SOLUTIONS

Section A : LEARNING POINTS:

Units of concentration of Solution :

(i) Mass Percentage(w/w): Amount of solute present in grams dissolved per 100g of solution.

Ex : 10% (w/w) glucose in water by mass, it means that 10 g of glucose is dissolved in 90 g of water resulting in a 100 g solution.

(ii) Volume percentage (v/v): Volume of solute present in 100ml of solution.

Ex : 10%(v/v) Alcohol in water by volume, it means that 10 ml of alcohol present in a 100 ml of solution.

(iii) Parts per million (ppm)- Amount of substance present in grams in 10^6 gm of solution.

(iv) Mole fraction (X) – It is the ratio of number of mole of a particular component to the total number of moles of all the components present in the solution.

$$X_A = n_A \ / \ (n_A + n_B), \qquad X_B = n_B \ / \ (n_A + n_B),$$

Where X_A & X_B are the mole fractions of Solvent and Solute respectively.
n_A & n_B are the number of moles of Solvent and Solute respectively.

(v) Molarity (M) – No. of mole of solute present per litre of solution.

$$M = (n_B / V) \times 1000$$

$$\boxed{m = \frac{w_B \times 1000}{M_B \times V}}$$

Where , n_B – No. of moles of solute
V – Volume of solution is ml.
w_B – amount of solute
M_B- Molar Mass of solute

(vi) Molality :-(m) No. of moles of solute present per 1000 g or 1. kg solvent.

$$\boxed{m = \frac{w_B \times 1000}{M_B \times w_A}}$$

Where , n_B – No. of moles of solute
w_A – amount of solvent in gm.
w_B – amount of solute
M_B- Molar Mass of solute

Solubility:

1) Solubility of a Solid in Liquids: It is observed that polar solutes dissolve in polar solvents and non polar solutes in nonpolar solvents.

Ex: Sodium chloride and sugar dissolve readily in water.Naphthalene and anthracene dissolve readily in benzene.

In general, a solute dissolves in a solvent if the intermolecular interactions are similar in the two.

Dissolution: When a solid solute is added to the solvent, some solute dissolves and its concentration increases in solution. This process is known as dissolution.

Crystallisation: Some solute particles in solution collide with the solid solute particles and get separated out of solution. This process is known as crystallisation.

A stage is reached when the two processes occur at the same rate. Under such conditions, number of solute particles going into solution will be equal to the solute particles separating out and a state of dynamic equilibrium is reached.

Solute + Solvent ↔Solution

Effect of temperature:

The solubility of a solid in a liquid is significantly affected by temperature changes. According to **Le Chateliers Principle**, the dissolution process is endothermic ($\Delta_{sol} H > 0$), the solubility should increase with rise in temperature and if it is exothermic ($\Delta_{sol} H < 0$) 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.

2. Solubility of gases in Liquids:

1) Oxygen dissolves only to a small extent in water. It is this dissolved oxygen which sustains all aquatic life.
2) Hydrogen chloride gas (HCl) is highly soluble in water.

Solubility of gases in liquids is greatly affected by pressure and temperature. The solubility of gases increase with increase of pressure.

Henry was the first to give a quantitative relation between pressure and solubility of a gas in a solvent which is known as

Henry's law:. The law states that at a constant temperature, **the solubility of a gas in a liquid is directly proportional to the pressure of the gas.**

The solubility of a gas in a liquid solution is a function of partial pressure of the gas. If we use the mole fraction of a gas in the solution as a measure of its solubility, then it can be said that
"The **mole fraction of gas in the solution is proportional to the partial pressure of the gas over the solution.**"

$$p = K_H x$$

Here K_H →is the Henry's law constant.
X→ is molefraction of gas in the solution.
P→ partial pressure of the gas.

Henry's law & its applications:

1) Solubility of a gas increases with decrease of temperature. It is due to this reason that aquatic species are more comfortable in cold waters rather than in warm waters.

2) To increase the solubility of CO_2 in soft drinks and soda water, the bottle is sealed under high pressure.

3) Scuba divers must cope with high concentrations of dissolved gases while breathing air at high pressure underwater. Increased pressure increases the solubility of atmospheric gases in blood. When the divers come towards surface, the pressure gradually decreases. This releases the dissolved gases and leads to the formation of bubbles of nitrogen in the blood. This blocks capillaries and creates a medical condition known as *bends*, which are painful and dangerous to life.
To avoid bends, as well as, the toxic effects of high concentrations of nitrogen in the blood, the tanks used by scuba divers are filled with air diluted with helium (11.7% helium, 56.2% nitrogen and 32.1% oxygen).

4) At high altitudes the partial pressure of oxygen is less than that at the ground level. This leads to low concentrations of oxygen in the blood and tissues of people living at high altitudes or climbers. Low blood oxygen causes climbers to become weak and unable to think clearly, symptoms of a condition known as *anoxia*.

Effect of Temperature:

Solubility of gases in liquids decreases with rise in temperature. The dissolution of a gas in liquid is an exothermic process involves dynamic equilibrium and thus must follow Le Chatelier's Principle. As the temperature increases the solubility of gas decreases.

From Henry's law:

As the temperature increases the value of K_H increases, we know that K_H is inversely proportional to Mole fraction of the gas i.e solubility of the gas (From Henry's law). So, as the temperature increases the solubility of a gas decreases.

Vapour Pressure and Raoult's law :-

a) Raoult's law for binary solutions of volatile liquids: At a given temperature, for a solution of volatile liquids, the partial vapour pressure of each component is equal to the product of the vapour pressure of the pure component and its mole fraction .

If the solution contains A & B are two volatile liquids, then

i.e $P_A = p^0_A X_A$ & $P_B = p^0_B X_B$

Where → p_A and p_B are the vapour pressures of A and B in solution respectively.
p^0_A α p^0_B are the vapour pressures of A and B in their pure state respectively.
x_A and x_B are the molefractions of A and B in solution respectively.

b) Raoult's law for solution containing Non – volatile solute: At a given temperature, the relative lowering vapour pressure of a solution is equal to the mole fraction of the solute.

Derivation :

We know that, from Raoults law,

$P_{\text{solution}} = p_A + p_B$
$P_{\text{solution}} = p^0_A X_A + p^0_B X_B$
$P_{\text{solution}} = p^0_A X_A + 0$ (Since $p^0_B = 0$, because B is anon volatile solute)
$P_{\text{solution}} = X_A P^0_A$

⇒

$P_{\text{solution}} = (1-X_B) P^0_A$ (Since $X_A + X_B = 1$)

⇒

$(P_{\text{solution}} - P^0_A) = - X_B P^0_A$

⇒

$$(P^0_A - P_{solution}) / X_B = P^0_A$$

Where , $(P^0_A - P_{solution})$ →Lowering of vapour pressure

$(P^0_A - P_{solution}) / P^0_A$ → Relative lowering of vapour pressure.

X_B → Mole fraction of Solute.

Ideal and Non-Ideal Solutions :

a) Ideal Solutions : The solution which obeys Raoult's law exactly at all concentration and all temperatures.

The ideal solution also have following characteristics:

i) It should obey the Raoult's law (Total pressure, $P = p^0_A x_A + p^0_B x_B$)

ii) Heat exchange on mixing is zero ($\Delta H_{mix} = 0$)

iii) Volume exchange on mixing is zero ($\Delta V_{mix} = 0$)

b) Non ideal solutions:

The solution which do not obey Raoult's law are called non ideal solutions. For these solutions

i) $p_A \# p^0_A x_A$ and $p_B \# p^0_B x_B$, ii) ΔH mixing # 0, iii) ΔV mixing # 0

Types of non ideal Solutions :

A) Non ideal solution showing positive deviation from Raoult's law :

→ The Mixture contains two components namely A and B, If the interaction between A-B molecules is weaker than the interactions between A-A or B-B or both, then the solution deviate from the ideal behavior and Each component of solution has a partial vapour pressure greater than expected on the basis of Raoult's law .

→The total vap. Pressure will be greater than corresponding Vap. Pressure expecterd in case of ideal solution of the same composition

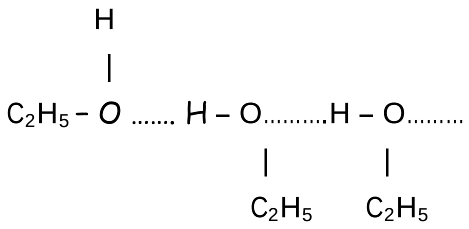
$$p_A > p^0_A x_A \quad \text{and} \quad p_B > p^0_B x_B$$

Total Vapour pressure P , $P = (p_A + p_B) > (p^0_A x_A + p^0_B x_B)$

→ Example: Mixture of Ethyal alcohol and Cyclohexane.

→Explanation for positive deviation:

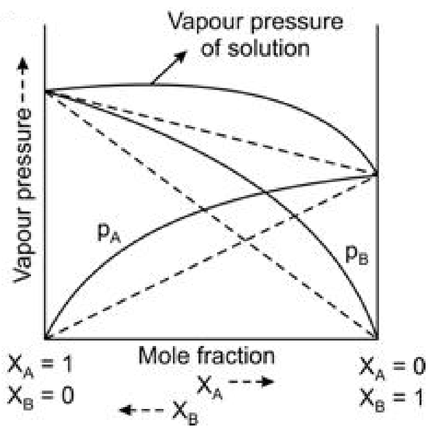
→In ethyl alcohol the molecules are held together by hydrogen bonding.



$$C_2H_5-\overset{\overset{H}{|}}{O}\cdots H-\underset{\underset{C_2H_5}{|}}{O}\cdots H-\underset{\underset{C_2H_5}{|}}{O}\cdots$$

→When Cyclohexane is added to ethyl alcohol, the Cyclohexane molecules occupy spaces between ethyl alcohol molecules. As a result , some hydrogen bonds in alcohol molecules break. The escaping tendency of alcohol and Cyclohexane molecules from solution increases. So there is increase in vapour pressure. In such case i) $\Delta H_{mixing} > 0$, ii) $\Delta V_{mixing} > 0$

Graph of a Positive deviation:



B) Non ideal solution showing negative deviations from Raoult's law:

→ The Mixture contains two components namely A and B, If the interaction between A-B molecules is stronger than the interactions between A-A or B-B or both, then the solution deviate from the ideal behavior and Each component of the solution has a partial vapour pressure less than from pure liquids. As a result, each component of solution has a partial vapour pressure less than expected on the basis of Raoult's law. Hence total vap.pressure becomes less than the corresponding vap. pressure expected in case of ideal solution

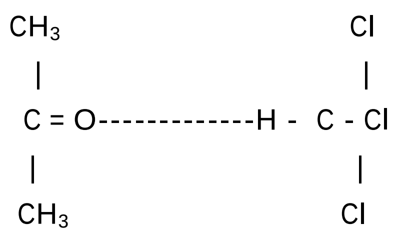
$p_A < p^0_A \; x_A$ and $p_B < p^0_B \; x_B$

Total Vapour pressure P , $P = (p_A + p_B) < (p^0_A \; x_A + p^0_B \; x_B)$

→ Example: Mixture of Acetone and Chloroform

→Explanation for negative deviation:

→By Mixing Acetone and chloroform, a new attractive forces are formed due to in intermolecular hydrogen bonding. Hence the attractive forces become stronger and the escaping tendency of each liquid from the solution decreases.

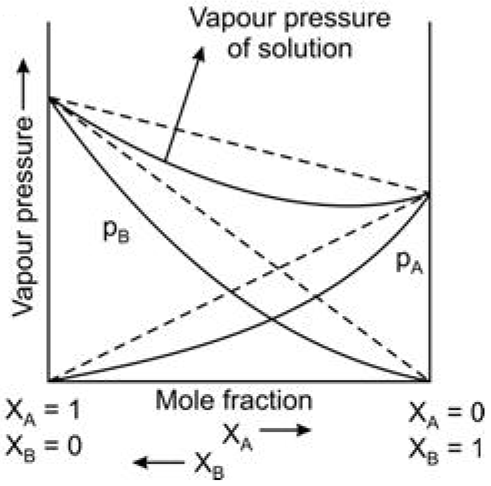


$$\begin{array}{ccc} CH_3 & & Cl \\ | & & | \\ C=O & \text{-------------} & H-C-Cl \\ | & & | \\ CH_3 & & Cl \end{array}$$

→So, the Vapour pressure of the solution is less than expected for an ideal solution . In such case

(i) ΔH mixing < 0
(ii) ΔV mixing < 0

Graph of a Negative deviation:



Azeotropes:

The binary mixtures having the same composition in liquid and vapour phase and boil at a constant temperature.

Types of Azeotropes:

i) **Minimum boiling azeotropes:**

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The solutions which show a large positive deviation from Raoult's law form minimum boiling azeotrope at a specific composition.

Example: Ethanol-water mixture (obtained by fermentation of sugars) on fractional distillation gives a solution containing approximately 95% by volume of ethanol. Once this composition, known as azeotrope composition, has been achieved, the liquid and vapour have the same composition, and no further separation occurs.

ii) **Maximum boiling azeotropes:**

The solutions that show large negative deviation from Raoult's law form maximum boiling azeotrope at a specific composition.

Example: Nitric acid and water is an example of this class of azeotrope. This azeotrope has the approximate composition, 68% nitric acid and 32% water by mass, with a boiling point of 393.5 K.

Colligative Properties:

The properties of the solutions which depends only on the number of the solute particles but not on the nature of the solute are called Colligative properties.

Four important Colligative properties are:

a) Relative lowering of vapour pressure.
b) Elevation of boiling point.
c) Depression of freezing point
d) Osmotic pressure.

a) Relative lowering of vapour pressure:

According to Raoult's law , At a given temperature, the relative lowering vapour pressure of a solution is equal to the mole fraction of the solute.

$$\boxed{X_b = \frac{(P^0_A - P_{solution})}{P^0_A}}$$

$$\boxed{\frac{n_b}{n_A + n_B} = \frac{(P^0_A - P_{solution})}{P^0_A}}$$

$$\boxed{\frac{n_B}{n_A} = \frac{(P^0_A - P_{solution})}{P^0_A}}$$

(since for dilute solutions $n_A >>> n_B$)

Where, $n_A \rightarrow$ No. of moles of solvent in solution
$n_B \rightarrow$ No. of moles of solute in solution

$$\boxed{\frac{w_B/M_B}{w_A/M_A} = \frac{(P^0_A - P_{solution})}{P^0_A}}$$

$$\boxed{M_B = \frac{w_B \times P^0_A \times M_A}{(P^0_A - P_{solution}) \times w_A}}$$

Where, w_A and w_B are the Mass of solvent and solute respectively.
M_A and M_B are the Molar mass of Solvent and Solute respectively.
P^0_A and $P_{solution}$ are the vapour pressure of pure solvent and solution respectively.

Thus, molar mass of the solute can be determined if the other quantities w_A , w_B, M_A and relative lowering of vapour pressure ($P^0_A - P_{solution}$) are known .

 # 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:

$$p_1 = y_1 \, p_{total} \quad \& \quad p_2 = y_2 \, p_{total}$$

In general, $p_i = y_i \, p_{total}$

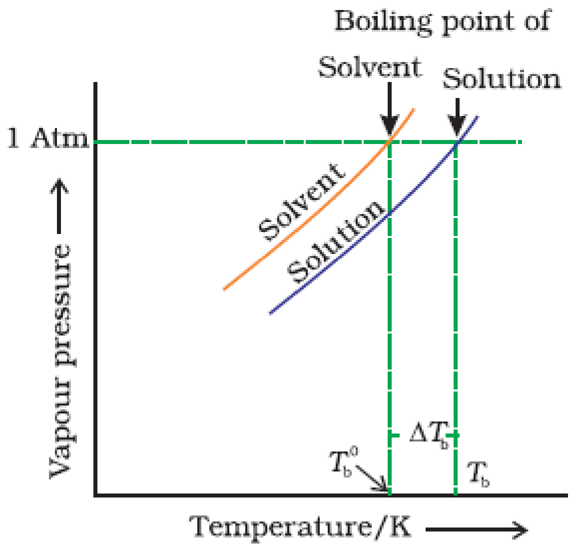
b) Elevation of boiling point:.

This vapour pressure of the solution at a given temperature is found to be lower than the vapour pressure of the pure solvent at the same temperature. In the solution, the non volatile solute particles obstruct the escaping of solvent molecules from the surface,thus, the vapour pressure is also reduced.

As you know that, as the temperature increases, the vapour pressure also increases.

The temperature at which the vapour pressure of the solvent/solution is equal to the atmospheric pressure, the solvent/solution boils.

Let T_b^0 be the boiling point of pure solvent and T_b be the boiling point of solution. The increase in the boiling point $\Delta T_b = T^0_b - T_b$ is known as elevation of boiling point.



$$\Delta T_b = K_b .m$$

$$(T_2 - T_1) = K_b .m \text{ ------------}\rightarrow \text{Eq (1)}$$

Where T_1 = Boiling point of Solvent in absolute temperature

T_2 = Boiling point of Solution in absolute temperature

m = Molality of the solution.

$$\Delta T_b = \frac{1000 \times K_b \times w_B}{M_b \times w_A}$$

Where W_B gram of a non volatile solute dissolved in W_A grams of the solvent and M_b is the Molar mass of the solute .

$$M_b = \frac{1000 \times K_b \times w_B}{\Delta T_b \times w_A}$$

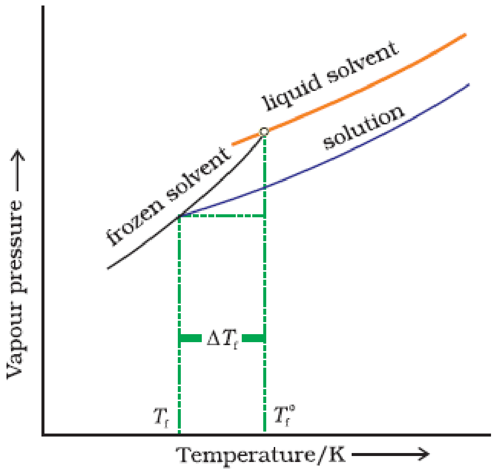
Thus using the above equation, molar mass of the solute is calculated.

c) Depression of freezing point:

The lowering of vapour pressure of a solution causes a lowering of the freezing point compared to that of the pure solvent.

A solution will freeze when its vapour pressure equals the vapour pressure of the pure solid solvent.

According to Raoult's law, when a non-volatile solid is added to the solvent it's vapour pressure decreases and now it would become equal to that of solid solvent at lower temperature. Thus, the freezing point of the solvent decreases.



$$\Delta T_f = K_f .m$$

$$(T_1 - T_2) = K_f .m \text{ ------------} \rightarrow \text{ Eq (1)}$$

Where T_1 = Freezing point of Solvent in absolute temperature
T_2 = Freezing point of Solution in absolute temperature
m = Molality of the solution.

$$\Delta T_f = \frac{1000 \times K_f \times w_B}{M_b \times w_A}$$

Where W_B gram of a non volatile solute dissolved in W_A grams of the solvent and M_b is the Molar mass of the solute .

$$M_b = \frac{1000 \times K_f \times w_B}{\Delta T_f \times w_A}$$

Thus using the above equation, molar mass of the solute is calculated.

d) Osmotic pressure:

Osmosis:

The phenomenon of the flow of solvent through a semi permeable membrane from pure solvent(low concentration) to the solution (high concentration) is called osmosis.

Osmotic Pressure:

The pressure applied on a solution to prevent the passage of solvent into the pure solvent through a semi permeable membrane is called osmotic pressure.

Reverse Osmosis:

The phenomenon of the flow of solvent through a semi permeable membrane from solution to pure solvent is called Reverse osmosis.

This process occurs, when the pressure applied on the solution is more than the Osmotic pressure.

Reverse Osmosis is used for the purification of Water.

Semi Permeable Membrane:

The membrane which allows only the solvent and not the solute to pass through it. E.g. Parch meat, cellophane membrane etc.

Van't Hoff observed that for dilute solutions, the osmotic pressure (π) is given as :-

$$\pi = C R T$$

Where π = Osmotic Pressure of the solution in atm

C = Molar concentration of the solution in mol/lit

R = Gas constant =0.0821 lit-atm-mol^{-1}-K^{-1}

T = Absolute Temperature in K

According to Van't Hoff equation

$$\Pi = CRT$$, where Π = Osmotic pressure of the solution

But Concentration, C = n / V

'n' is the number of moles of solute dissolved in 'V' litre of the solution,

$$\Pi = (n / V) \times (RT) \dashrightarrow \text{Eq (1)}$$

Again, $n = W_B / M_B$ Where W_B is the mass of solute and M_B is the molar mass of the solute.

Substituting 'n' in the Eq(1)

$$\boxed{\Pi = \frac{w_B \times R \times T}{w_B \times V}}$$

Molecular mass of Solute = $$\boxed{M_B = \frac{w_B \times R \times T}{\Pi \times V}}$$

Thus using the above equation, molar mass of the solute is calculated.

Advantage :

a) The osmotic pressure method has the advantage over other methods as pressure measurement is around the room temperature and the molarity of the solution is used instead of molality.

b) For determination of molar mass of solutes is particularly useful for biomolecules as they are generally not stable at higher temperatures.and polymers have poor solubility.

Isotonic Solutions: Two solutions having same osmotic pressure at a given temperature are called isotonic solutions. When such solutions are separated by semipermeable membrane no osmosis occurs between them.

Example: The osmotic pressure associated with the fluid inside the blood cell is equivalent to that of 0.9% (mass/volume) sodium chloride solution, called normal saline solution and it is safe to inject intravenously.

Hypertonic : If the salt concentration is more than 0.9% (mass/volume), the solution is said to be hypertonic. In this case, water will flow out of the cells if placed in this solution and they would shrink.

Hypotonic: If the salt concentration is less than 0.9% (mass/volume), the solution is said to be hypotonic. In this case, water will flow into the cells if placed in this solution and they would swell.

Application of Osmosis:

1) A raw mango placed in concentrated salt solution loses water via osmosis and shrivel into pickle. 2) Wilted flowers revive when placed in fresh water.

3) Water will move out of the blood cells through osmosis, when placed in water containing more than 0.9% (mass/volume) salt, blood cells collapse due to loss of water by osmosis.

4) People taking a lot of salt or salty food experience water retention in tissue cells and intercellular spaces because of osmosis. The resulting puffiness or swelling is called **edema**.

5) Water movement from soil into plant roots and subsequently into upper portion of the plant is partly due to osmosis.

6) The preservation of meat by salting and of fruits by adding sugar protects against bacterial action. Through the process of osmosis, a bacterium on salted meat or candid fruit loses water,
shrivels and dies.

Abnormal Molar Masses:

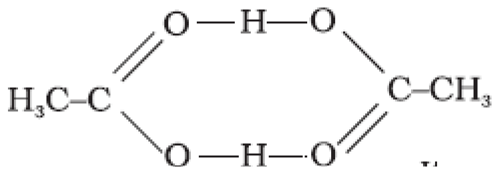
In some cases, the molar masses determined by colligative properties do not agree with the expected or theoretical values . This may be due to the association **or** the dissociation of the solute molecule in the solution. The two changes are being discussed below:

1) Association of solute particles: – In certain polar solvents, the solute molecules undergo association i.e. two three or even more molecules exists in combination with each other to form bigger molecules .

Total no. of molecules in solution become less than the no. of molecules of the substance added and, therefore colligative properties will be lower, since colligative properties inversely proportional to the molar mass will be greater than theoretical values in such cases.

Example: Acetic acid CH_3 COOH exist in dimmer by hydrogen bond.

$$2\ CH_3COOH \rightleftharpoons (CH_3COOH)_2$$



2) Dissociation of solute particles: Molecules of some substance dissociate or ionize in a solvent to give two or more particles.

The total no.of particles increases in solution and consequently colligative properties of such solution will be larger so the molar mass will be less than theoretical values in such cases.

Example: KCl , KCl = K^+ + Cl^-

Van’t hoff factor (i):

It is defined as

$$i = \frac{\text{Normal molar mass}}{\text{Observed molar mass}}$$

In case of association, the observed molar mass more than the normal, so the factor “ $i < 1$”

In case of dissociation, the observed molar mass has a lesser value, so “ $i > 1$”.

In case of solute where neither association nor dissociation in solvent, the Van”t factor “ $i = 1$”

In terms of colligative property ‘i’ may be defined as :

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

Inclusion of Van’t Hoff Factor ‘i’ modifies the equations for colligative properties as follows:-

Elevation of boiling point,	$\Delta T_b = i K_b m$
Depression of Freezing point,	$\Delta T_f = i K_f m$
Osmotic Pressure	$\pi = i C R T$

Section – B : Important Conceptual Questions and Answers:

1Q) A solution of glycol containing 1.82 gm/litre has an osmotic pressure of 51.8 cm of mercury at 10^0. What is the molecular weight of glycol? (1 Mark)

(**Ans**) M_2= w R t / P V =1.82 x 0.0821 x 2.83/ (51.8/76) x 1 = 62.04

2Q) Which one of the following salts will have the same value of van't Hoff factor (i) As that of $K_4[Fe(CN)_6]$. (1 Mark)

(**Ans**) $Al_2(SO_4)_3$

Explanation :

$K_4[Fe(CN)_6]$ and $Al_2(SO_4)_3$ each dissociates to give 5 ions or i = 5

$$K_4[Fe(CN)_6] \rightleftharpoons 4K^+ + [Fe(CN)_6]^-$$

3Q) A 5% solution of cane sugar (mol. wt. =342) is isotonic with 1% solution of a substance X. Calculate the molecular weight of X. (2 Mark)

(**Ans**) Osmotic pressure of 5% cane sugar solution (p_1) = C × R× T = $\frac{50\ g/L}{342} \times 0.0821 \times T$

Osmotic pressure of 1% solution of subtance (p_2) = $\frac{10\ g/L}{M} \times 0.0821 \times T$

Both are isotonic, So $p_1 = p_2$

$$\frac{50}{342} \times 0.0821 \times T = \frac{10}{M} \times 0.0821 \times T$$

Hence, M (mol. wt. of X) = 342/5 =68.4

4Q) During osmosis, Mention the flow of solvent through a semipermeable membrane. (1 Mark)

(**Ans**) During osmosis solvent flows through semipermeable membrane from lower concentration to higher concentration.

5Q) A solution containing 10 g per dm^3 of urea (molecular mass = 60 g mol^{-1}) is isotonic with a 5% solution of a non-volatile solute. The molecular mass of this non-volatile solute is (2 Mark)

(**Ans**) Osmotic pressure of urea from the formula ,

$$\pi V = nRT \Rightarrow \pi = \frac{nRT}{V}$$

$$\Rightarrow \pi = \frac{10}{60} \times .082 \times T$$ [1 dm = 1 litre]

Osmotic pressure of solution having non-volatile solute $= \frac{5}{M \times .1} \times .082 \times T$

Equating these pressures,

$$\frac{10}{60} \times .082 \times T = \frac{5}{M \times .1} \times .082 \times T$$

M = 300 gm mol^{-1}

6Q) Vapour pressure of CCl_4at 25?C is 143 mm Hg. 0.5 g of a non-volatile solute (mol. wt. 65) is dissolved in 100 ml of CCl_4. Find the vapour pressure of the solution. (Density of CCl_4= 1.58 g/cm3) (2 Mark)

(**Ans**) Relative lowering of V.P. = mole fraction of Solute

$$\frac{p^\circ - p_s}{p^\circ} = x_2$$

Density of CCl_4 = Mass of CCl_4 Volume

Mass of CCl_4 = Density x Volume

= 1.58 x 100

= 158 g

$$\frac{143 - p_s}{143} = \frac{0.5/65}{158/154};\ p_s = 141.93\,\text{mm}$$

7Q) The relationship between osmotic pressure at 273 K when 10g glucose (P_1), 10 g urea (P_2), and 10g sucrose (P_3) are dissolved in 250 ml of water is (2 Mark)

(**Ans**) Moles of glucose =

$$\frac{10}{180} = 0.05, \quad \text{Urea} = \frac{10}{60} = 0.16, \quad \text{Sucrose} = \frac{10}{342} = 0.029$$

Hence osmotic pressure $p_2 > p_1 > p_3$

8Q) Two elements A and B form compounds of formula AB_2 and AB_4. When dissolved in 20.0 g of benzene 1.0 g of AB_2 lowers freezing point by 2.3?C whereas 1.0 g of AB_4 lowers freezing point by 1.3?C. The K_f for benzene is 5.4. The atomic masses of A and B will be (3 Mark)

(**Ans**) Let the masses of A and B be a and b. The molecular mass of AB_2 will be (a + 2b) mol^{-1}g and AB_4will be (a + 4b) g mol^{-1}

For AB_2, 2.3 = (5.1 x 1 x 1000)/[(a+2) x 20]

$$\text{or}\quad a + 2b = \frac{5.1 \times 1 \times 1000}{20 \times 2.3} \quad ...(I)$$

For AB_4, 1.3 = (5.1 x 1 x 1000)/[(a+4b) x 20]

$$\text{or}\quad a + 4b = \frac{5.1 \times 1000}{20 \times 1.3} \quad ....(II)$$

On solving (I) and (II),

a = 25.49 and b = 42.64

9Q) 250 ml of sodium carbonate solution contains 2.65gm of sodium carbonate. If 10 ml of this solution is diluted to one litre, What is the concentration of resultant solution? (mol. Wt. of sodium carbonate=106) (2 Mark)

(**Ans**) Initial concentration =2.65 gm / 250 cc = 10.60 gm/L= (10.6/106) = 0.1M.

When 10 ml is diluted to one litre solution =100 times.

Hence final conc.= 0.1/100 = 0.001M .

10Q) Write Henry's law. **(1 Mark)**

(**Ans**) The solubility of a gas in a liquid is directly proportional to the partial pressure of the gas at a given temperature.

11Q) What happens when blood cells are placed in pure water? **(1 Mark)**

(**Ans**) Water molecules move into blood cells through the cell walls. So, blood cells swell and may even burst.

12Q) Two liquids A and B boil at 120^0c and 160^0c respectively. Which of them has higher vapour pressure at 70^0 c? **(1 Mark)**

(**Ans**) Lower the boiling point, more volatile it is .So liquid A will have higher vapour pressure at 70^0c

13Q) A solution of 1.25gm of a non-electrolyte in 20gm of water freezes at 271.94 K. If K_f= 1.86K/m, then calculate the molar mass of the solute. **(2 Mark)**

(**Ans**) According to depression of freezing point.

$$M_b = \frac{1000 \times K_f \times w_B}{\Delta T_f \times w_A}$$

= (1000 X 1.86 X 1.25) / ((273-271.94) x 20)

= 109.67gm/mol.

14Q) Osmotic pressure of a solution is 0.0821 atm at a temperature of 300 K. The concentration in mole/litre is: **(1 Mark)**

(**Ans**) C= π/RT = 0.0821/(0.0821 x 300) = 0.003

15Q) A solution of ethanol in water is 1.6 molal. How many grams of ethanol are present in 500g of the solution. **(2 Marks)**

(**Ans**) Mass of ethanol = Molality × M.wt.

= 1 6 × 46 = 73.6g

Total of mass of solution = 73.6 + 1000 = 1073.6g

∴ 1073.6g of solution contain 7.6g of ethanol.

∴ Mass of ethanol in 500g of solution $= \frac{73.6}{1073.6} \times 500$

$= 34.28g$

16Q) On a hill station pure water boils at 99.82^{0}C. The K_b of water is 0.513?C Kg mol^{-1}. Calculate the boiling point of 0.69m solution of urea. (1 Marks)

(**Ans**)

$$\Delta T_b = K_b m$$
$$= 0.513 \times 0.69$$
$$= 0.3539\,°C$$

Boiling point of solution = Boiling point of water + ΔT_b

= 99.82 + 0.3539

= 100.17^0 C

17Q) Find the vant Hoff factor for $Al_2(SO_4)_3$ (1 Marks)

(**Ans**) $Al_2(SO_4)_3 \longrightarrow 2Al^{3+} + 3SO_4^{2-}$

Total ions produced = 2 + 3 = 5

$$i = \frac{\text{observed no. of ions}}{\text{Normal no. of ions}} = \frac{5}{1} = 5$$

18Q) Ethylene glycol solution having molality 0.5 is used as coolant in a car. Calculate the freezing point of solution (given K_f=1.86 K Kg/mole) (2 Marks)

(**Ans**)

$$\Delta T_f = k_f \times m$$
$$= 1.86 \times 0.5$$
$$= 0.93°C.$$
$$\Delta T_f = T_f^\circ - T_f$$
$$\text{or} \quad T_f = T_f^\circ - \Delta T_f$$
$$= 0 - 0.93$$
$$= -0.93°C$$

19Q) An aqueous solution freezes at –0.186^{0}C. K_f = 1.86?, K_b = 0.512. Find elevation in boiling point. (2 Marks)

(**Ans**) ΔT_f= 0 – (0.186) = 0.186^{0}C

ΔT_f = K_f x m

Or $$m = \frac{\Delta T_f}{K_f} = \frac{0.186}{1.86} = 0.$$

ΔT_b = K_b x m

$$= 0.512 \times 0.1 = 0.0512\,°C$$

20Q) Density of 1 M solution of glucose 1.18g/cm^3. K_f for H_2O is 1.86 Km^{-1}. Find freezing point of solution. (2 Marks)

(**Ans**) Mass of solution = Volume x Density

= 1000 x 1.18 =1180g

Mass of water = 1180 -180 =1000g

∴ 1M C_2H_5OH = 1 m

ΔT_f = K_f x m

= 1.86 x 1 = 1.86

$T_f = T^0{}_f - \Delta T_f$

= 0 – 1.86

= –1.86^{0}C

21Q) Vapour pressure of two liquid A and B are 120 and 180mm Hg at a given temperature. If 2 mole of A and 3 mole of B are mixed to form an ideal solution, calculate the vapour pressure of solution at the same temperature. (2Marks)

(**Ans**) Total moles = 2 + 3 = 5

$P_{\text{solution}} = P^0{}_A X_A + P^0{}_B X_B$

= (120 x 2/5) + (180 x 3/5)

= 48 + 108

= 156mm.

22Q) The Osmotic pressure of human blood is 7.65 atm at 37^{0}C. For injecting glucose solution it is necessary the glucose solution has same osmotic pressure as of human blood. Find the molarity of glucose solution having same osmotic pressure as of human blood. (2 Marks)

(**Ans**) π = CRT =(n/v) RT

7.65 = (n/v) x 0.0821 x 310

n/v = 7.65 / (0.0821 x 310)

= 0.30mol/lit

Molarity = 0.30M

23Q) A solution contains 25% water, 25% ethanol and 50% acetic acid by mass. Find mole fraction of each of the component. (3 Marks)

(**Ans**) Moles of water = $n_1 = {}^{25}/_{18} = 1.39$

Moles of C_2H_5OH = $n_2 = {}^{25}/_{46} = 0.54$

Moles of CH_3COOH = $n_3 = {}^{50}/_{60} = 0.83$

Total moles in solution = $n_1 + n_2 + n_3 = 1.39 + 0.54 + 0.83 = 2.76$

Mole fraction of water = $\dfrac{n_1}{n_1 + n_2 + n_3} = \dfrac{\mathbf{1}.39}{2.76} = 0.503$

Mole fraction of C_2H_5OH = $\dfrac{n_2}{n_1 + n_2 + n_3} = \dfrac{0.54}{2.76} = 0.196$

Mole fraction of CH_3COOH = $\dfrac{n_3}{n_1 + n_2 + n_3} = \dfrac{0.83}{2.76} = 0.301$

24Q) Conc. H_2SO_4 has a density 1.9g/ml and is 99% H_2SO_4 by weight. Find molarity of solution. (3 Marks)

(**Ans**) Mass of 1000 ml of H_2SO_4 = density × volume

=1.9 x 1000

= 1900 g

Mass of H_2SO_4 present in 1900 g (1L) of H_2SO_4

= (99/100) x 1900

= 1881 g

Mole of H_2SO_4 present in 1L = Molarity = mass / molar mass

=1881/ 98

= 19.197 M

25Q) Vapour pressure of pure water is 40mm. If a non-volatile solute is added to it, vapour pressure falls by 4 mm. Calculate molality of solution. (3 Marks)

(**Ans**) $\dfrac{\Delta p}{p^{\circ}} = x_B = {}^{4}/_{40} = 0.1$

∴ The solution has 0.1 moles of solute in 0.9 moles of water.

Mass of water = $Moles \times M.wt. = 0.9 \times 18 = 16.2g$

Molality of solution = $\dfrac{Moles\ of\ solute}{Mass\ of\ solvent\ in\ kg} = \dfrac{0.1}{16.2/1000}$

= 6.17m

26Q) The vapour pressure of benzene at certain temperature is 640mm Hg. To 39.08 of benzene, non-volatile and non-electrolyte solid-weighing 2.175g was added. The vapour pressure of solution was 600mm of Hg. Find the mass of the solute? (3 Marks)

(**Ans**)

$$\frac{p^\circ - p}{p^\circ} = x_B$$

$$\frac{p - p}{p^\circ} = \frac{w/m}{W/M} = \frac{w \times M}{m \times W}$$

$$\frac{640 \quad 600}{640} = \frac{2.175 \times 78}{m \times 39.08}$$

$$\frac{40}{640} = \frac{2.175 \times 78}{39.08m}$$

$$m = \frac{2.175 \times 78}{39.00} \times \frac{640}{40} = 59.63g$$

27Q) An aqueous solution of a non-volatile and non-electrolyte substance boils at 100.5^{0}C. Calculate osmotic pressure of this solution at 27^{0}C. K_b (for water) per 1000g = 0.50. (3 Marks)

(**Ans**) $\Delta T_b = T_b - T_b^\circ$

$100.5 - 100 = 0.5^0C$

$$\Delta T_b = K_b m \quad \text{or} \quad m = \frac{\Delta T_b}{K_b} = \frac{0.5}{0.50} = 1$$

∴ Molality of solution = 1, (∵ Solvent is water)

∴ density of solution = 1

Volume of solution = volume of solvent = 1000/1 = 1000ml = 1 L

$$T = 27°C = 27 + 273 = 300K, \ n = 1$$

$\pi = nRT$

$$\pi = \frac{nRT}{v} = \frac{1 \times 0.0821}{1} \times 300 = 24.63atm$$

28Q) A decimolar solution of NaCl exerts osmotic pressure of 4.6 atm at 300K. Find the degree of dissolution. (3 Marks)

(**Ans**) $NaCl \rightarrow Na^+ + Cl^-$

Initial moles:	1	0	0
Moles at Equilibrium:	1-α	α	α

Equilibrium

Total moles at equilibrium = $1 - \alpha + \alpha + \alpha = 1 + \alpha$