# class 12          Chapter -1  Solutions

Solution is a homogeneous mixture of two or more substances in same or different physical phases. The substances forming the solution are called components of the solution. On the basis of number of components a solution of two components is called binary solution.

**Solute and Solvent**

In a binary solution, solvent is the component which is present in large quantity while the other component is known as solute.

**Classification of Solutions**

**(A)** Following types of solutions are seen on the basis of physical state of solute and solvent.

| S.No. | Solute | Solvent | Examples |
|---|---|---|---|
| **Solid solutions** | | | |
| 1 | Solid | Solid | Alloys |
| 2. | Liquid | Solid | Hydrated salts, Amalgam of Hg with Na |
| 3. | Gas | Solid | Dissolved gases in mineral |
| **Liquid solutions** | | | |
| 4. | Solid | Liquid | Salt/sugar solution in water |
| 5 | Liquid | Liquid | Alcohol in water |
| 6 | Gas | Liquid | Aerated drinks, $O_2$ in water |
| **Gaseous solutions** | | | |
| 7. | Solid | Gas | Iodine vapour in air |
| 8 | Liquid | Gas | Water vapour in air |
| 9. | Gas | Gas | Air $(O_2 + N_2)$ |

[if water is used as a solvent, the solution is called aqueous solution and if not, the solution is called non-aqueous solution.]

**(B)** Depending upon the amount of solute dissolved in a solvent we have the following types of solutions:

(i) **Unsaturated solution** A solution in which more solute can be dissolved without raising temperature is called an unsaturated solution.

**(ii)Saturated solution** A solution in which no solute can be dissolved further at a given temperature is called a saturated solution.

**(iii)Supersaturated solution** A solution which contains more solute than that would be necessary to saturate it at a given temperature is called a supersaturated solution.

## Solubility

The maximum amount of a solute that can be dissolved in a given amount of solvent (generally 100 g) at a given temperature is termed as its solubility at that temperature.

The solubility of a solute in a liquid depends upon the following factors:

(i) Nature of the solute
(ii) Nature of the solvent
(iii) Temperature of the solution
(iv) Pressure (in case of gases)

## Henry's Law

The most commonly used form of Henry's law states "the partial pressure (P) of the gas in vapour phase is proportional to the mole fraction (x) of the gas in the solution" and is expressed as

$$p = K_H \cdot x$$

Greater the value of $K_H$, higher the solubility of the gas. The value of $K_H$ decreases with increase in the temperature. Thus, aquatic species are more comfortable in cold water [more dissolved $O_2$] rather than Warm water.

## Applications

1.In manufacture of soft drinks and soda water, $CO_2$ is passed at high pressure to increase its solubility.
2.To minimise the painful effects (bends) accompanying the decompression of deep sea divers. $O_2$diluted with less soluble. He gas is used as breathing gas.
3. At high altitudes, the partial pressure of $O_2$ is less then that at the ground level. This leads to low concentrations of $O_2$ in the blood of climbers which causes „anoxia".

## Concentration of Solutions

The concentration of a solution is defined as the relative amount of solute present in a solution. On the basis of concentration of solution there are two types of solutions.

(i) Dilute solution
(ii) Concentrated solution

**Methods of Expressing Concentration of Solutions**

Various expression for the concentrations of solutions can be summarised as

**(i)Percentage by weight** (w / w %) It is defined as the amount of solute present in 100 g of solution.

w / w % = weight of solute / weight of solution * 100

**(ii)Percentage by volume** (w / V%) It is defined as the weight 01 solute present in 100 mL of solution.

w / V % = weight of solute / weight of solution * 100

or the volume of solute present in 100 mL of solution.

u / V % = volume of solute / volume of solution * 100

**(iii)Mole fraction** (x) It is defined as the ratio of the number of moles of a component to the total number of moles of all the components. For a binary solution, if the number of moles of A and B are $n_A$ and $n_B$ respectively, the mole fraction of A will be

$$\chi_A = \frac{n_A}{n_A + n_B}$$

Similarly, $$\chi_B = \frac{n_B}{n_A + n_B} \qquad \because \chi_A + \chi_B = 1$$

**(iv)Parts per million** (ppm) It is defined as the parts of a component per million parts ($10^6$) of the solution. It is widely used when a solute is present in trace quantities.

ppm = number of parts of the component / total number of parts of all the components * $10^6$

**(v)Molarity** (M) It is the number of moles of solute present in 1L($dm^3$) of the solution.

M = number of moles of solute / volume of solution (L)

M = mass of solute (in gram) * 1000 / mol. wt. of solute x volume of solution (in mL)

Molarity varies with temperature due to change in volume of solution.

[When molarity of a solution is 1 M, it is called a molar solution. 0.1 M solution is called a decimolar solution while 0.5 M solution is known as semi molar solution]

Molarity = Percent by mass * density * 10 / molecular weight

Dilution law, $M_1 V_1 = M_2 V_2$ (for dilution from volume $V_1$ to $V_2$)

For reaction between two reactants, $M_1 V_1 / n_1 = M_2 V_2 / n_2$

where, $n_1$ and $n_2$ arc stoichiometric coefficient in balanced equation.

**(vi)Molality** (m) It is the number of moles of solute per kilogram of the solvent.

Molality = mass of solute in gram * 1000 / mol. wt. of solute * mass of solvent (in g)

Molality is independent of temperature.

[Whcn solvent used is water, a molar (1 M) solution is more concentrated than a molal (1 M) solution.]

**(vii)Normality** (N) The number of gram equivalents of solute present in 1 L of solution.

Normality = number of grams – equivalent of solute / volume of solution in L

Number of gram-equivalents of solute = mass of solute in gram / equivalent weight

[Relationship between normality and molarity N x Eq. weight = M x mol. weight ]

If two solutions of the same solute having volumes and molarities $V_1$, $M_1$ and $V_2$, $M_2$ are mixed, the molarity of the resulting solution is

$$M = \frac{V_1 M_1 + V_2 M_2}{V_1 + V_2}$$

Similarly, Normality $(N) = \dfrac{N_1 V_1 + N_2 V_2}{V_1 + V_2}$

To dilute $V_1$ mL of a solution having molarity $M_1$ to molarity $M_2$ up to the final volume $V_2$ mL, the volume of water added is

$$V_2 - V_1 = \left( \frac{M_1 - M_2}{M_2} \right) V_1.$$

Similarly, $V_2 - V_1 = \left( \dfrac{N_1 - N_2}{N_2} \right) V_1$

(viii) **Formality** (F) It is the number of formula weights of solute present per litre of the solution.

Formality = moles of substance added to solution / volume of solution (in L))

**(ix)Mass fraction** Mass fraction of any component in the solution is the mass of that component divided by the total mass of the solution.

Molality, mole fraction and mass fraction are preferred over molarity, normality, etc., because former involve weights which do not change with temperature.

**(x) Demal** (D) It represents one mole of solute present in 1L of solution at O°C.

**Raoult's Law**

The Raoult''s law states "For a solution of two volatile liquids, the vapour pressure of each liquid in the solution is less than the respective vapour pressure of the pure liquids and the equilibrium partial vapour pressure of the liquid is directly proportional to its mole fraction.

For a solution containing two liquids A and B, the partial vapour pressure of liquid A is

$$p_A \propto \chi_A \quad \text{or} \quad p_A = k\chi_A$$

$$\text{where, } \chi_A = \frac{n_A}{(n_A + n_B)} = \text{mole fraction of liquid } A$$

The proportionality constant is obtained by considering the pure liquid when $\chi_A = 1$ then $k = P°_A$, the vapour pressure of pure liquid, hence

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

Similarly,

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

The total vapour pressure of the solution,

$$P_T = p_A + p_B = p_A^\circ \chi_A + p_B^\circ \chi_B$$
$$= p_A^\circ + (p_B^\circ - p_A^\circ)\chi_B$$

**Konowaloff Rule**

At any fixed temperature, the vapour phase is always richer in the more volatile component as compared to the solution phase. In other words, mole fraction of the more volatile component is always greater in the vapour phase than in the solution phase.

The composition of vapour phase in equilibrium with the solution is determined by the partial pressure of components. If $Y_1$ and $Y_2$ are the
component 1 and 2 respectively in the vapour phase then. using Dalton''s law of partial pressure,

$p_1 = y_1 * P_{total}$

$p_2 = y_2 * P_{total}$

## Ideal Solutions

Those solutions in which solute-solute (B-B) and solvent-solvent (A-A) interactions are almost similar to solvent solute (A-B) interactions are called ideal solutions. These solutions satisfy the following conditions :
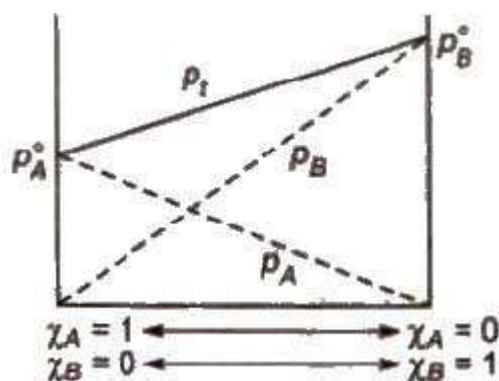


(i) Solution must obey Raoult"s law, i.e.,

$$p_A = p_A^\circ \chi_A, \quad p_B = p_B^\circ \chi_B$$

(ii) $\Delta H_{mix} = 0$ (No energy evolved or absorbed)

(iii) $\Delta V_{mix} = 0$ (No expansion or contraction on mixing)

Some solutions behave like nearly ideal solutions, e.g., benzene + toluene. n-hexane + n-heptane, ethyl iodide + ethyl bromide, chlorobenzene + bromobenzene.

## Non-ideal Solutions

Those solutions which shows deviation from Raoult"s law is called non-ideal solution.
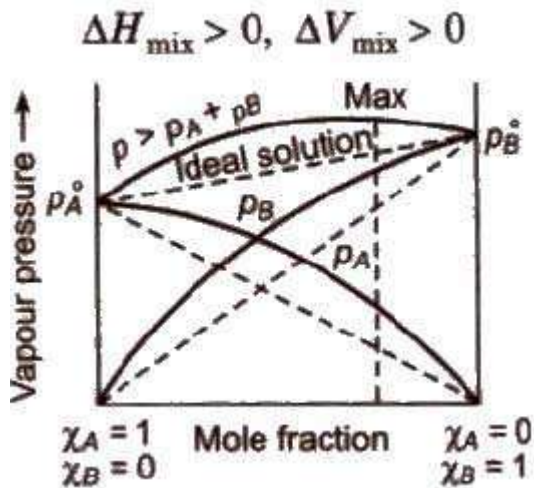
For such solutions,

$\Delta H_{mix} \neq 0$

$\Delta V_{mix} \neq 0$

(a) **Non-ideal solutions showing positive deviation** In such a case, the A – B interactions are weaker than A – A or B – B interactions and the observed vapour pressure of each component and the total vapour pressure are greater than that predicted by Raoult"s law.

$$p_A > p_A^\circ \chi_A, \quad p_B > p_B^\circ \chi_B$$

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

For such solutions

$$\Delta H_{mix} > 0, \quad \Delta V_{mix} > 0$$



**Examples** : Ethanol + water, $CS_2$ + acetone, $CCl_4 + C_6H_6$, $CCl_4 + C_6H_5CH_3$, ethanol + cyclohexane, $CCl_4 + CHCl_3$.
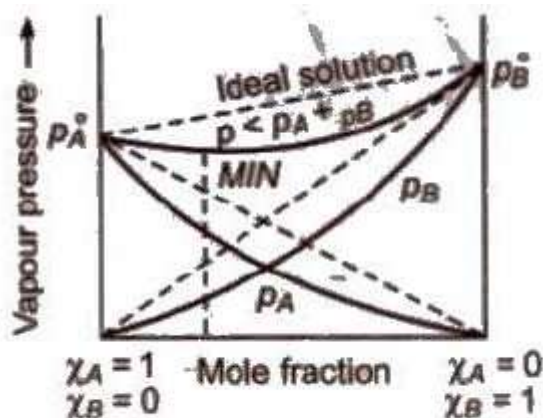
(b) **Non-ideal solution showing negative deviation** In such a case, the A – B interactions are stronger than A – A or B – B interactions and the observed vapour pressure of each component and the total vapour pressure are lesser than that predicted by Raoult's law.

$$p_A < p_A^\circ \chi_A, \quad p_B < p_B^\circ \chi_B$$

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

For such solutions,

$$\Delta H_{mix} < 0, \quad \Delta V_{mix} < 0$$

**Non-ideal solution showing negative deviation**

Examples : $CHCl_3 + CH_3COCH_3$, $CHCl_3 + C_6H_6$, $H_2O + HCl$, $H_2O + HNO_3$, methanol + acetic acid.

## Azeotropic Mixture

A mixture of two liquids which boils at a particular temperature like a pure liquid and distils over in the same composition is known as constant boiling mixtures. These are formed by non-ideal solutions.

**(i)Minimum boiling azeotropes** are formed by those liquid pairs which show positive deviation from ideal behaviour. Such azeotropes have boiling points lower than either of the components, e.g., $C_2H_5OH$ (95.57%) + $H_2O$ (4.43%)(by mass).

**(ii)Maximum boiling azeotropes** are formed by those liquid pain; which show negative deviation from ideal behaviour. Such azeotropes have boiling points higher than either of the components. e.g., $H_2O(20.22O\%)+ HCl$ (79.78%] by mass.

## Colligative Properties

[Colligatil1e : from Latin. = Co mean „together'; ligare means „to bind".]

Colligative properties are those properties which depends only upon the number of solute particles in a solution irrespective of their nature.

## Relative Lowering of Vapour Pressure

It is the ratio of lowering in vapour pressure to vapour pressure of pure solvent. The relative lowering in vapour pressure of solution containing a nonvolatile solute is equal to the mole fraction of solute in the solution.

$$\frac{p_A^\circ - p_A}{p_A^\circ} = \chi_B$$

where, $\dfrac{p_A^\circ - p_A}{p_A^\circ}$ = relative lowering of vapour pressure

$$\frac{p_A^\circ - p_A}{p_A^\circ} = \frac{n_B}{n_A + n_B}$$

for dilute solutions, $n_B \ll n_A$. Hence,

$$\frac{p_A^\circ - p_A}{p_A^\circ} = \frac{n_B}{n_A}$$

or

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

$$M_B = \frac{W_B}{W_A} \times M_A \times \frac{p_A}{(p_A^\circ - p_A)}$$

Above expression is used to find the molecular weight of an unknown solute dissolved in a given solvent. Where, $W_B$ and $W_A$ = mass of Solute and solvent respectively. $M_B$ and $M_A$ = molecular weight of solute and solvent respectively.

Ostwald and Walker method is used to determine the relative lowering of vapour pressure.

**Elevation in Boiling Point** ($\Delta T_b$)

Boiling point of a liquid is the temperature at which its vapour pressure becomes equal to the atmospheric pressure. As the vapour pressure of a solution containing a nonvolatile solute is lower than that of the pure solvent, it boiling point will be higher than that of the pure solvent as shown in figure. The increase in boiling point is known as elevation in boiling point, $\Delta T_b$
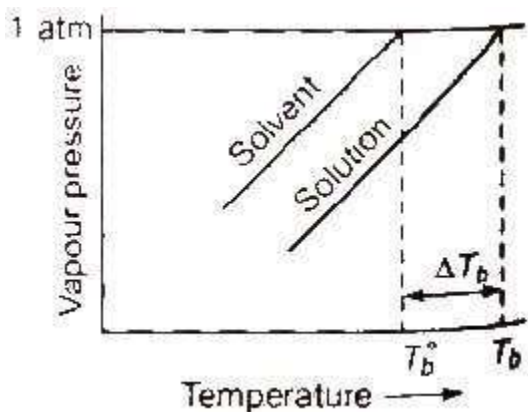
$\Delta T_b = T_b - T^\circ_b$

$\Delta T_b = K_b \, m$ (where; m = molality)

$K_b$ is molal elevation constant or ebullioscopic constant. Molecular mass of solute can be calculated as

$$\Delta T_b = \frac{K_b \cdot W_B \times 1000}{M_B \times W_A}$$

$$M_B = K_b \cdot \frac{W_B}{W_A} \times \frac{1000}{\Delta T_b}$$

where, $W_B$ and $W_A$ = mass of solute and solvent respectively.

$K_b$ has units of K / m or K kg mol$^{-1}$, for water, $K_b = 0.52$ K kg mol$^{-1}$

The boiling point elevation of a solution is determined by

(i) Landsberger''s method

(ii) Cottrell''s method

**Depression in Freezing Point** ($\Delta T_f$)

Freezing point of a liquid is the temperature at which vapour pressure of the solvent in its liquid and solid phase become equal. As we know that vapour pressure of solution containing non-volatile solute is lower than that of pure solvent, solid form gets separated out at a lower temperature as shown in the figure.
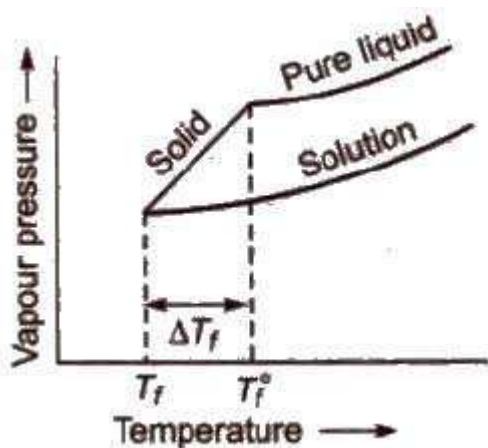
This decrease in freezing point of a liquid is known as depression in freezing point.

Depression in freezing point $(\Delta T_f) = T°_f - T_f$

$$\Delta T_f = K_f \cdot m = K_f \frac{W_B}{M_B} \times \frac{1000}{W_A}$$

To find molecular mass of solute,

$$M_B = \frac{K_f \cdot W_B \times 1000}{\Delta T_f \cdot W_A}$$

where, $K_f$ is molal depression constant or cryoscopic constant.

$K_f$ has units of K / m or K kg mol$^{-1}$.

Ethylene glycol is usually added to water in the radiator to lower its freezing point. It is called antifreeze solution.

[Common salt (NaCI) and anhydrous CaC12 are used to clear snow on the roads because they depress the freezing point of water. The freezing point depression is determined by Beckmann method or Rast method.]

Calculations of molal elevation constant ($K_b$) and molal depression constant ($K_f$)

$$K_b = \frac{M_A R(T_b^\circ)^2}{\Delta H_v}, \quad K_f = \frac{M_A \cdot R(T_f^\circ)^2}{\Delta H_f}$$

$T_b^\circ$ = boiling point of solvent

$T_f^\circ$ = freezing point of solvent

$\Delta H_f$ = molar enthalpy of fusion

$\Delta H_n$ = molar enthalpy of vaporisation

**Osmotic Pressure** $(\pi)$

**Osmosis** is the phenomenon of spontaneous flow of the solvent molecules through a semipermeable membrane from pure solvent to solution or from a dilute solution to concentrated solution. It was first observed by Abbe Nollet.

Some natural semipermeable membranes are animal bladder, cell membrane etc.

$CU_2[Fe(CN)_6]$ is an artificial semipermeable membrane which does not work in non-aqueous solutions as it dissolves in them.

Osmosis may be

**(i)Exosmosis** It is outward flow of water or solvent from a cell through semipermeable membrane.
**(ii)Endosmosis** It is inward flow of water or solvent from a cell through a semipermeable membrane.

The hydrostatic pressure developed on the solution which just prevents the osmosis of pure solvent into the solution through a semipermeable membrane is called **osmotic pressure.**

$$\text{Osmotic pressure } (\pi) = RCT; \left( C = \frac{n_B}{V} = \frac{W_B}{M_B V} \right)$$

$$M_B = \frac{W_B RT}{\pi V}$$

$$\pi = \frac{dRT}{M_B} ; \left( d = \frac{W_B}{V} \right)$$

where, d = density, R = solution constant,

T = temperature, $M_B$ = molar mass of solute

Osmotic pressure can be determined by anyone of the method listed below

(i) Pfeffer''s method
(ii) Berkeley and Hartley''s method (very good method)
(iii) Morse and Frazer''s method

On the basis of osmotic pressure, -the solution can be

**(i)Hypertonic solution** A solution is called hypertonic if its osmotic pressure is higher than that of the solution from which it is separated by a semipermeable membrane.

When a plant cell is placed in a hypertonic solution, the fluid from the plant cell comes out and cell shrinks, this phenomenon is called plasmolysis.

**(ii)Hypotonic solution** A solution is called hypotonic if its osmotic pressure is lower than that of the solution from which it is separated by a semipermeable membrane.

**(iii)Isotonic solution** Two solutions are called isotonic if they exert the same osmotic pressure. These solutions have same molar concentration. $0.91\%$ solution of pure NaCl is isotonic with human RBC''s.

Two solutions are isotonic if they have the same molar concentration, e.g., if x % solution of X is isotonic with y % solution of Y, this means molar concentration of X = Molar concentration of Y

$$\frac{x}{100} \times \frac{1000}{M_X} = \frac{y \times 1000}{100 \times M_Y} \implies \frac{x}{M_X} = \frac{y}{M_Y}$$

Osmotic pressure method is the best method for determining the molecular masses of polymers since observed value of any other colligative property is too small to be measured with reasonable accuracy.

**Reverse osmosis** When the external pressure applied on the solution is more than osmotic pressure, the solvent flows from the solution to the pure solvent, I which is called reverse osmosis. Desalination of sea water is done by reverse Osmosis.

**Abnormal Molecular Masses**

In some cases, observed colligative properties deviate from their normal calculated values due to association or dissociation of molecules. As we know,

Colligative property $\propto 1 / M_B$

lienee, higher and lower values of molar mass is observed in case of association and dissociation respectively, e.g., in benzene, acetic acid gets associated, so, its observed

molecular mass is 120. Similarly KCI undergoes dissociation in aqueous solution, so its observed molecular mass is 37.25.

These observed values are corrected by multiplying with van''t Hoff factor (i).

**van't Hoff Factor** (i)

It is the ratio of observed value of colligative property to the calculated value of colligative property.

i = observed value of colligative property / calculated value of colligative property

or i = normal molecular mass / observed molecular mass

or i = number of particles after association or dissociation / number of particles initially

So to correct the observed value of molar mass, van''t Hoff factor (i) must be included in different expressions for colligative properties.
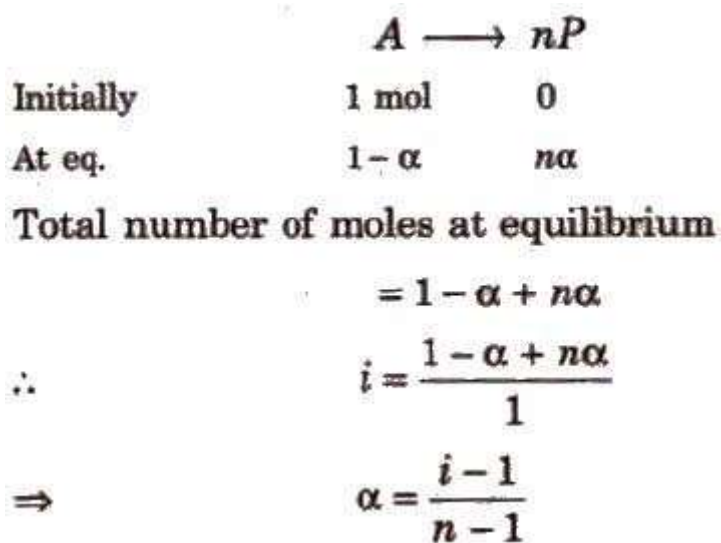
$$\Delta T_b = i\,K_b \cdot m$$
$$\Delta T_f = iK_f \cdot m$$
$$\pi = i\,CRT$$
$$\frac{\Delta p}{p_A^\circ} = ix_B$$

**Degree of Dissociation (α) and van't Hoff Factor (i)**

(i) If one molecule of a substance gets dissociated into n particles or molecules and α is the degree of dissociation then

$$A \longrightarrow nP$$

| | | |
|---|---|---|
| Initially | 1 mol | 0 |
| At eq. | $1 - \alpha$ | $n\alpha$ |

Total number of moles at equilibrium

$$= 1 - \alpha + n\alpha$$

$$\therefore \quad i = \frac{1 - \alpha + n\alpha}{1}$$

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

**Degree of Association ($\alpha$) and van't Hoff Factor (i)**

If n molecules of a substance A associate to form $A_n$ and $\alpha$ is the degree of association then

$$nA \longrightarrow A_n$$

Initially

1 mol — 0

At equilibrium

$1 - \alpha$ — $\dfrac{\alpha}{n}$

Total number of moles at equilibrium

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

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

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

van't Hoff factor (i) > 1 for solutes undergoing dissociation and it is < 1 for solutes undergoing association.

Chapter 1

Solutions

$$Molality = \frac{\text{Number of moles of solute}}{\text{Mass of solvent in kilograms}}$$

$$Molarity = \frac{\text{Number of moles of solute}}{\text{Volume of solution in litres}}$$

$$P = KHx$$

The difference in boiling points of solution $T_b$ and pure solvent $T_b^0$ is called elevation in boiling point $\Delta T = T_b - T_b^0$

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

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

$$\frac{p_1^0 - p}{p_1^0} = i . \frac{n_2}{n_1}$$

$$\Delta T_b = i . \frac{k_b \times 1000 \times w_2}{M_2 \times w_1}$$

$$\Delta T_f = i . \frac{k_f \times 1000 \times w_2}{M_2 \times w_1}$$

- **Solutions:** Solutions are the homogeneous mixtures of two or more than two components.
- **Binary solution:** A solution having two components is called a binary solution.
- Components of a binary solution.

  It includes solute and solvent.

  1. When the solvent is in solid state, solution is called solid solution.
  2. When the solvent is in liquid state, solution is called liquid solution.
  3. When the solvent is in gaseous state, solution is called gaseous solution.

- **Concentration:** It is the amount of solute in given amount of solution.
- **Mass by volume percentage (w/v):** Mass of the solute dissolved in 100 mL of solution.
- **Molality** (m) is the number of moles of solute present in 1kg of solvent.
- **Molarity** (M) is the number of moles of solute present in 1L of solution.
- **Normality** is the number of gram equivalent of solute dissolved per litre of solution.
- **Solubility:** It is the maximum amount that can be dissolved in a specified amount of solvent at a specified temperature.
- **Saturated solution:** It is a solution in which no more solute can be dissolved at the same temperature and pressure.
- In a nearly saturated solution if dissolution process is an endothermic process, solubility increases with increase in temperature.
- In a nearly saturated solution if dissolution process is an exothermic process, solubility decreases with increase in temperature.
- **Henry's Law:** It states "at a constant temperature the solubility of gas in a liquid is directly proportional to the pressure of gas". In other words, "the partial pressure of gas in vapour phase is proportional to the mole fraction of the gas in the solution".
- When a non-volatile solute is dissolved in a volatile solvent, the vapour pressure of solution is less than that of pure solvent.
- **Raoult's law:** It states that "for a solution of volatile liquids the partial vapour pressure of each component in the solution is directly proportional to its mole fraction".

$$p_1 = p^0{}_1 X_1 ; p_2 = p^0{}_2 X_2$$

- Using Dalton's law of partial pressure the total pressure of solution is calculated.
$$p_{total} = p^0{}_1 + (p_2 - p^0{}_1) X_2$$
- Comparison of Raoult' law and Henry's law: It is observed that the partial pressure of volatile component or gas is directly proportional to its mole fraction in solution. In case of Henry's Law the proportionality constant is KH and it is different from p10 which is partial pressure of pure component. Raoult's Law becomes a special case of Henry's Law when KH becomes equal to p10 in Henry's law.
- Classification of liquid–liquid solutions: It can be classified into ideal and non-ideal solutions on basis of Raoult's Law.

- Ideal solutions:
  1. The solutions that obey Raoult's Law over the entire range of concentrations are known as ideal solutions.
  2. $\Delta_{mix}H = 0$ and $\Delta_{mix}V = 0$
  3. The intermolecular attractive forces between solute molecules and solvent molecules are nearly equal to those present between solute and solvent molecules i.e. A-A and B-B interactions are nearly equal to those between A-B.
- Non-ideal solutions:
  1. When a solution does not obey Raoult's Law over the entire range of concentration, then it is called non-ideal solution.
  2. $\Delta_{mix}H \neq 0$ and $\Delta_{mix}V \neq 0$
  3. The intermolecular attractive forces between solute molecules and solvent molecules are not equal to those present between solute and solvent molecules i.e. A-A and B-B interactions are not equal to those between A-B
- Types of non- ideal solutions:
  1. Non ideal solution showing positive deviation
  2. Non ideal solution showing negative deviation
- Non ideal solution showing positive deviation
  3. The vapour pressure of a solution is higher than that predicted by Raoult's Law.
  4. The intermolecular attractive forces between solute-solvent molecules are weaker than those between solute-solute and solvent-solvent molecules i.e., A-B < A-A and B-B interactions.
- Non ideal solution showing negative deviation
  1. The vapour pressure of a solution is lower than that predicted by Raoult's Law.
  2. The intermolecular attractive forces between solute-solvent molecules are stronger than those between solute-solute and solvent-solvent molecules i.e. A-B > A-A and B-B interactions.
- Azeotopes: These are binary mixtures having same composition in liquid and vapour phase and boil at constant temperature. Liquids forming azeotrope cannot be separated by fractional distillation.
- Types of azeotropes: There are two types of azeotropes namely,
  1. Minimum boiling azeotrope
  2. Maximum boiling azeotrope

- The solutions which show a large positive deviation from Raoult's law form minimum boiling azeotrope at a specific composition.
- The solutions that show large negative deviation from Raoult's law form maximum boiling azeotrope at a specific composition.
- Colligative properties: The properties of solution which depends on only the number of solute particles but not on the nature of solute are called colligative properties.
- Types of colligative properties: There are four colligative properties namely,
    1. Relative lowering of vapour pressure
    2. Elevation of boiling point
    3. Depression of freezing point
    4. Osmotic pressure
- Relative lowering of vapour pressure: The difference in the vapour pressure of pure solvent $p_1^0$ and solution $p_1$ represents lowering in vapour pressure $(p_1^0 - p_1)$.
- Relative lowering of vapour pressure: Dividing lowering in vapour pressure by vapour pressure of pure solvent is called relative lowering of vapour pressure $\left(\frac{p_1^0 - p_1}{p_1^0}\right)$
- Relative lowering of vapour pressure is directly proportional to mole fraction of solute. Hence it is a colligative property.
- Elevation of boiling point: $\left(\frac{p_1^0 - p_1}{p_1^0}\right) = X_2$
- For a dilute solution elevation of boiling point is directly proportional to molal concentration of the solute in solution. Hence it is a colligative property.
- Depression of freezing point: The lowering of vapour pressure of solution causes a lowering of freezing point compared to that of pure solvent. The difference in freezing point of the pure solvent $T_f^0$ and solution $T_f$ is called the depression in freezing point.

$$\Delta T = T_f^0 - T_f$$

- For a dilute solution depression in freezing point is a colligative property because it is directly proportional to molal concentration of solute.
- Osmosis: The phenomenon of flow of solvent molecules through a semi permeable membrane from pure solvent to solution is called osmosis.
- Osmotic pressure: The excess pressure that must be applied to solution to prevent the

passage of solvent into solution through a semipermeable membrane is called osmotic pressure.

- Osmotic pressure is a colligative property as it depends on the number of solute particles and not on their identity.
- For a dilute solution, osmotic pressure ($\pi$) is directly proportional to the molarity (C) of the solution i.e. $\pi$ = CRT
- Osmotic pressure can also be used to determine the molar mass of solute using the equation $M_2 = \dfrac{w_2 RT}{\pi V}$
- Isotonic solution: Two solutions having same osmotic pressure at a given temperature are called isotonic solution.
- Hypertonic solution: If a solution has more osmotic pressure than other solution it is called hypertonic solution.
- Hypotonic solution: If a solution has less osmotic pressure than other solution it is called hypotonic solution.
- Reverse osmosis: The process of movement of solvent through a semipermeable membrane from the solution to the pure solvent by applyingexcess pressure on the solution side is called reverse osmosis.
- Colligative properties help in calculation of molar mass of solutes.
- Abnormal molar mass: Molar mass that is either lower or higher than expected or normalmolar mass is called as abnormal molar mass.
- Van't Hoff factor: Van't Hoff factor (i)accounts for the extent of dissociation or association.

$$i = \dfrac{\text{Normal molar mass}}{\text{Abnormal molar mass}}$$

$$= \dfrac{\text{Observed collogative property}}{\text{Calculated collogative property}}$$

$$= \dfrac{\text{Total number of moles of particles after association / dissociation}}{\text{Total number of moles of particles before association / dissociation}}$$

- Value of i is less than unity in case solute undergo association and the value of i is greater than unity in case solute undergo dissociation.
- Inclusion of van't Hoff factor modifies the equations for colligative properties as: