

# SOLUTION

## CLASS : XII



## **Solution :-**

**Homogenous mixture of two or more substance in single phase .**

**e.g. Brass mixture of copper and zinc different from German silver (Copper ,Zinc and Nickel )**



**Brass**



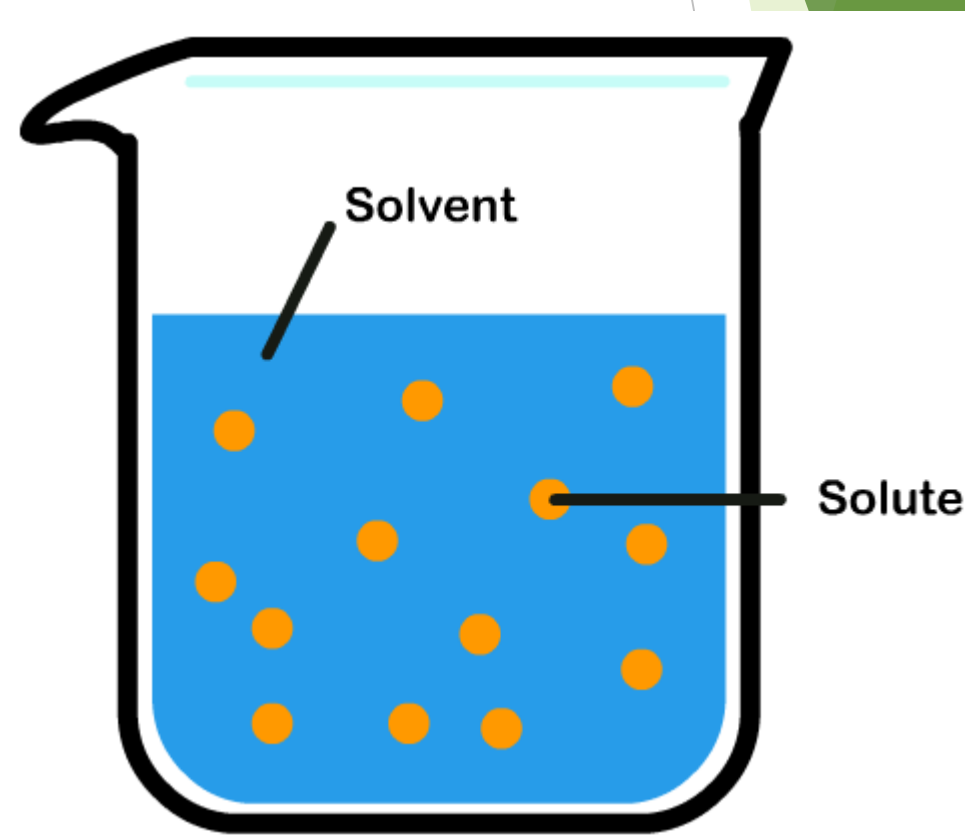
**German silver**

## Components of binary solution:

Solution = Solvent + Solute

**Solvent** : It is the component of the solution which is present in a relatively large proportion in the solution whose physical state is same as that of the resulting solution.

**Solute** : It is the second component of a binary solution which is present relatively small proportion.



## **Types of Solutions :**

In a binary solution components can be solid, liquid or gas.

There are three types of the Solutions:

- **Gaseous Solution**
- **Liquid Solutions**
- **Solid Solutions**

Type of Solution	Solute	Solvent	Common Examples
Gaseous Solutions	Gas	Gas	Mixture of oxygen and nitrogen gases
	Liquid	Gas	Chloroform mixed with nitrogen gas
	Solid	Gas	Camphor in nitrogen gas
Liquid Solutions	Gas	Liquid	Oxygen dissolved in water
	Liquid	Liquid	Ethanol dissolved in water
	Solid	Liquid	Glucose dissolved in water
Solid Solutions	Gas	Solid	Solid Solution of hydrogen in palladium
	Liquid	Solid	Amalgam of mercury with sodium
	Solid	Solid	Copper dissolved in gold

## Expressing the Concentration of Solutions :

The Concentration of the solution refers to the amount of solute present in the given quantity of solution or solvent.

**1. Mass percent (w/w) :** It is defined as the number of parts by mass of solute per hundred parts by mass of solution

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

If  $W_B$  be the mass of solute (B) and  $W_A$  be the mass of solvent (A), then

$$\text{Mass \% of B} = \frac{W_B}{W_A + W_B} \times 100$$

**Example :** Calculate the mass percentage of glucose in the solution if 10 g of glucose is dissolved in 90 g of water.

**Solution :**

Mass of glucose/solute ( $W_B$ ) = 10 g

Mass of water/solvent ( $W_A$ ) = 90 g

$$\begin{aligned}\text{mass \% of glucose} &= \frac{\text{mass of the glucose}}{\text{total mass of solution}} \times 100 \\ &= \frac{10 \text{ g}}{10 \text{ g} + 90 \text{ g}} \times 100 \\ &= \frac{10 \text{ g}}{100 \text{ g}} \times 100 \\ &= 10 \%\end{aligned}$$

## 2. Volume percentage (V/V):

It is defined as the number of part by volume of solute per hundred parts by volume of solution.

$$\text{Volume \% of a components} = \frac{\text{Volume of the component}}{\text{Total volume of solution}} \times 100$$

If  $V_B$  be the volume of solute (B) and  $V_A$  be the volume of solvent (A), Then

$$\text{Volume \% of B} = \frac{V_B}{V_A + V_B} \times 100$$





### 3. Mass by volume percentage (w/v) :

It is the mass of solute dissolved in 100 mL of the solution.

$$\text{Percentage(w / V)} = \frac{\text{Mass of solute X 100}}{\text{Volume of solution}}$$

### 4. Parts per million (ppm) :

It is defined as the number of parts by mass of solute per million parts by mass of solution.

$$\text{ppm} = \frac{\text{Number of parts of the component}}{\text{Total number of all components of the solution}} \times 10^6$$



## 5. Mole Fraction (x) :

Mole fraction is defined as the ratio of number of moles of one component to the total number of moles of all the components present in the solution.

$$\text{Mole fraction of a component} = \frac{\text{Number of the moles of the components}}{\text{Total number of moles of all the components}}$$

Let us consider a solution that contains the components A and B. If the number of moles of A and B are  $n_A$  and  $n_B$  respectively, then

$$\text{Mole fraction of A } (x_A) = \frac{n_A}{n_A + n_B}$$

We know

$$\text{Number of moles} = \frac{\text{Grams of compound}}{\text{molar mass of compound}}$$

if the mole fraction of A is  $x_A$  and mole fraction of B is  $x_B$ , then

$$x_A + x_B = 1$$

The sum of mole fractions of all the components in a solution is always unity.

**Example :** Calculate the mole fraction of ethylene glycol ( $C_2H_6O_2$ ) in a solution containing 20% of  $C_2H_6O_2$  by mass.

**Solution :**

Assume that total mass of solution = 100 g

Mass of  $H_2O$  = 80 g

Mass of  $C_2H_6O_2$  = 20 g

Molar mass of  $H_2O$  =  $1 \times 2 + 16 \times 1 = 18 \text{ g mol}^{-1}$

Molar mass of  $C_2H_6O_2$  =  $12 \times 2 + 1 \times 6 + 16 \times 2 = 62 \text{ g mol}^{-1}$

$$\begin{aligned}\text{Moles of } H_2O &= \frac{\text{Grams of } H_2O}{\text{molar mass of } H_2O} \\ &= \frac{80 \text{ g}}{18 \text{ g mol}^{-1}} \\ &= 4.444 \text{ mol}\end{aligned}$$

$$\begin{aligned}\text{Moles of } C_2H_6O_2 &= \frac{\text{Grams of } C_2H_6O_2}{\text{molar mass of } C_2H_6O_2} \\ &= \frac{20 \text{ g}}{62 \text{ g mol}^{-1}} \\ &= 0.322 \text{ mol}\end{aligned}$$

$$\begin{aligned}x_{C_2H_6O_2} &= \frac{\text{Moles of } C_2H_6O_2}{\text{Moles of } C_2H_6O_2 + \text{Moles of } H_2O} \\ &= \frac{0.322 \text{ mol}}{0.322 \text{ mol} + 4.444 \text{ mol}} \\ &= 0.068\end{aligned}$$

$$\begin{aligned}x_{H_2O} &= \frac{\text{Moles of } H_2O}{\text{Moles of } C_2H_6O_2 + \text{Moles of } H_2O} \\ &= \frac{4.444 \text{ mol}}{0.322 \text{ mol} + 4.444 \text{ mol}} \\ &= 0.932\end{aligned}$$

As we know that the sum of mole fractions of all the components is in a solution is always unity,

so mole fraction of water we can also be calculate as :

$$1 - 0.068 = 0.932$$

## 6. Molarity (M) :

It is defined as the number of gram mole of the solute present in one litre of the solution.

**Note : Molarity is a function of temperature, Because Volume depends on Temperature**

$$\begin{aligned}\text{Molarity} &= \frac{\text{mass of solute in gram}}{\text{Molar mass of solute} \times \text{Volume of solution in litre}} \\ &= \frac{W_B}{M_B \times V(l)} = \frac{W_B \times 1000}{M_B \times V(ml)}\end{aligned}$$

Here -

$W_B$  = mass of solute

$M_B$  = molar mass of solute

$V$  = Volume of solution in litter

**Example :** Calculate the molarity of a solution containing 20 g of NaOH in 500 ml solution.

**Solution :**

$$\text{Mole of NaOH} = \frac{20g}{40 g mol^{-1}} = 0.5 \text{ mol}$$

$$\begin{aligned} \text{Molarity of solution} &= \frac{\text{mole of NaOH}}{\text{Volume of Solution}} \\ &= \frac{0.5 \text{ mol}}{500 \text{ ml}} \times 1000 \\ &= 1 \text{ mol } L^{-1} \end{aligned}$$

If one mole of solute present in one litre of solution then solution known as **molar solution**.

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## 7. Molality (m):

It is defined as the number of moles of the solute present in one kilogram (1000 g) of solvent.

$$\text{Molality} = \frac{\text{mole of solute}}{\text{Mass of the solvent in kg}}$$

As we know

$$\text{Mole} = \frac{\text{Mass in gram of compound}}{\text{Molar mass of compound}}$$

$$\text{Molality} = \frac{\text{mass of solute in gram} \times 1000}{\text{Molar mass of solute} \times \text{Mass of solvent in g}}$$

$$= \frac{W_B \times 1000}{M_B \times W_A(g)}$$

Here -

$W_B$  = mass solute

$M_B$  = molar mass of solute

$W_A$  = mass of solvent

**Note :** Molality does not depend on temperature.

If one mole of solute present in one kilogram of solvent then solution known as molal solution.

**Example :** Calculate the molality of a solution if 100 g of NaOH present in 1 kg of solution.

**Solution :**

$$\text{Mass of Solute} = 100 \text{ g}$$

$$\text{Mole of solute} = \frac{100 \text{ g}}{40 \text{ g mol}^{-1}}$$

$$= 2.5 \text{ mol}$$

$$\text{Mass of solvent} = \text{Mass of solution} - \text{mass of solute}$$

$$= 1000 \text{ g} - 100 \text{ g}$$

$$= 900 \text{ g}$$

$$\text{Molality} = \frac{\text{Mole of solute}}{\text{Mass of the solvent in kg}}$$

$$= \frac{2.5 \text{ mol}}{900 \text{ g}} \times 1000$$

$$= 2.77 \text{ mol kg}^{-1}$$

**Solubility : Solubility of a substance is the maximum amount of substance dissolved in a given amount of solvent at a particular temperature**

## **Solubility of a Solid in a Liquid**

In general, a solute dissolves in a solvent if the intermolecular interactions are similar in the two or we may say **like dissolves like**. Solution in which no more solute can be dissolved at the same temperature and pressure is called a **saturated solution**. An *unsaturated solution* is one in which more solute can be dissolved at the same temperature. solubility of one substance into another depends on the nature of the substances, Temperature and pressure

### *Effect of temperature*

The solubility of a solid in a liquid is significantly affected by temperature changes. the dissolution process is endothermic ( $\Delta_{\text{sol}} H > 0$ ), the solubility should increase with rise in temperature and if it is exothermic ( $\Delta_{\text{sol}} H < 0$ ) the solubility should decrease. These trends are also observed experimentally.

### *Effect of pressure*

Pressure does not have any significant effect on solubility of solids in liquids. It is so because solids and liquids are highly incompressible and practically remain unaffected by changes in pressure



# Solubility of a gas in a liquid

Solubility of a gas in liquids depends on the nature of the substance, pressure and temp.

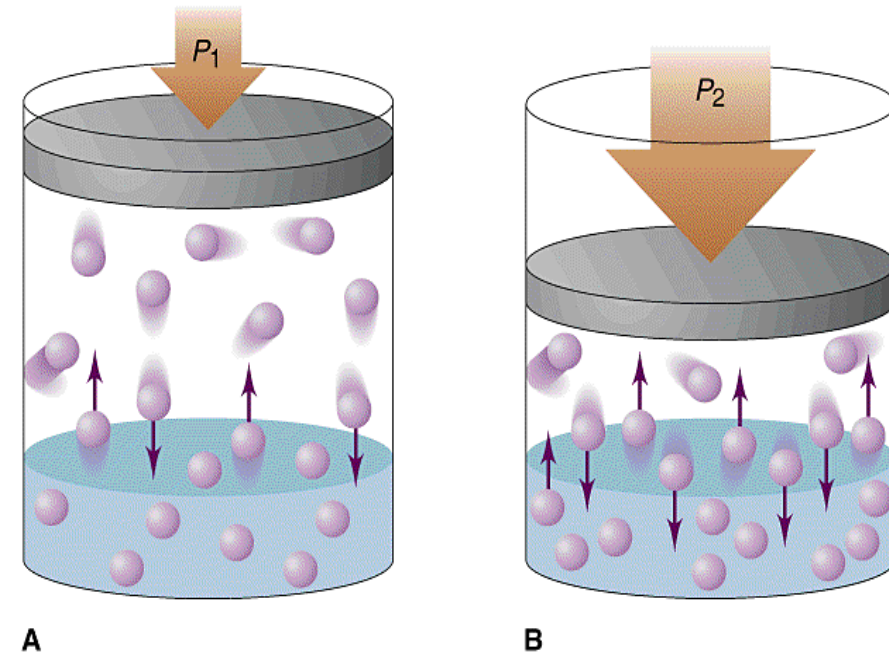
Solubility of gases increases with increase in pressure.

## (Henry's Law):-

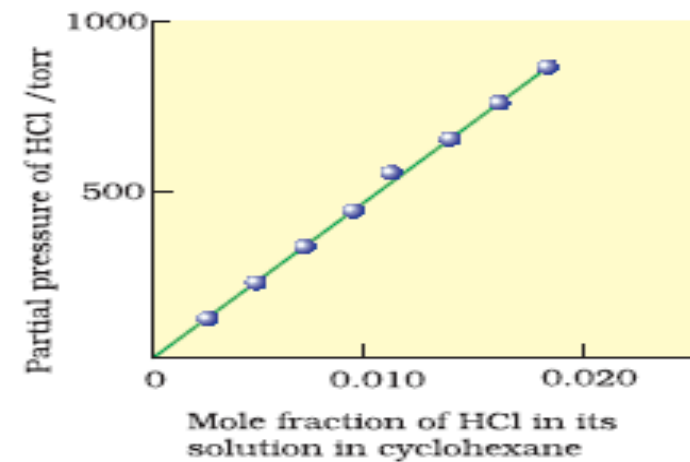
Partial pressure of gas in vapour phase is directly proportional to mole fraction of gas in the solution.

$$P_g = K_H \cdot X_g$$

$P_g$  = Partial pressure of gas  
 $K_H$  = Henry Law gas constant  
 $X_g$  = Mole fraction of gas

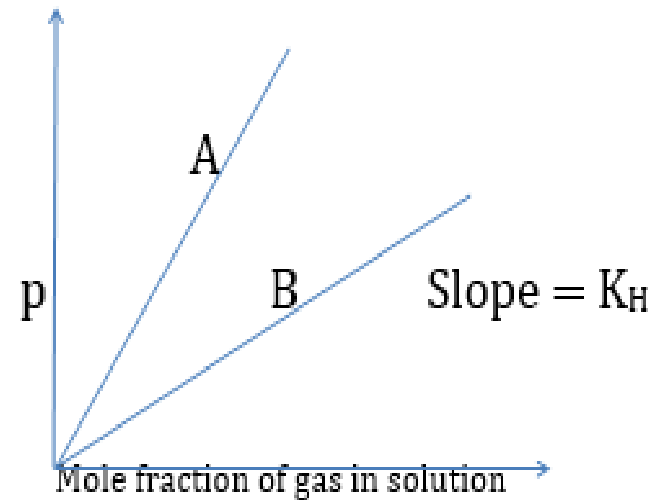


pressure vs solubility



**Lower the value of  $K_H$  at a given pressure , higher is the solubility of the gas in the liquid. Increase in temp increases the  $K_H$  value and decreases the solubility of the gas.**

**Gas B has lower  $K_H$  , hence it is more soluble than gas A**



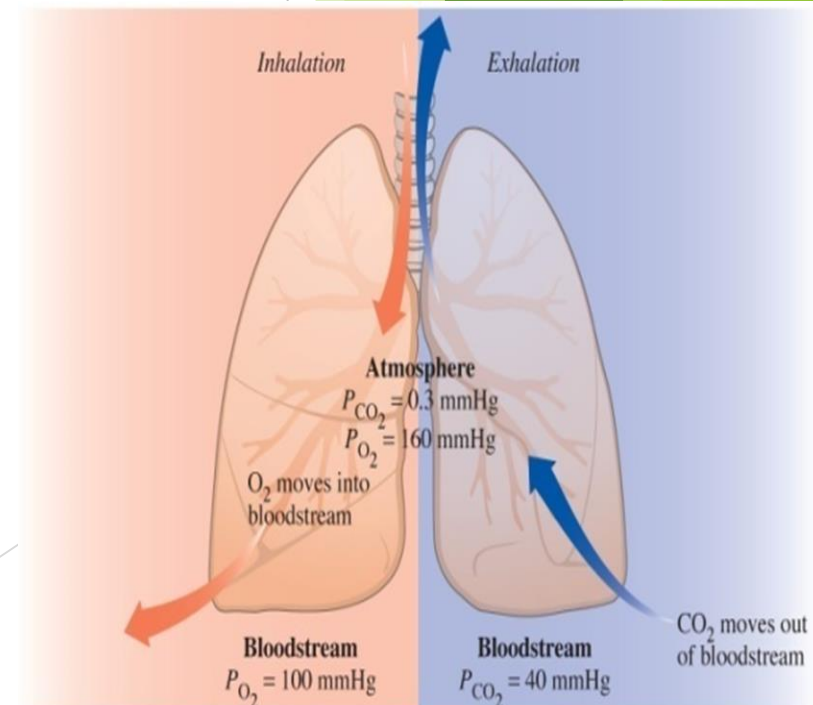
**According to Le Chatelier's principle increase in temperature shifts the equilibrium in the backward direction. Hence, solubility of gases decreases with increase of temperature. Moreover the kinetic energy of gas molecules increases with increase In temperature and thus they escape from solution.**

# Application of Henry Law

- ❑ Aquatic species are more comfortable in cold water rather than in warm water
- ❑ Deep sea divers [scuba divers] carry cylinders containing oxygen mixed with less soluble helium to avoid bends
- ❑ Soft drinks and soda water bottles are sealed under high pressure.
- ❑ Partial pressure of oxygen in high altitude is less, the concentration of oxygen in blood of people living at higher altitude or climbers is reduced. This condition is known as anoxia which causes weakness and inability to think.
- ❑ In lungs the partial pressure of oxygen is high. Hence it dissolves in blood hemoglobin to form oxy haemoglobin, whereas in tissues partial pressure of oxygen is low hence oxygen is released by oxyhaemoglobin for utilization in cellular functions



c.



Q1. Calculate the number of millimoles of  $\text{N}_2$  that would dissolve in 1 litre of water if its partial pressure is 0.987 bar.  $K_{\text{H}} = 76.48$  kbar.

Q2. If the solubility of  $\text{H}_2\text{S}$  in water at STP is 0.195 m, calculate Henry's law constant.

# SOLUTION OF TWO VOLATILE LIQUIDS

## Vapour Pressure of Liquid Solutions:

Pressure exerted by vapours of liquid at constant temperature and in equilibrium state is called Vapour Pressure of liquid. It depends on Nature of liquid, Surface area of the liquid, Temperature

### Raoult's Law:

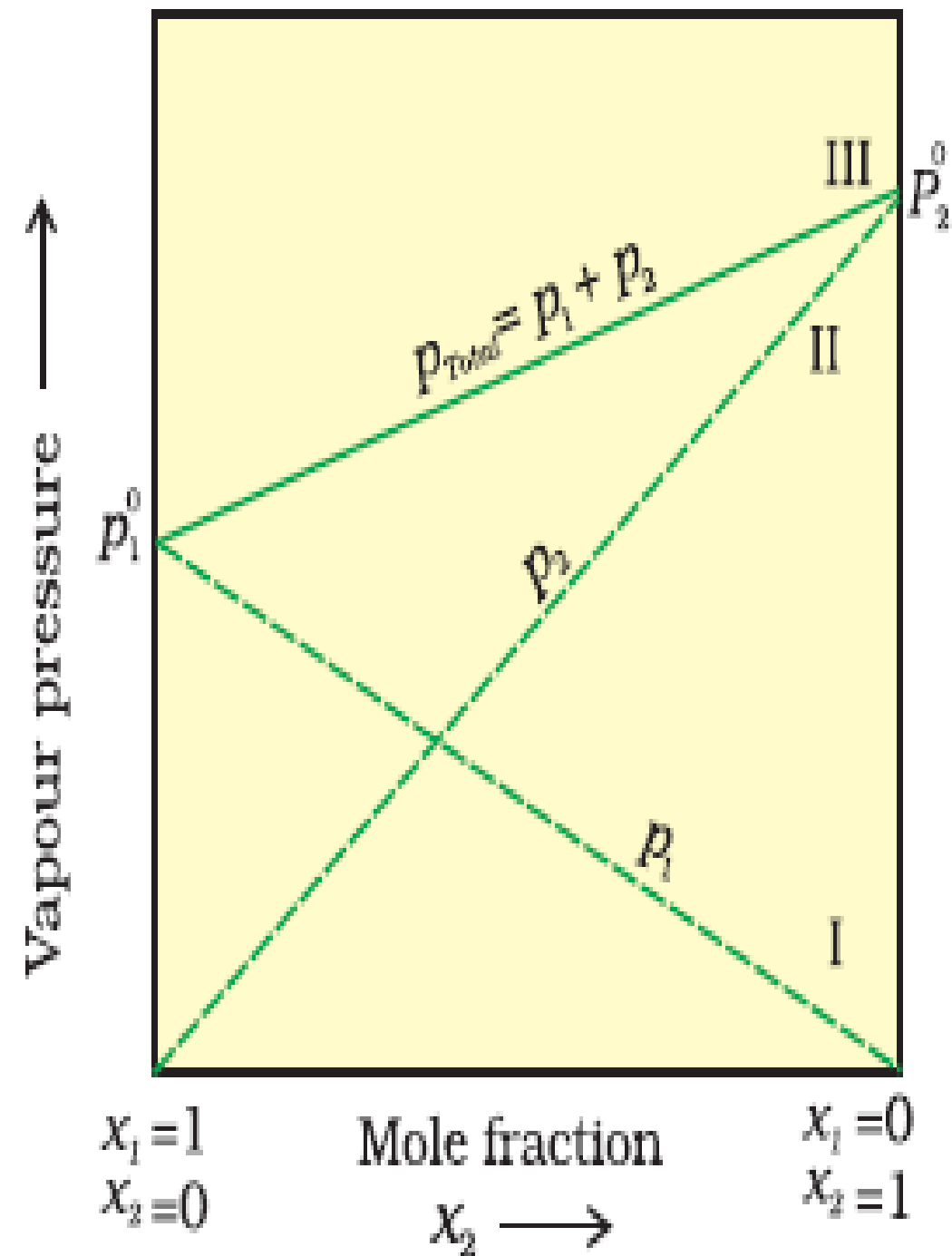
*For a solution of two or more volatile liquids, the partial vapour pressure of each component is directly proportional to its mole fraction at a given temperature.* The vapour pressure of the solution will be the sum of the partial vapour pressure of the components.

Consider a solution having two components A and B. Let  $p_A$  and  $p_B$  are the partial vapour pressures,  $\chi_A$  and  $\chi_B$  are the mole fractions of the two components and  $p_A^0$  and  $p_B^0$  are the vapour pressures of pure components.

$$p_A = p_A^0 \chi_A$$

$$p_B = p_B^0 \chi_B$$

$$\text{Total vapour pressure, } p_{\text{total}} = p_A + p_B$$



$$p_A = p_A^\circ \chi_A$$

$$p_B = p_B^\circ \chi_B$$

$$\text{Total vapour pressure, } p_{\text{total}} = p_A + p_B$$

On substituting the values of  $p_A$  and  $p_B$  in  $p_{\text{total}}$

$$p_{\text{total}} = p_A^\circ \chi_A + p_B^\circ \chi_B$$

$$= p_A^\circ [1 - \chi_B] + p_B^\circ \chi_B$$

$$= p_A^\circ - p_A^\circ \chi_B + p_B^\circ \chi_B$$

$$p_{\text{total}} = p_A^\circ + [p_B^\circ - p_A^\circ] \chi_B$$

The composition of vapour phase in equilibrium with the solution is determined by the partial pressures of the components. If  $y_1$  and  $y_2$  are the mole fractions of the components 1 and 2 respectively in the vapour phase then, using Dalton's law of partial pressures:

$$p_1 = y_1 p_{\text{total}}$$

$$p_2 = y_2 p_{\text{total}}$$

In general

$$p_i = y_i p_{\text{total}}$$

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

According to Raoult's law the V.P of a volatile component ( $p_A$ ) of the solution is equal to the product of its mole fraction ( $\chi_A$ ) in solution and V.P in pure state. ( $P_A^\circ$ )

$$p_A = p_A^\circ \chi_A$$

In the solution of a gas in liquid, one of the component is volatile and its solubility is given by Henry's law

$$P_A = K_H \chi_A$$

If we compare the two equations, they differ only in their proportionality constant  $K_H$  &  $P_A^\circ$ .

Thus Raoult's law is a special case of Henry's law.

# Ideal solution:-

Ideal Solution - Solution which follows Raoult's Law over all range of concentration

Ex. Solution made by benzene and toluene

Intermolecular attractive forces between A-A and B-B are same those between A-B

Positive Deviation:

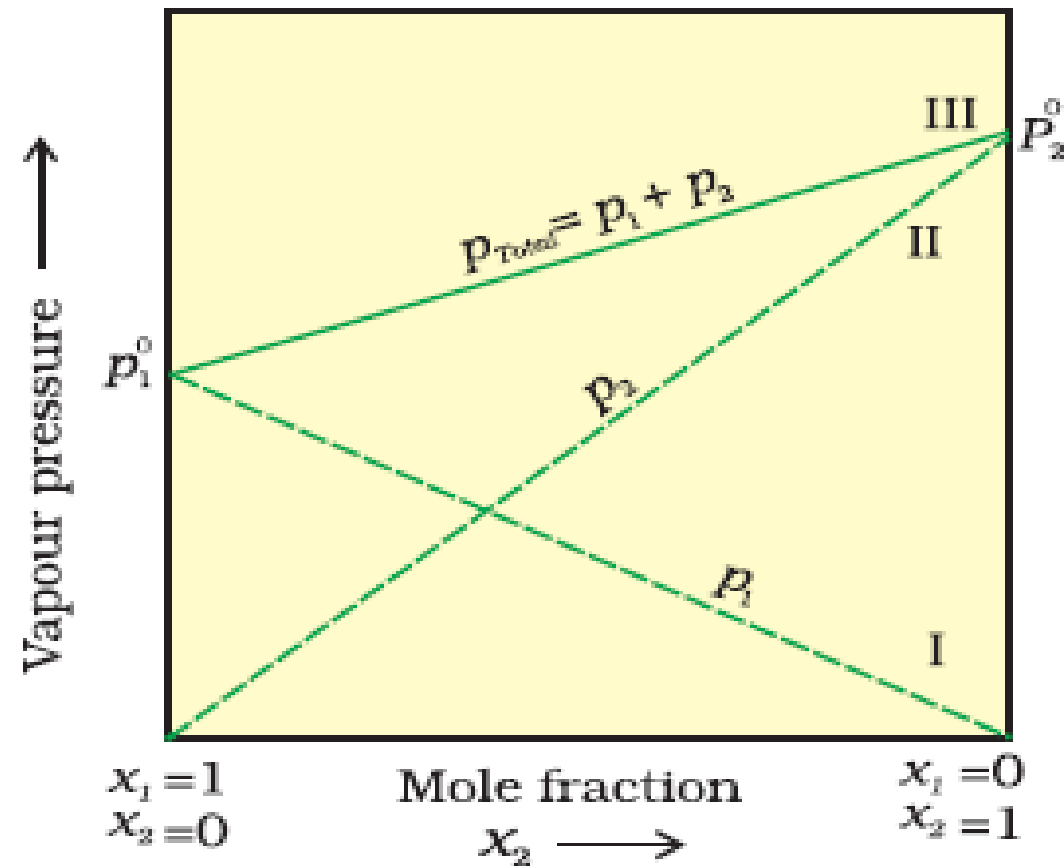
$$1. P_A = P_A^\circ \chi_A \text{ \& \cdot } P_B = P_B^\circ \chi_B$$

$$2. \Delta H_{\text{mix}} = 0$$

$$3. \Delta V_{\text{mix}} = 0$$

Example:

Benzene and Toluene



# Non ideal solution

Solution which does not follows Raoult's Law over all range of concentration

- **Cause for these deviations lie in the nature of interactions at the molecular level.**

## Types of Non Ideal Solution

- **Solution show positive deviation form Raoults' law**
- **Solution show negative deviation form Raoult law**



# Positive Deviation

➤ In the solution which shows positive deviation the attractive forces between A-A and B-B are stronger than A-B, leads to decrease in vapour pressure resulting.

In Positive Deviation:

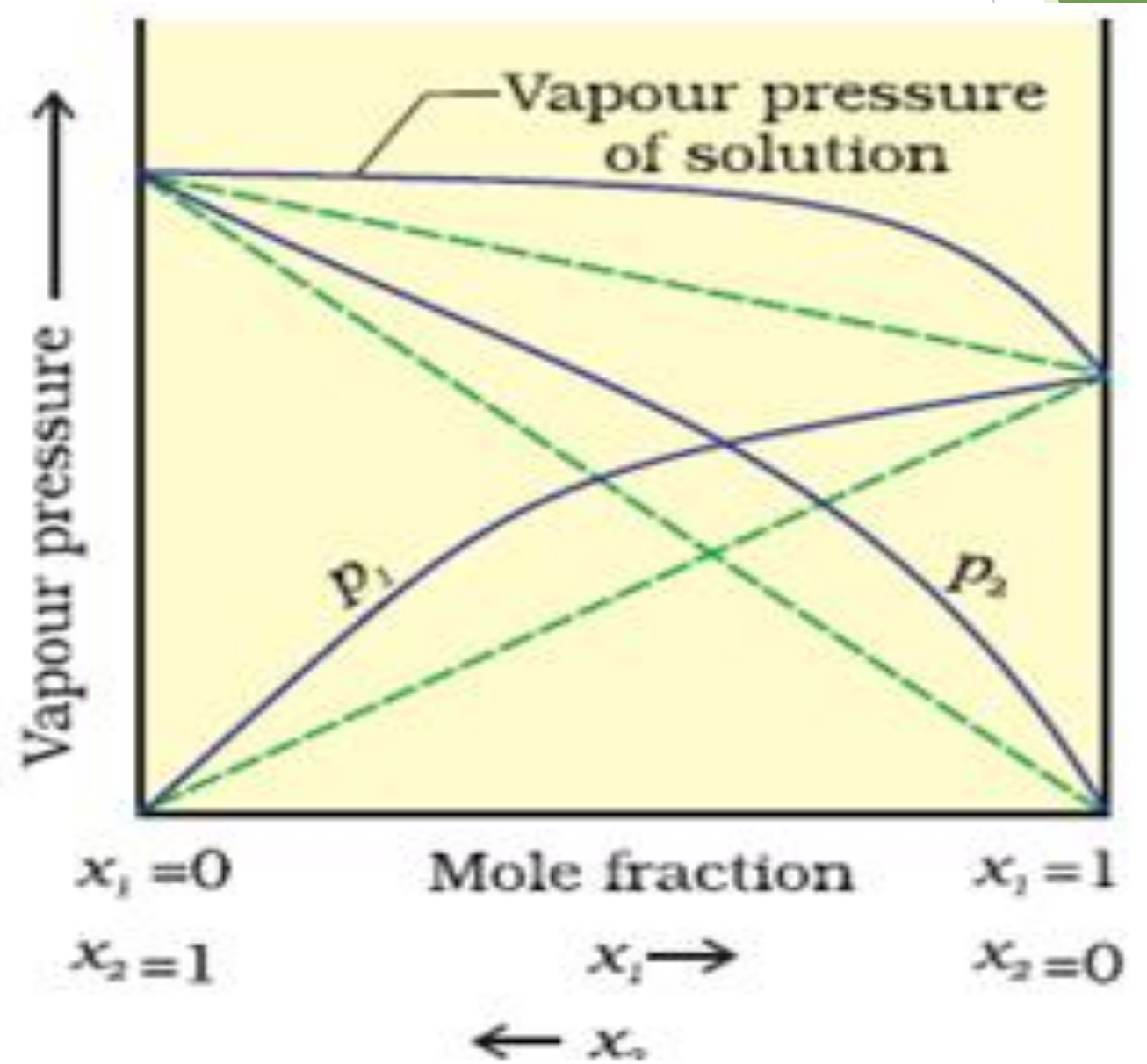
1.  $P_A > P_A^0 \chi_A$  &  $P_B > P_B^0 \chi_B$

2.  $\Delta H_{\text{mix}} > 0$

3.  $\Delta V_{\text{mix}} > 0$

Example:

Cyclo hexane and Ethanol



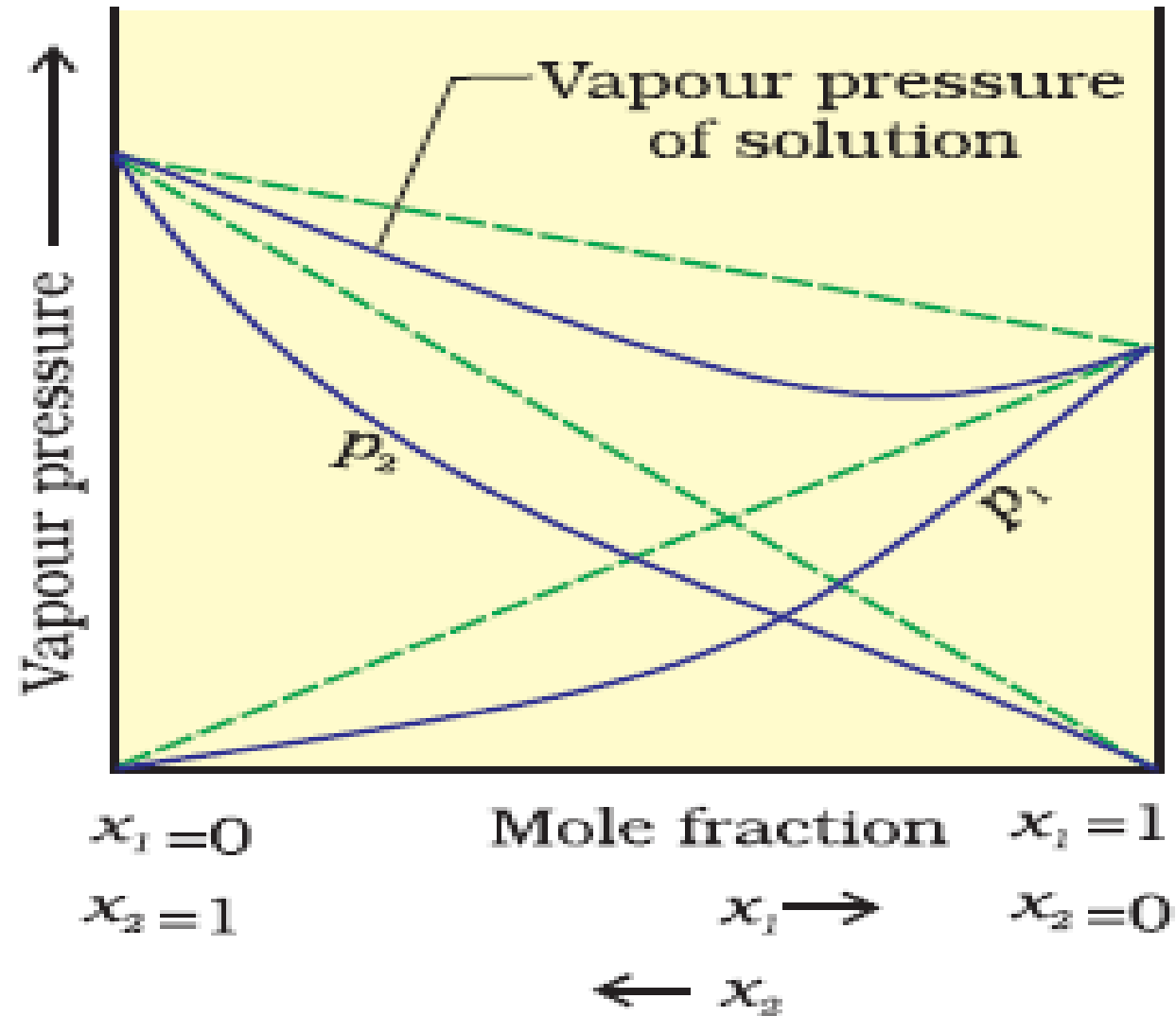
# Negative deviation

Attractive forces between A-A and B-B are weaker than A-B, leads to decrease in vapour pressure .

In Negative Deviation:

1.  $P_A < P_A^0 \chi_A$
- &  $P_B < P_B^0 \chi_B$
2.  $\Delta H_{\text{mix}} < 0$
3.  $\Delta V_{\text{mix}} < 0$

Chloroform and Acetone



# Azeotropes

Binary mixtures having the same composition in liquid and vapour phase and boil at a constant temperature. In such cases, it is not possible to separate the components by fractional distillation.

## Two types of Azeotropes

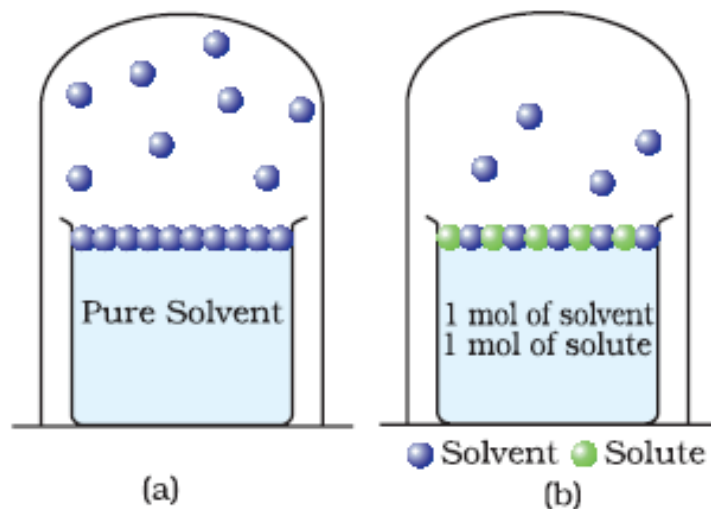
- **Minimum boiling Azeotrope** formed by those liquid pairs which show +ve deviation from ideal behavior .Such azeotropes have boiling point lower than either of the components  
e.g. Ethanol(95%) and water(5%) mixture
- **Maximum boiling azeotrope.** formed by those liquid pairs which show -ve deviation from ideal behavior .Such azeotropes have boiling point higher than either of the components.  
e.g. Nitric acid(68%) and water(32%)

Q1. Vapour pressure of chloroform ( $\text{CHCl}_3$ ) and dichloromethane ( $\text{CH}_2\text{Cl}_2$ ) at 298 K are 200 mm Hg and 415 mm Hg respectively. (i) Calculate the vapour pressure of the solution prepared by mixing 25.5 g of  $\text{CHCl}_3$  and 40 g of  $\text{CH}_2\text{Cl}_2$  at 298 K and (ii) mole fractions of each component in vapour phase.

Q2. Heptane and octane form an ideal solution. At 373 K, the vapour pressures of the two liquid components are 105.2 kPa and 46.8 kPa respectively. What will be the vapour pressure of a mixture of 26.0 g of heptane and 35 g of octane?

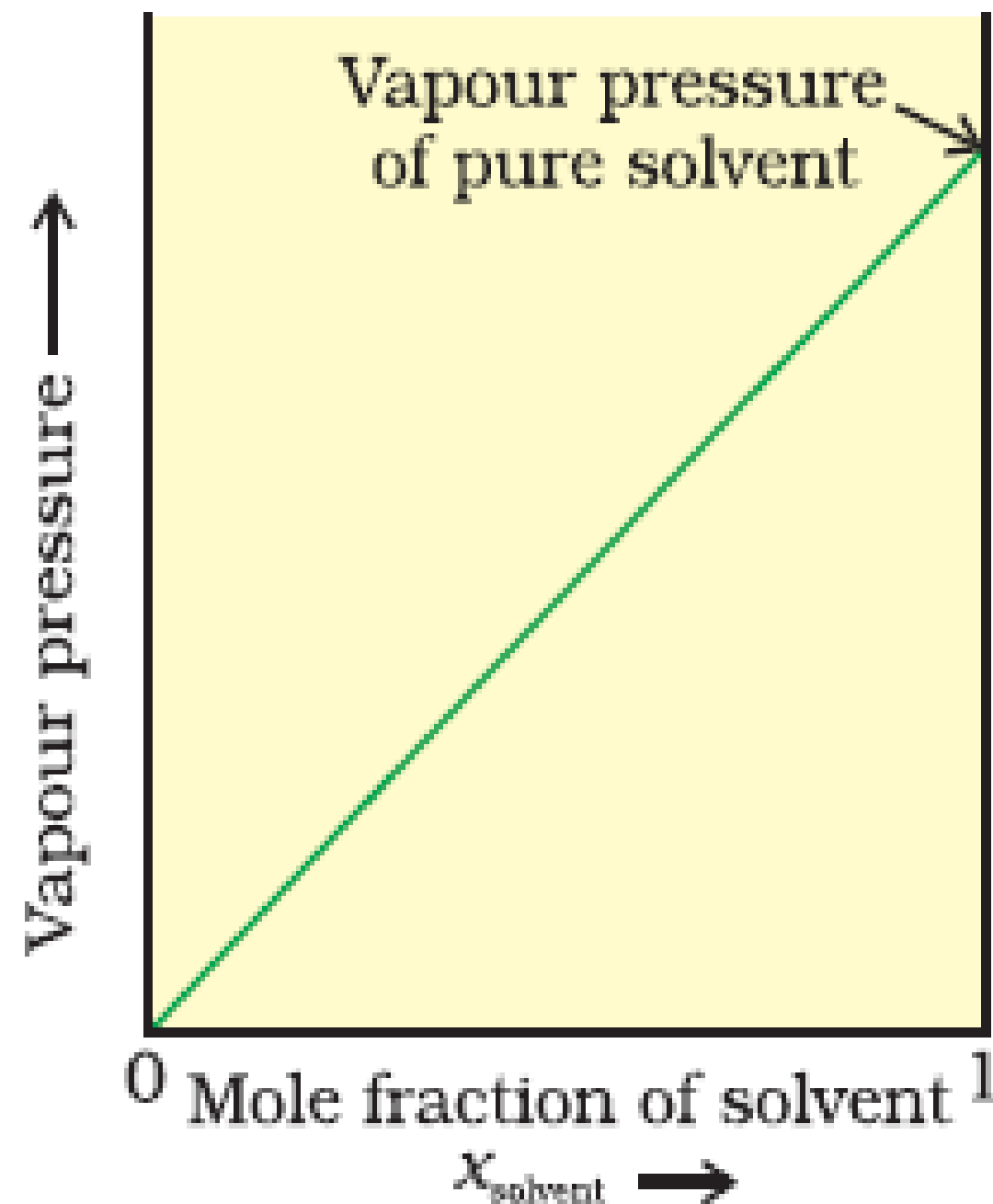
# Vapour Pressure of Solutions of Solids in Liquids

In case of non volatile solute the vapor pressure of a solution is lower than that of the pure solvent



the partial vapour pressure of each volatile component in the solution is directly proportional to its mole fraction.

$$P_1 = X_1 P_1^\circ$$



### **Colligative properties:**

**The properties of dilute solutions containing non-volatile solute which depends on the relative no; of solute and solvent particles but do not depend upon the nature are called colligative properties. The important colligative properties are**

- 1. Relative lowering of vapour pressure**
- 2. Elevation in boiling point**
- 3. Depression in freezing point**
- 4. Osmotic pressure**

# 1. RELATIVE LOWERING OF VAPOUR PRESSURE

Consider a solution containing a non-volatile solute. According to Raoult's law,

$P_s = P_1 + P_2$  but for a non-volatile solute, the solute does not contribute towards the total vapour pressure of the solution

$$P_1 = x_1 P_1^0$$

The reduction in the vapour pressure of solvent ( $\Delta p_1$ ) is given

$$\begin{aligned}\Delta p_1 &= P_1^0 - P_1 = P_1^0 - P_1^0 x_1 \\ &= P_1^0 (1 - x_1)\end{aligned}$$

Knowing that  $x_2 = 1 - x_1$ ,

$$\Delta p_1 = x_2 P_1^0$$

$$\frac{\Delta p_1}{P_1^0} = \frac{P_1^0 - P_1}{P_1^0} = x_2$$

Also

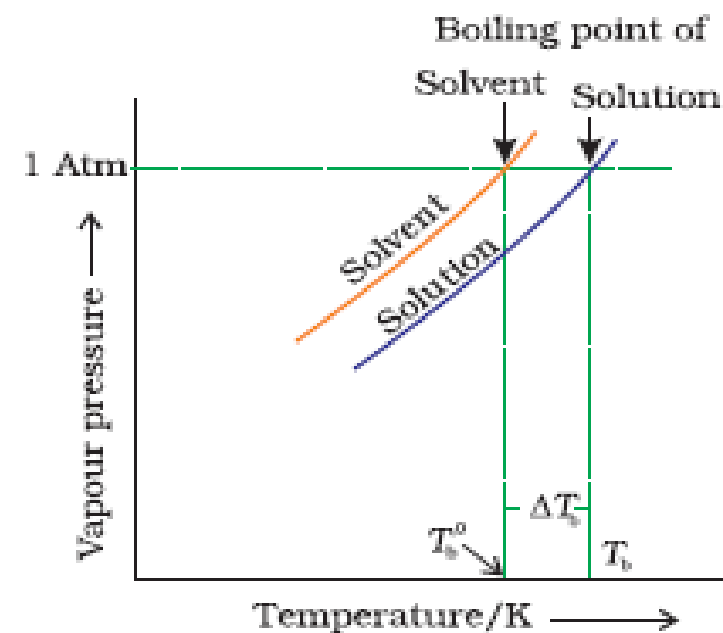
$$\begin{aligned}\frac{P_1^0 - P_1}{P_1^0} &= \chi_B \\ \therefore \frac{P_1^0 - P_1}{P_1^0} &= \frac{w_B / M_B}{w_A / M_A} \\ \therefore \frac{P_1^0 - P_1}{P_1^0} &= \frac{w_B}{M_B} \times \frac{M_A}{w_A}\end{aligned}$$

Using this equation the molar mass of the solute can be calculated.

## 2.ELEVATION IN BOILING POINT

The boiling point of a liquid is the temp at which the V.P of liquid becomes equal to atmospheric pressure. When a non- volatile solute is added to a solvent , the solute Molecules occupy certain surface area. As a result the fraction of the surface covered by solvent molecules gets reduced. Thus the V.P of the solution is less than that of the pure solvent . In order to make this solution boil, its V.P must be increased to atmospheric pressure by raising the temp above boiling point of pure solvent.

Elevation in boiling point is directly proportional to molality of solution



$$\Delta T_b \propto m$$

$$\Delta T_b = K_b \times m$$

$K_b$  is **Ebullioscopic constant** (or) molal elevation constant

$K_b$  is defined as the elevation in boiling point produced when one mole of a solute is dissolved in one Kg of the solvent.

$$\begin{aligned} \Delta T_b &= K_b \times m \\ &= K_b \times \frac{w_B}{M_B} \times \frac{1000}{w_A} \end{aligned}$$



### 3. DEPRESSION IN FREEZING POINT

Freezing point of a solution is defined as the temp: at which the vapour pressure of liquid is equal to the VP of corresponding solid.

When a non-volatile solute is added to a solvent, its V.P decreases. In order to make this solution freeze, its V.P must be decreased by lowering the temp below the freezing point of pure solvent.

$\Delta T_f$  – Depression in freezing point ( $T_o - T_f$ )

Depression in freezing point is directly proportional to molality of solution.

$$\Delta T_f \propto m$$

$$\Delta T_f = K_f \times m$$

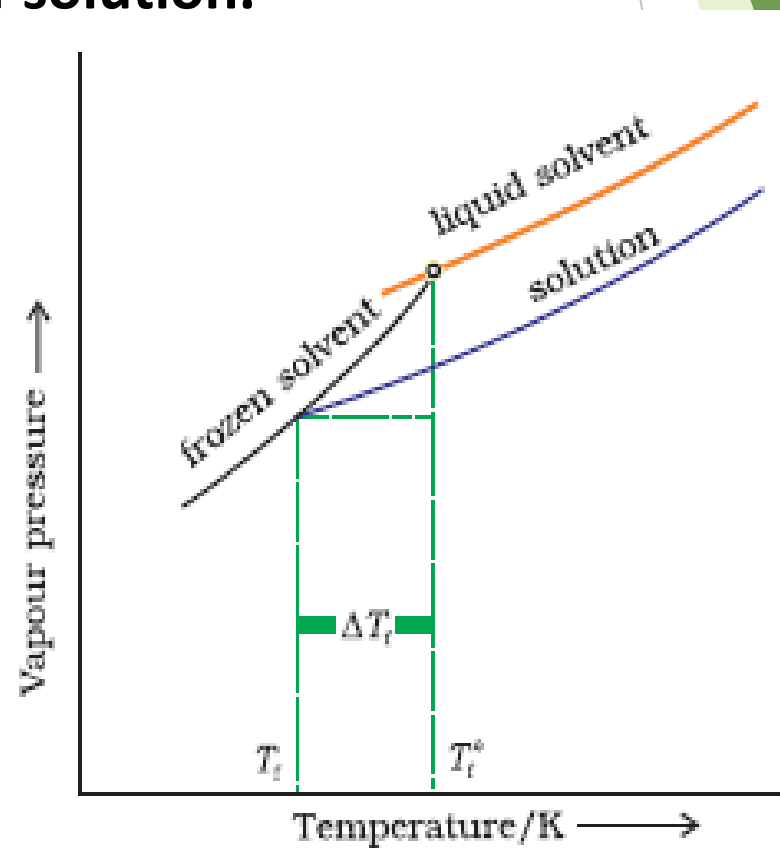
$K_f$  is Cryoscopic constant (or) molal depression constant

$K_f$  is defined as the depression constant when one mole of a solute is dissolved in one Kg of the solvent.

$$\Delta T_f = K_f \times m$$

$$= K_f \times \frac{w_B}{M_B} \times \frac{1000}{w_A}$$

Using  $\Delta T_f$ , molar mass of non-volatile can be calculated.



## **4.OSMOTIC PRESSURE**

**Osmotic pressure is defined as the excess pressure that must be applied to the concentrated solution side to prevent the movement of solvent through semi – permeable membrane. The membrane which allow the passage of solvent molecules through them but do not allow the passage of solute molecules are called semipermeable membranes. Example : Animal bladder, cellophane**

Osmotic pressure  $\pi = CRT$

$$= \frac{n_B RT}{V(L)}$$

$$\pi = \frac{w_B RT}{M_B V(L)}$$

### **Reverse osmosis :**

**If the pressure applied on concentrated solution side is more than osmotic pressure then the solvent particles move from a solution of higher concentration to a region of lower concentration. This is known as reverse osmosis. Reverse osmosis is used for the purification/desalination of sea water.**

**Isotonic solutions** :These are solutions having same osmotic pressure at a given temperature.0.9% solution of NaCl is isotonic with human blood . Hence it is mixed with medicines during intravenous injections .

**Hypertonic solutions** : These are solutions having greater osmotic pressure than that of blood . Eg: salt solution of concentration more than 0.9% are hypertonic compared to blood.

**Hypotonic solutions** : These are solutions having lesser osmotic pressure than that of blood . Eg: salt solution of concentration less than 0.9% are hypotonic compared to blood.

## Abnormal molecular mass

The experimental molar mass determined by colligative properties when there is association or dissociation of solute particles is known as abnormal molecular mass

- If there is association of solute, the no: of particles would decrease and the colligative properties will be lower than expected. Since the colligative properties are inversely proportional to molar mass , the experimental molar mass will be higher than the correct value.
- If there is dissociation of solute, the no: of particles would increase and the colligative properties will be higher than expected. Since the colligative properties are inversely proportional to molar mass , the experimental molar mass will be lower than the correct value.

Van't Hoff factor :It is defined as the ratio of experimental value of a colligative properties to the theoretical value of colligative property.

$$i = \frac{\text{Normal molecular mass}}$$

Observed or abnormal molecular mass

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

Theoretical value of colligative property

$$i = \frac{\text{No: of particles after dissociation or association}}$$

No: of particles before dissociation or association

If solute dissociates  $i > 1$

If solute associates  $i < 1$

For non electrolytes  $i = 1$

**For strong electrolytes** : Van't Hoff factor is equal to the no: of ions produced by the dissociation of one molecule of the electrolyte.



### For weak electrolytes:

For dissociation :

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

$\alpha$  is the degree of dissociation .

$m$  is the no: of ions or particles produced from one molecule.



For association

$$\alpha = \frac{i - 1}{1/m - 1}$$

$\alpha$  is the degree of association . Degree of association is the no: of molecules associated to the total no: of molecules.  $m$  is the no: of particles that associate to form dimer, trimer , tetramer etc.

## Colligative properties for non-ideal solutions

(i) Relative lowering of vapour pressure: 
$$\frac{P_A^0 - P_A}{P_A^0} = i \chi_B$$

$$\frac{P_A^0 - P_A}{P_A^0} = i \frac{w_B}{M_B} \times \frac{M_A}{w_A}$$

(i) Elevation in boiling point :  $\Delta T_b = i \cdot K_b \cdot m$

$$\Delta T_b = i \cdot K_b \cdot \frac{w_B}{M_B} \times \frac{1000}{w_A}$$

(i) Depression in freezing point :  $\Delta T_f = i \cdot K_f \cdot m$

$$\Delta T_f = i \cdot K_f \cdot \frac{w_B}{M_B} \times \frac{1000}{w_A}$$

(i) Osmotic pressure :  $\pi = i \cdot CRT$

$$= i \cdot \frac{w_B RT}{M_B V(L)}$$

Q1. The vapour pressure of pure benzene at a certain temperature is 0.850 bar. A non-volatile, non-electrolyte solid weighing 0.5 g when added to 39.0 g of benzene (molar mass  $78 \text{ g mol}^{-1}$ ). Vapour pressure of the solution, then, is 0.845 bar. What is the molar mass of the solid substance?

Q2. A solution containing 30 g of non-volatile solute exactly in 90 g of water has a vapour pressure of 2.8 kPa at 298 K. Further, 18 g of water is then added to the solution and the new vapour pressure becomes 2.9 kPa at 298 K. Calculate: (i) molar mass of the solute (ii) vapour pressure of water at 298 K.



Q3. 18 g of glucose,  $C_6H_{12}O_6$ , is dissolved in 1 kg of water. At what temperature will water boil at 1.013 bar?  $K_b$  for water is  $0.52 \text{ K kg mol}^{-1}$ .

Q4. 45 g of ethylene glycol ( $C_2H_6O_2$ ) is mixed with 600 g of water. Calculate (a) the freezing point depression and (b) the freezing point of the solution.  $K_f = 1.86 \text{ K kg mol}^{-1}$ .

Q5. Two elements A and B form compounds having formula  $AB_2$  and  $AB_4$ . When dissolved in 20 g of benzene ( $C_6H_6$ ), 1 g of  $AB_2$  lowers the freezing point by 2.3 K whereas 1.0 g of  $AB_4$  lowers it by 1.3 K. The molar depression constant for benzene is  $5.1 \text{ K kg mol}^{-1}$ . Calculate atomic masses of A and B.

Q6.  $200 \text{ cm}^3$  of an aqueous solution of a protein contains  $1.26 \text{ g}$  of the protein. The osmotic pressure of such a solution at  $300 \text{ K}$  is found to be  $2.57 \times 10^{-3} \text{ bar}$ . Calculate the molar mass of the protein.

Q7.  $2 \text{ g}$  of benzoic acid ( $\text{C}_6\text{H}_5\text{COOH}$ ), dissolved in  $25 \text{ g}$  of benzene shows a depression in freezing point equal to  $1.62 \text{ K}$ .  $K_f$  for benzene is  $4.9 \text{ K Kg/mol}$ . What is the percentage association of the acid if it forms dimer in solution.

Q8.  $0.6 \text{ mL}$  of acetic acid ( $\text{CH}_3\text{COOH}$ ), having density  $1.06 \text{ g mL}^{-1}$ , is dissolved in  $1 \text{ litre}$  of water. The depression in freezing point observed for this strength of acid was  $0.0205^\circ\text{C}$ . Calculate the van't Hoff factor and the dissociation constant of acid.  $K_f=1.86 \text{ K Kg/mol}$