1. SOLUTIONS

STUDY NOTES

1.0 In this first chapter **Solutions** we are going to learn (i) Types of Solutions (ii) Concentrations of solutions (iii) Solubility of a gas in a liquid - **Henry's law** (iv) Vapour pressure of liquid solutions - **Raoult's law** (v) Ideal and non -ideal solutions **(vi) Colligative properties:** RLVP, ΔT_b, ΔT_f , **Osmotic pressure(Π)** (vii) Abnormal molar masses - **Van't Hoff factor (i)**

1.1 Solutions [Solvent + Solute $\rightleftharpoons$ Solution]

i) Solution: A **homogeneous mixture** of two or more substances is called a solution.

Ex: Glucose water is a **solution.**

ii) Solvent: The **larger** component of a binary solution is called **solvent.**

Ex: Water is the **solvent** in the glucose water.

iii) Solute: The **lesser** component of a binary solution is called **solute.**

Ex: Glucose is the **solute** in the glucose water.

iv) Component: Any substance present in the solution (including solutes & solvent)

v) Types of solutions : i) Gaseous solutions ii) Liquid solutions iii) Solid solutions

1.2. Concentration of solution:

Composition of a solution is expressed interms of quantity (mass, volume, moles) of the component (solute) present in the solution.

Types of concentrations: i) Mass percentage (w/w) (ii) Volume percentage (V/V) iii) Mass by volume percentage (w/V) iv) Parts per million (ppm) v) Mole fraction (χ) (vi) Molarity (M) (vii) Molality (m)

i) Mass percentage (w/w): Mass of the component (in g) present in 100g of solution.

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

ii) Volume percentage (V/V): Volume of the component (in mL) present in 100 mL of solution.

$$\text{Volume \% (V/V) of a component} = \frac{\text{Volume of the component}}{\text{Total volume of the solution}} \times 100$$

iii) Mass by volume % (w/V): The mass of solute (in g) present in 100 mL of solution.

$$\text{Mass by volume \% (w/V) of solute} = \frac{\text{Mass of the solute in the solution}}{\text{Total volume of the solution}} \times 100$$

iv) Parts per million(ppm): When a component solute is present in very small trace of quantities, then concentration of the solution is expressed in parts per million.

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

v) Mole Fraction(χ): In a binary mixture, if the number of moles of components A,B are and n_A and n_B then

(i) Mole fraction of A is $\chi_A = \frac{n_A}{n_A + n_B}$ (ii) Mole fraction of B is $\chi_B = \frac{n_B}{n_A + n_B}$

Note: Sum of mole fractions of solvent A and solute B is $\chi_A + \chi_B = 1$

vi) Molarity (M): The number of moles of the solute dissolved per litre of the solution

$$\text{Molarity(M)} = \frac{\text{No.of moles of solute (n)}}{\text{Volume of solution(L)}} = \frac{\text{wt of solute}}{\text{GMM of solute}} \times \frac{1000}{\text{V(mL)}}$$

Formula: $M = \frac{n}{V(L)} = \frac{w}{GMM} \times \frac{1}{V(L)} = \frac{w}{GMM} \times \frac{1000}{V(mL)}$

Tip: Molarity depends on Volume (hence, Temperature dependent - Charle's law)

vii) Molality(m): The number of moles of the solute present in per Kg of the solvent.

$$\text{Molality(m)} = \frac{\text{No.of moles of solute}}{\text{Mass of solvent (in Kg)}} = \frac{\text{wt. of solute}}{\text{GMM of solute}} \times \frac{1000}{\text{Wt.of solvent}}$$

Formula: $m = \frac{w}{GMM} \times \frac{1000}{W(g)} = \frac{w}{GMM} \times \frac{1}{W(Kg)}$; w= wt. of solute W= wt. of solvent.

Tip: Molality depends on Mass (hence, Temperature independent)

1.3 Solubility of a substance: It is the maximum amount of the substance that can be dissolved in a solvent. It depends upon the nature of solute and solvent, as well as temperature and pressure.

1.3.1 Solubility of a Solid in liquid: It depends on the temperature - Le Chatelier's principle.

1.3.2 Solubility of a Gas in liquid: It depends on the pressure - Henry's law

Solubility of a gas in a liquid is directly proportional to the partial pressure of the gas present above the surface of the liquid, at constant temperature.

Henry's law: The partial pressure (p) of a gas in vapour phase is directly proportional to the mole fraction (x) of the gas in the solution.

Partial pressure of the gas, $\mathbf{p=K_H x}$.

Here K_H is Henry's law constant, x is mole fraction of the gas in the solution.

1.4 Vapour pressure of Liquid - Liquid solutions:

Raoult's law: It states that for a solution of volatile liquids, the partial vapour pressure of each component of the solution is directly proportional to its mole fraction.

Thus $p_1 \propto x_1 \Rightarrow p_1 = p_1^0 x_1$. Here p_1 is the partial vapour pressure of pure component,

p_1^0 is vapour pressure of pure component 1 at the same temperature, x_1 is mole fraction of component 1.

Also, from the Dalton's law of partial pressures, we have $p_{total} = p_1 + p_2 = x_1 p_1^0 + x_2 p_2^0$

1.5 Ideal & Non-ideal Solutions:

The solutions which obey Raoult's law over the entire range of concentration are known as ideal solutions . **Ex:** C_6H_6+ $C_6H_5CH_3$

The solutions which don't obey Raoult's law are known as Non-ideal solutions.

The vapour pressure of such Non-ideal solutions will be either higher (positive deviation) or lower (negative deviation) than that predicted by Raoult's law.

1.6 Colligative Properties (RLVP, ΔT_b, ΔT_f, Osmotic pressure):

Colligative properties describe how a solute changes a solvents behaviour .

These properties depend on the number of solute particles and are related to vapour pressure.

The RLVP is a colligative property of a solution which describes that when a Non-volatile solute (like sugar or salt) is added to a solution then there is a decrease in the vapour pressure of the solvent compared to the vapour pressure of the pure solvent.

(a) RLVP: Vapour pressure of solvent in a solution is always less than that of pure solvent and it depends only on the concentration of the solute particles.

LVP is the difference between the vapour pressure of pure solvent p_1^o an vapour pressure of solution p_1 in pure state. Thus LVP = $\Delta p_1 = p_1^o - p_1$

Here, p_1^o is the VP of solvent or component 1 and p_1 is the VP of the solution .

Relative lowering of vapour pressure $\text{RLVP} = \dfrac{\Delta p_1}{p_1^o} = \dfrac{p_1^o - p_1}{p_1^o}$

According to Raoult's law, the RLVP of a solution containing a Non-volatile solute is equal to the mole fraction of the solute.

If n_1= no.of moles of solvent, n_2= no.of moles of solute then

$\text{RLVP} = \dfrac{p_1^o - p_1}{p_1^o} = \chi_2 = \dfrac{n_2}{n_1 + n_2} = \dfrac{n_2}{n_1}$ For dil. solutions ($n_2 <<<<< n_1$) . Hence $n_1 + n_2 = n_1$

$\therefore \text{RLVP} = \dfrac{p_1^o - p_1}{p_1^o} = \dfrac{n_2}{n_1} = \dfrac{w_2 / M_2}{w_1 / M_1} = \dfrac{w_2}{M_2} \times \dfrac{M_1}{w_1} \Rightarrow \boxed{M_2 = \left(\dfrac{p_1^o}{p_1^o - p_1}\right)\left(\dfrac{M_1 w_2}{w_1}\right)}$

Here, w_1, M_1 are mass & molar mass of the solvent.

w_2, M_2 are mass & molar mass of the solute.

b) Elevation of boiling point (ΔT_b): It is a colligative property. The boiling point of a solution higher than that of a pure solvent, when a non-volatile solution is added, causing reduction in vapour pressure.

For a dilute solution, elevation of boiling point $\boxed{\Delta T_b = K_b m}$

Here K_b is Boiling point elevation constant (or) Ebullioscopic constant, m is molality,

c) Depression of freezing point (ΔT_f): It is the decrease in the freezing point of a solvent caused by adding a non-volatile solute, making the solution freeze at a lower temperature than pure solvent.

For a dilute solution, **depression of freezing point,** $\boxed{\Delta T_f = K_f m}$

Here K_f is Freezing point depression constant or Cryoscopic constant, m is molality .

d) Osmosis- Osmotic pressure(Π) :

The spontaneous flow of solvent particles through a semi permeable membrane in a solution is called **osmosis.** It happens from low concentration to high concentration.

The pressure required to stop osmosis is called **Osmotic pressure.**

Osmotic pressure **(Π)** is directly proportional to the molar concentration(molarity) C of the solution and absolute temperature (T).

Van't Hoff Equation: Osmotic pressure $\Pi = iCRT$, i= 1 for non-electrolytes.

$$\Pi = CRT = \left(\frac{n}{V}\right)RT \Rightarrow \Pi V = nRT = \left(\frac{w}{M}\right)RT \Rightarrow M = \frac{wRT}{\Pi V} = \text{Molar mass of the solute.}$$

Terminology: Π = Osmotic pressure, C= molarity= n/V, R = Universal gas constant T= Absolute temperature in Kelvin, n= no.of moles of solute, V= Volume of solution in liters i(Van't Hoff factor) represents the no.of particles the solution dissociates into (i>1). (or) associates into (i<1) (or) do not interact (i=1).

(e) Reverse osmosis & Water purification:

The direction of osmosis can be reversed if a pressure larger than the osmotic pressure is applied to the solution side. This phenomenon is called reverse osmosis.

Reverse osmosis is used in desalination of sea water. (Purification of water)

(f) Abnormal molar masses - Van't Hoff factor (i):

Molar mass that is either lower or higher than the expected or normal value is called as abnormal molar mass.

Dissociation (i>1) is the splitting of the molecules of a solute in to smaller ions.

Here molar mass is lesser than expected (abnormal decrease). **Ex:** $NaCl \rightarrow Na^+ + Cl^-$

Association (i<1) is the joining of solute particles in to larger molecules.

Here molar mass is higher than expected (abnormal increase). **Ex:** Acetic acid in benzene

Van't Hoff factor **i** is a measure of the affect of solute on a solution's colligative property.

The Van't Hoff factor i accounts for dissociation or association of solute.

Van't Hoff factor is defined as ratio of normal molar mass to experimentally determined molar mass (or) as the ratio of observed colligative property to the calculated colligative property.

$$i = \frac{\text{Normal molar mass}}{\text{Abnormal molar mass}} = \frac{\text{No.of particles after dissociation or association}}{\text{No.of partciles before dissociation or association}}$$

☛ **Degree of dissociation,** $\alpha = \frac{i-1}{n-1}$; Degree of association, $\alpha = \frac{1-i}{1-(1/n)}$

n= no. of moles of solute.

Textual Notation on Symbols

For solvents, use the suffix 1:

w_1= Mass of solvent, M_1= Molar mass of solvent, χ_1 or χ_A = Mole fraction of solvent

For solute, use the suffix 2:

w_2= Mass of solute, M_2= Molar mass of solute, χ_2 or χ_B = Mole fraction of solute

☛ Any other valid/ habituated notation can also be followed.

FORMULA BOX

1. **Mass % (w/w)** means w grams of solute is present in 100g of solution.

2. **Mass by volume % (w/V)** means w grams of solute is present in 100mL of solution.

3. **No.of moles (n)** $= \frac{\text{weight}}{\text{Gram molar mass}} = \frac{w}{GMM}$

4. **Mole fraction** of A is $\chi_A = \frac{n_A}{n_A + n_B}$ (ii) Mole fraction of B is $\chi_B = \frac{n_B}{n_A + n_B}$

   Note: $\chi_A + \chi_B = 1$ **(or)** $\chi_B = 1 - \chi_A$

5. **Molarity(M)** $= \frac{\text{No.of moles of solute}}{\text{Volume of solution(L)}}$ (or) $M = \frac{n}{V} = n \times \frac{1000}{V(mL)} = \frac{w}{GMM} \times \frac{1000}{V(mL)}$

6. **Molality(m)** $= \frac{\text{No.of moles of solute}}{\text{Mass of solvent (Kg)}}$ (or) $m = \frac{n}{W} = n \times \frac{1000}{W(g)} = \frac{w}{GMM} \times \frac{1000}{W(g)}$

7. **Henry's law:** Partial pressure of the gas, $\mathbf{p = K_H x}$.

8. **Raoult's law:** partial vapour pressure in a liquid $p_1 = p_1^0 x_1$

9. **Relative lowering of vapour pressure** $RLVP = \frac{\Delta p_1}{p_1^o} = \frac{p_1^o - p_1}{p_1^o}$

10. $RLVP = \frac{p_1^o - p_1}{p_1^o} = \chi_2 = \frac{n_2}{n_1 + n_2} = \frac{n_2}{n_1}$ For dil. solutions $(n_2 <<<<< n_1)$.

11. RLVP for a dilute solution $= \frac{p_1^o - p_1}{p_1^o} = \frac{n_2}{n_1} = \frac{w_2 / M_2}{w_1 / M_1} = \frac{w_2}{M_2} \times \frac{M_1}{w_1}$

12. Molar mass of solute $\mathbf{M_2 = \left(\frac{p_1^o}{p_1^o - p_1}\right)\left(\frac{M_1 w_2}{w_1}\right)}$

13. **Elevation of boiling point** $\Delta T_b = K_b m$

14. **Depression of freezing point,** $\Delta T_f = K_f m$

15. **Osmotic pressure** $\Pi = iCRT$, i= 1 for non-electrolytes.

16. $\Pi = CRT = \left(\frac{n}{V}\right)RT \Rightarrow \Pi V = nRT = \left(\frac{w}{M}\right)RT \Rightarrow M = \frac{wRT}{\Pi V}$ = Molar mass of the solute.

17. **Van't Hoff factor,** $i = \frac{\text{No.of particles after dissociation or association}}{\text{No.of partciles before dissociation or association}}$

18. **Degree of dissociation,** $\alpha = \frac{i-1}{n-1}$; Degree of association, $\alpha = \frac{1-i}{1-(1/n)}$