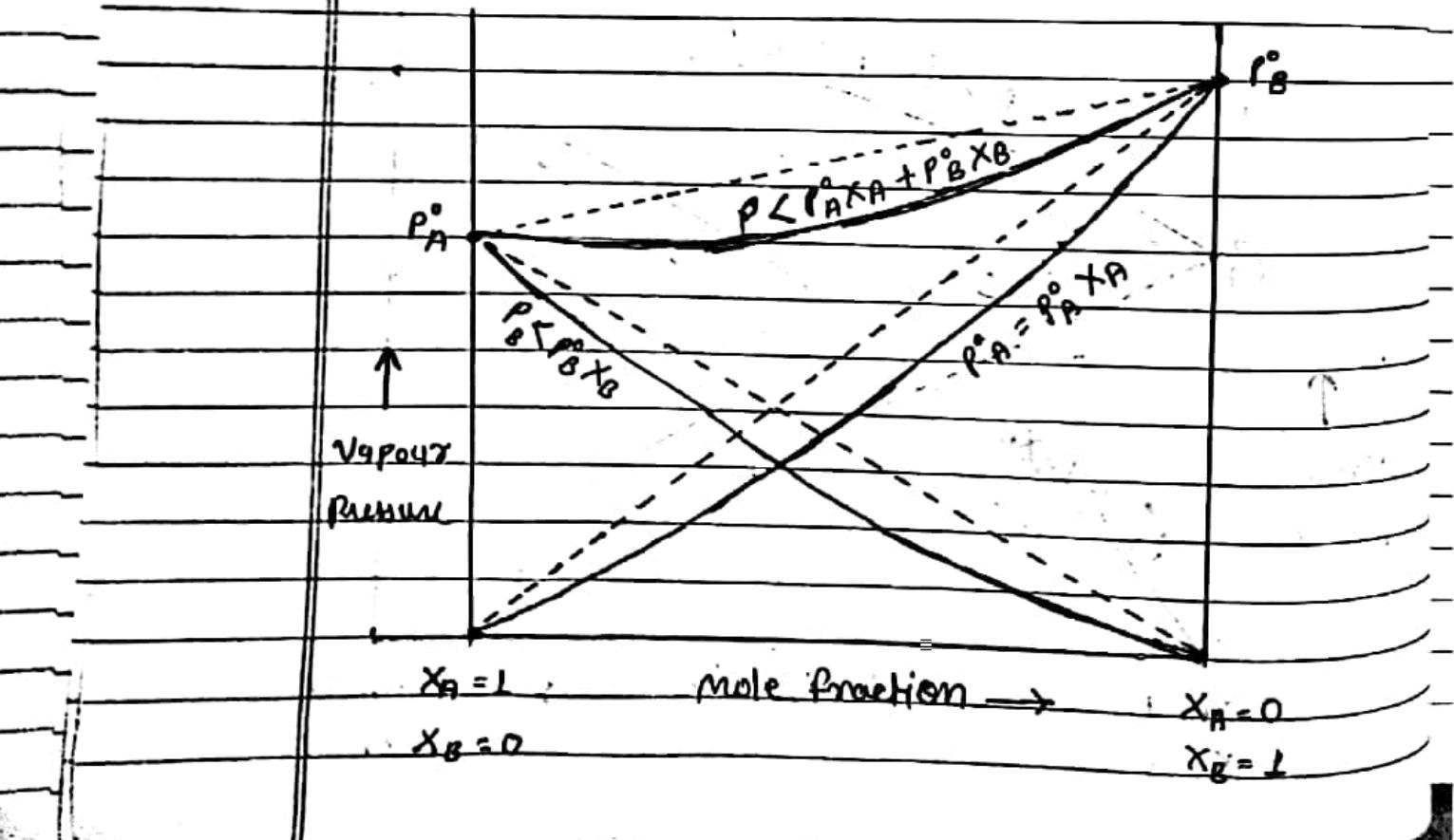
→ $\Delta H < 0$ [Exothermic]

→ $\Delta V < 0$ [Negative]

Examples: →

Chloroform + Acetone, Water + HCl
Acetone + Aniline, Chloroform + Bi

Graph: →



* Azeotropic Mixtures

- A binary solution of liquid solute and liquid solvent.
 - fixed composition of components.
 - boils at constant temperature.
 - The composition in liquid phase and vapour phase remains same for an azeotropic mixture.

$$X_A = Y_A \quad \& \quad X_B = Y_B$$

(liquid) (Vapour) (liquid) (Vapour)

- fractional distillation is not possible.

"Azeotrops are the constant boiling mixture which boils at constant temperature without any change in composition."

- There are two types of Azeotropic Mixtures:-

Minimum Boiling Azeotrops:-

- It is a non-ideal solution with positive deviation.
 - The boiling point of azeotropic solution is lower than the boiling point of either of pure component.
 - Example:- Water + Ethanol, Ethanol + chloroform.

Maximum Boiling Azeotropes:-

- It is a non-ideal solution with negative deviation.
 - The boiling point of Azeotropic solution is higher than the boiling point of either of pure Component

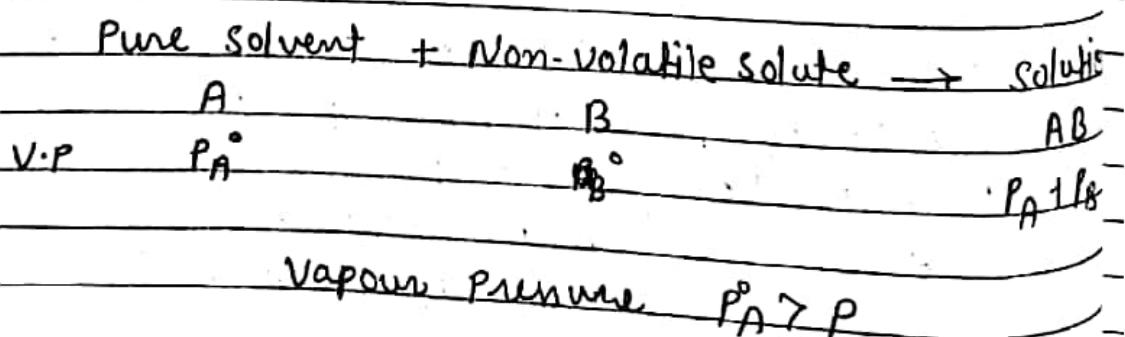
Example:- Acetone + chloroform
Methanol + Acetic acid.

* Colligative properties:-

- The properties of any homogeneous sol' which, all only depends upon the quantity of solute.
- The nature of solute does not affect any colligative property.
- The nature of solvent may affect the colligative property.
- There are four colligative properties which are as follows:-
 - i) Relative lowering of Vapour pressure
 - ii) Elevation in Boiling point
 - iii) Depression in freezing point
 - iv) Osmotic pressure

* Relative lowering of Vapour pressure :-

- When a non-volatile solute is added to pure solvent, its vapour pressure decreases. This decrease in vapour pressure is termed as lowering of Vapour pressure.



According to Raoult's law

$$P_A = P_A^{\circ} X_A \dots \dots \text{(i)}$$

$$P_B = P_B^{\circ} X_B$$

Now according to Dalton's law of Partial Pressure

$$P = P_A + P_B$$

Solute is non-volatile in nature so its vapour pressure is negligible.

$$\text{Now } P = P_A \dots \dots \text{(ii)}$$

by the eq. (i) & (ii)

$$P = P_A^{\circ} X_A \dots \dots \text{(iii)}$$

we know that

$$X_A + X_B = 1$$

$$X_A = 1 - X_B \dots \dots \text{(iv)}$$

By the equation (iii) & (iv)

$$P = P_A^{\circ} (1 - X_B)$$

$$P = P_A^{\circ} - P_A^{\circ} X_B$$

$$P_A^{\circ} X_B = P_A^{\circ} - P$$

$$X_B = \frac{P_A^{\circ} - P}{P_A^{\circ}} \dots \dots \text{(v)}$$

We know that

$$x_B = \frac{n_B}{n_A + n_B} \quad (\text{vi})$$

and

$$n_B = \frac{W_B}{M_B} \quad (\text{vii})$$

$$n_A = \frac{W_A}{M_A} \quad (\text{viii})$$

by the eq. (vi), (vii), (viii)

$$x_B = \frac{\frac{W_B}{M_B}}{\frac{W_A}{M_A} + \frac{W_B}{M_B}} \quad (\text{ix})$$

By the eq. (v) & (ix)

$$\frac{P_A^\circ - P}{P_A^\circ} = \frac{\frac{W_B}{M_B}}{\frac{W_A}{M_A} + \frac{W_B}{M_B}}$$

For a dilute solution:

$$\frac{W_B}{M_B} \ll \frac{W_A}{M_A}$$

so, $\frac{W_B}{M_B}$ is negligible

then

$$\frac{P_A^\circ - P}{P_A^\circ} = \frac{\frac{W_B}{M_B}}{\frac{W_A}{M_A}}$$

$$\frac{P_A^\circ - P}{P_A^\circ} = \frac{W_B}{M_B} \times \frac{M_A}{W_A}$$

||

on cross multiply

$$M_B = \frac{W_A \times M_A}{W_N} \left[\frac{P_A^e}{P_A^e - P} \right]$$

where

W_A = weight of solvent

W_B = weight of solute

M_A = molar mass of solvent

M_B = molar mass of solute

P_A^e = vapour pressure of pure solvent.

P_A = vapour pressure of solution.

* Elevation in Boiling point \Rightarrow

Boiling point \Rightarrow

The temperature at which the vapour pressure of the solution becomes equal to atmospheric pressure or external pressure.

→ Boiling point depends on external pressure. As external pressure change, boiling point also change.

→ When a non-volatile solute is added to a solvent, its boiling point is increases, this increment in B.P. is known as elevation in Boiling point.

→ It means boiling point of the solution will always be greater than the boiling point of the solvent.

B.P. of soln. $>$ B.P. of solvent

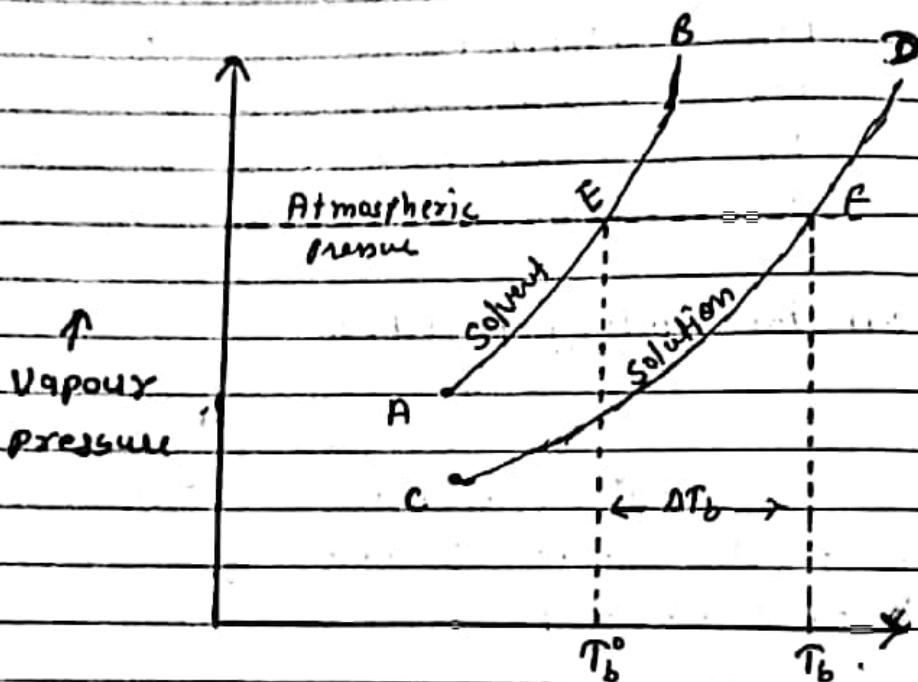
$$T_b > T_b^o$$

[$\because T_b = \text{Boiling Pt. Temp}$]

If the boiling point of a solution is T_b and
boiling point of solvent is T_b^o then

Elevation in B.P. = B.P. of Solⁿ - B.P. of Solvent

$$\Delta T_b = T_b - T_b^o \dots \dots \text{(i)}$$



By the above graph

$$\Delta T_b \propto \Delta P \dots \dots \text{(ii)}$$

Now according to Raoult's law

$$\Delta P \propto X_B \dots \dots \text{(iii)}$$

By the equation (ii) & (iii)

$$\Delta T_b \propto X_B$$

$$\Delta T_b = K_b \cdot X_B \quad \text{--- (iv)}$$

we know that

$$X_B = \frac{n_B}{n_A + n_B} \quad \text{--- (v)}$$

now;

$$n_B = \frac{W_B}{M_B} \quad \text{&} \quad n_A = \frac{W_A}{M_A} \quad \text{--- (vi)}$$

By the eq. (v) & (vi)

$$X_B = \frac{\frac{W_B}{M_B}}{\frac{W_A}{M_A} + \frac{W_B}{M_B}}$$

for dilute solution $\frac{W_A}{M_A} \ggg \frac{W_B}{M_B}$

$$\text{Then; } X_B = \frac{\frac{W_B}{M_B}}{\frac{W_A}{M_A}} \quad \text{--- (vii)}$$

By the eq. (iv) & (vii)

$$\Delta T_b = K \cdot X_B$$

$$= K \cdot \frac{\frac{W_B}{M_B}}{\frac{W_A}{M_A}}$$

$$\Delta T_b = K \times \frac{W_B}{M_B} \times \frac{M_A}{W_A} \quad \text{--- (viii)}$$

By the eq. (vi) and (viii)

$$\Delta T_b = K \times \frac{n_B \times M_A}{W_A} \quad \text{--- (ix)}$$

we know that

$$\text{Molality} = \frac{\text{no. of moles of Solute}}{\text{weight of solvent (kg)}}$$

$$m = \frac{n_B}{W_A} \quad (X)$$

By the eq. (ix) & (x)

$$\Delta T_b = K \times M_A \times m$$

$$\Delta T_b = K_b \cdot m \quad \therefore K_b = K \cdot M_A$$

where $\Delta T_b \rightarrow$ elevation in boiling point
 $K_b \rightarrow$ molal elevation constant
 $m \rightarrow$ molality.

* Relation between elevation in Boiling Point & molecular mass of solute (non-volatile) \Rightarrow

We know that $\Delta T_b = K_b \cdot m \dots \textcircled{1}$

$$\text{Molality } m = \frac{n_B}{W_A} \quad \textcircled{2}$$

$$\text{and } n_B = \frac{W_B}{M_B} \quad \textcircled{3}$$

By the equation $\textcircled{2}$ and $\textcircled{3}$

$$m = \frac{W_B}{M_B \times W_A} \quad \textcircled{4}$$

By the equation $\textcircled{1}$ & $\textcircled{4}$

$$\Delta T_b = \frac{K_b \cdot W_B}{M_B \times W_A}$$

Then

$$M_B = \frac{k_b \cdot w_B}{\Delta T_b \cdot w_A}$$

where;

w_A → Weight of Solvent

w_B → weight of Solute

ΔT_b → Elevation in Boiling point

k_b → Molal elevation constant

M_B → Molecular Mass of Solute.

* Relation between k_b and latent heat of vapourization:

$$k_b = \frac{M_A R T_b^2}{\Delta H_{\text{vapourisation}} \times 1000}$$

where M_A = molecular mass of solvent

R = Gas constant

T_b = Boiling point of solvent

ΔH_{vapour} = Latent Heat of Vapourization

* Unit of k_b and k_f :

$$k_b \& k_f = \frac{\Delta T}{m}$$

$$= \frac{^\circ C}{\text{mol/kg}}$$

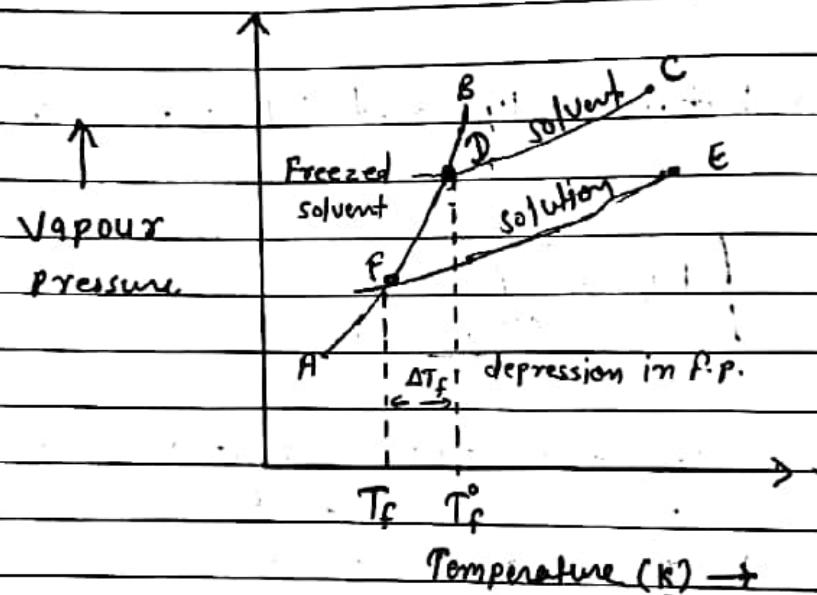
$$= \text{K} \cdot \text{kg} \cdot \text{mol}^{-1}$$

* Depression in freezing point: \Rightarrow

Freezing point \Rightarrow

The temperature at which the pressure of solid becomes equal to the atmospheric pressure.

\rightarrow When a non-volatile solute is added to a pure solvent, its freezing point decreases. This decrease in Freezing point is termed as depression in f.p.



If freezing point of pure solvent is T_f° and freezing point of solution T_f . Then:-

$$\text{Depression in F.P.} = \text{f.p. of solvent} - \text{f.p. of soln}$$

$$\Delta T_f = T_f^\circ - T_f \quad \text{--- (1)}$$

By the graph

$$\Delta P \propto \Delta T_f$$

(2)

According to Raoult's law.

$$\Delta P \propto X_B \quad \text{--- (3)}$$

By the eq. (2) & (3)

$$\Delta T_f \propto X_B$$

$$\Delta T_f = K X_B \quad \text{--- (4)}$$

Now,

$$X_B = \frac{n_B}{n_A + n_B} \times n_B = \frac{W_B}{M_B}, \quad n_A = \frac{W_A}{M_A}$$

$$X_B = \frac{\frac{W_B}{M_B}}{\frac{W_A}{M_A} + \frac{W_B}{M_B}}$$

$$\text{For dilute soln} \quad \frac{W_A}{M_A} \ggg \frac{W_B}{M_B}$$

$$\text{Then } X_B = \frac{\frac{W_B}{M_B}}{\frac{W_A}{M_A}}$$

$$X_B = \frac{W_B \times M_A}{M_B \cdot W_A} \quad \text{--- (5)}$$

By the eq (4) & (5)

$$\Delta T_p = K \cdot \frac{W_B \times M_A}{M_B \cdot W_A}$$

$$\Delta T_f = K \cdot \frac{n_B \times M_A}{W_A}$$

$$\boxed{\Delta T_f = k_f \cdot m}$$

$$\boxed{\begin{aligned} \therefore \frac{n_B}{W_A} &= m \\ b - k_f &= k \cdot M_A \end{aligned}}$$

where: ΔT_f - Depression in freezing point

k_f - Molal depression constant

m - molality

* Relation between depression in freezing point and molecular mass of non-volatile solute: \Rightarrow

We know that $\Delta T_f = K_f \cdot m$ — (i)

$$\text{Molality (m)} = \frac{W_B}{M_B \times W_A} \quad (\text{ii})$$

By the equation (i) & (ii)

$$\Delta T_f = K_f \cdot \frac{W_B}{M_B \times W_A}$$

$$M_B = \frac{K_f \times W_B}{\Delta T_f \times W_A}$$

where

W_B = weight of solute

M_B = molecular mass of solute.

W_A = weight of solvent.

K_f = molal depression constant

* Relation between K_f and latent Heat of Vapourisation

$$K_f = \frac{M_A R T_f^2}{\Delta H_{\text{vapour}} \times 1000}$$

where;

M_A = molecular mass of solvent

R = gas constant

T_f = freezing point of solvent

ΔH_{vapour} = Latent Heat of Vapourisation

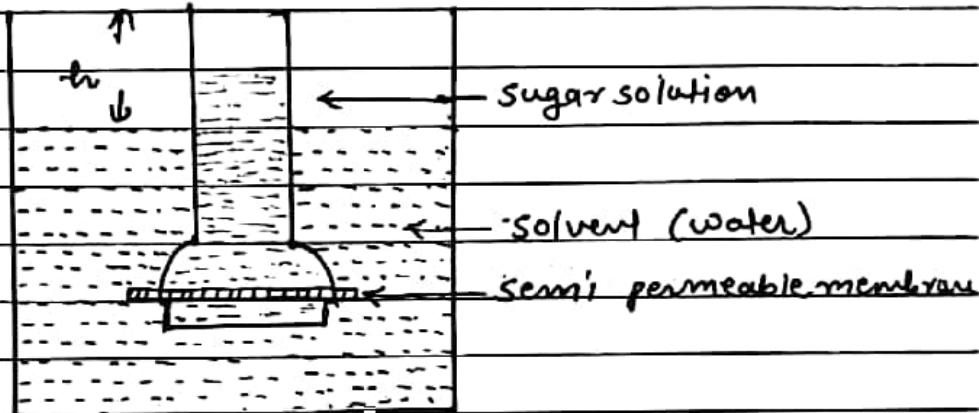
* Diffusion: \Rightarrow

It is process in which the movement of molecules takes place from higher concentration to lower concentration area.

e.g.: \Rightarrow Spread of odour of Camellie in room.

* Osmosis: \Rightarrow

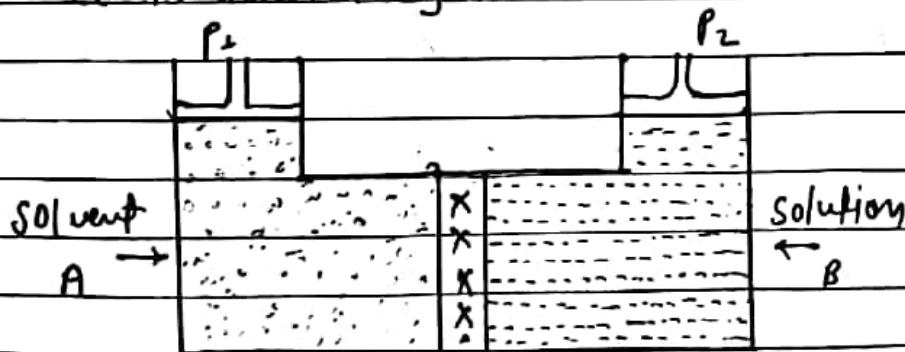
The spontaneous flow of solvent molecules from a dilute solution to a concentrated solution through a semi-permeable membrane is called osmosis.
It is represented by: \Rightarrow



* Osmotic pressure: \Rightarrow

The pressure required to stop the movement of solvent molecules from a dilute solution to a concentrate solution through a semi-permeable membrane is called osmotic pressure.

It is denoted by π .



* Rules for osmotic pressure: \Rightarrow

i) Van't Hoff Boyle Law: \Rightarrow

The osmotic pressure of solution is directly proportional to its concentration at a particular temperature.

$$\Pi \propto C \dots \dots \text{(i)}$$

$$C = \frac{L}{V} \dots \dots \text{(ii)}$$

By the equation (i) & (ii),

$$\boxed{\Pi \propto \frac{t}{V}}$$

ii) Van't Hoff Pressure-Temperature law: \Rightarrow

The osmotic pressure of solution is directly proportional to its absolute temperature at a particular concentration and dilution.

$$\Pi \propto T \dots \dots \text{(i)}$$

By the Van't Hoff Boyle law;

$$\Pi \propto \frac{t}{V} \dots \dots \text{(ii)}$$

By the eq (i) & (ii),

$$\Pi \propto \frac{T}{V}$$

$$\Pi = \frac{ST}{V}$$

$$\pi V = ST \quad [\text{for one mole}]$$

for n mole;

$$\pi V = nST$$

[$\because S = \text{solution constant}$]

The value of $S =$ The value of R

$$\boxed{\pi V = nRT}$$

* Relation between Osmotic pressure and molecular mass of solutes.

$$\pi V = nRT$$

$$\pi = \frac{n}{V} RT \quad \dots \dots \text{(i)}$$

We know that

$$n = \frac{w_B}{M_B} \quad \dots \dots \text{(ii)}$$

By the eq. (i) & (ii)

$$\pi = \frac{w_B \cdot R \cdot T}{M_B \cdot V}$$

$$\boxed{M_B = \frac{w_B \cdot R \cdot T}{\pi \cdot V}}$$

where;

w_B = Mass of solute

M_B = molecular mass of solute

R = gas constant

T = Absolute Temperature

π = Osmotic pressure

* Isotonic Solution: \Rightarrow Two solutions having equal osmotic pressure are called isotonic Solutions.

$$\pi_1 = \pi_2$$

$$c_1 RT_1 = c_2 RT_2$$

$$c_1 T_1 = c_2 T_2$$

* Hypertonic Solutions: \Rightarrow A solution having higher osmotic pressure with respect to other solution is called Hypertonic solution.

* Hypotonic Solution: \Rightarrow A solution having lower osmotic pressure with respect to other solution called Hypotonic solution.

* Reverse Osmosis: \Rightarrow The direction of is reversed when a applied pressure is larger than osmotic pressure. Due to which pure solvent flows out of the solution through semi-permeable membrane. This phenomenon is called Reverse osmosis.

\rightarrow It occurs when a pressure larger than osmotic pressure is applied to the solution.

* Van't Hoff factor \Rightarrow

- It is introduced by Van't Hoff.
- It Explain the Abnormal molecular mass.

* Abnormal Molecular Mass: \Rightarrow

It is the molecular mass of solute which is different from actual molecular mass due to dissociation and association of solute particle.

- Colligative property is inversely proportional to the molecular mass.
- Van't Hoff factor is represented by 'i'.

$$i = \frac{\text{Total no. of moles of solute after dissociation / Association}}{\text{Total no. of moles of solute before dissociation / Association}}$$

$$i = \frac{\text{Observed value of Colligative property}}{\text{Calculated value of colligative property}}$$

$$i = \frac{\Delta P_{\text{observed}}}{\Delta P_{\text{calculated}}} = \frac{\Delta T_b^{\text{observed}}}{\Delta T_b^{\text{calculated}}} = \frac{\Delta T_f^{\text{observed}}}{\Delta T_f^{\text{calculated}}} = \frac{T_{\text{observed}}}{T_{\text{calculated}}}$$

If $i = 1 \Rightarrow$ The solute does not dissociated or associated in the solution.

$i > 1 \Rightarrow$ The solute is dissociated.

$i < 1 \Rightarrow$ The solute is associated.

* Change in colligative properties: \rightarrow

i) Relative lowering of vapour pressure

$$\frac{P_A^o - P}{P_A^o} = i \cdot X_B$$

ii) Elevation in Boiling point

$$\Delta T_b = i \cdot K_b \cdot m$$

iii) Depression in freezing point

$$\Delta T_f = i \cdot K_f \cdot m$$

iv) Osmotic pressure:-

$$\Pi = i c R T$$

* Relation between Van't Hoff factor (i) and Degree of dissociation (α):-

→ Van't Hoff factor is represented by i .

→ Degree of dissociation is represented by α .

Now;



no. of moles before dissociation	1	0
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no. of moles after dissociation	$1 - \alpha$	$n\alpha$
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no. of moles before dissociation = $10 \times L$

no. of moles after dissociation: $L - \alpha + n\alpha$

$$L - \alpha + n\alpha = L + \alpha(n-1)$$

We know that

$$i = \frac{\text{no. of moles after dissociation}}{\text{no. of moles before dissociation}}$$

$$i = \frac{L + \alpha(n-1)}{L}$$

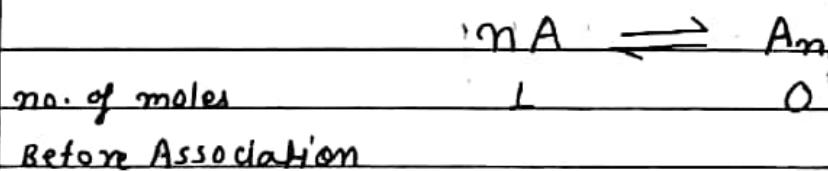
$$i = L + \alpha(n-1)$$

$$\alpha(n-1) = i - L$$

$$\boxed{\alpha = \frac{i - L}{(n-1)}}$$

* Relation between Van't Hoff factor and degree of association:

Let n molecules of solute associate to form a molecule and degree of association is α .



no. of mole	$1 - \alpha$	$\frac{\alpha}{n}$
After Association		

no. of moles before association = L_0 = L

no. of moles after association = $L - \alpha + \frac{\alpha}{n}$

we know that

$$i = \frac{\text{no. of moles after association}}{\text{no. of moles before association}}$$

$$i = \frac{L - \alpha + \frac{\alpha}{n}}{L}$$

$$i = \frac{L - \alpha + \frac{\alpha}{n}}{L}$$

$$i - 1 = \frac{\alpha}{n} - \alpha$$

$$i - 1 = \frac{\alpha - n\alpha}{n}$$

$$n(i-1) = \alpha(1-n)$$

$$\boxed{\alpha = \frac{n(i-1)}{(1-n)}}$$