

Class: 12
Subject: Chemistry
Topic: Solutions
No. of Questions: 23

1. A solution is prepared by dissolving 43 gm of naphthalene in 117 gm of benzene. Calculate the mole fractions of the two components of the solution.

Sol. Number of moles, $n = \frac{W}{M}$

$$\text{Moles of benzene, } n_1 = \frac{117}{78} = 1.50 \text{ mol}$$

$$\text{Moles of naphthalene, } n_2 = \frac{43}{128} = 0.34 \text{ mol}$$

$$\text{Mole fraction of naphthalene, } X_2 = \frac{0.34}{1.50 + 0.34} = 0.185$$

$$\text{Mole fraction of benzene} = 1 - X_2 = 1 - 0.185 = \mathbf{0.815}$$

2. 1.2 gm of a non-volatile substance was dissolved in 100 gm of acetone at 20°C. The vapour pressure of the solution was found to be 182.5 torr. Calculate the molar mass of the substance (vapour pressure of acetone at 20°C is 185.0 torr)

Sol. From equation $\frac{P_1^0 - P_1}{P_1^0} = \frac{W_1 M_1}{W_2 M_2}$

$$\frac{185 - 182.5}{185} = \frac{1.2 \times 58}{100 \times M_1} \text{ or } M_1 = 51.51 \text{ gm/mol.}$$

3. A solution containing 8.6 gm per dm³ of urea (molar mass = 60 gm/mole) was found to be isotonic with a 5 percent solution of an organic non - volatile solute. Calculate the molar mass of the latter.

Sol. According to Van't Hoff theory, isotonic solutions have the same osmotic pressure at the same temperature and the same molar concentration.

$$\text{Molar concentration of urea solution} = \frac{8.6}{60}$$

Let M_2 be the molar mass of unknown solute.

$$\text{Molar concentration of unknown solution} = \frac{50}{M_2}$$

Since both are isotonic,

$$\frac{50}{M_2} = \frac{8.6}{60} \text{ or } M_2 = 348.8 \text{ gm/mol}$$

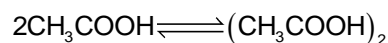
4. Acetic acid associate in benzene to form dimer. 1.65 gm of acetic acid when dissolved in 100 gm of benzene raised the boiling point by 0.36°C. Calculate the Van't Hoff factor and the degree of association of acetic acid in benzene. ($K_b = 2.57 \text{ K Kg mol}^{-1}$)

Sol. Normal molar mass of acetic acid = 60 gm/mol
 Observed molar mass of acetic acid, $M_2 = \frac{K_b W_2}{W_1 \Delta T_b}$

$$= \frac{2.57 \times 1.65 \times 10^{-3}}{100 \times 10^{-3} \times 0.36} = 0.1178 \text{ kg} = 117.8 \text{ gm/mol}^{-1}$$

$$i = \frac{\text{normal molar mass}}{\text{observed molar mass}} = \frac{60}{117.8} = 0.509$$

Since acetic acid associated to form double molecules. Hence



$$\begin{array}{ccc} 1 & & 0 \\ 1-\alpha & & \alpha/2 \end{array}$$

$$\begin{aligned} \text{The number of effective moles} &= 1 - \alpha + \alpha/2 \\ &= 1 - \alpha/2 \end{aligned}$$

$$i = \frac{1 - \alpha/2}{1}$$

$$0.508 = 1 - \alpha/2 \text{ or } \alpha = 0.984$$

5. At 25°C, the osmotic pressure of human blood due to the presence of various solutes in the blood is 7.65 atm. Assuming that molarity equals molality, calculate the freezing point of blood. $K_f = 1.86 \text{ K Kg mol}^{-1}$.

Sol. According to the Van't Hoff equation, $\pi = CRT$

$$C = \frac{\pi}{RT} = \frac{7.65}{0.0821 \times 298} = 0.313 \text{ mol dm}^{-3}$$

i.e. the concentration of various solutes present in blood = 0.313 mol dm⁻³.

In the present concentration, as given, molality of solution = molarity of solution

$$m = C = 0.313 \text{ mol kg}^{-1}$$

$$\Delta T_f = K_f m = 1.86 \times 0.313 = 0.582 \text{ K}$$

$$T_0 - T_f = \Delta T_f$$

$$T_f = T_0 - \Delta T_f = 273 - 0.582 = 272.418 \text{ K} \quad = -0.582 \text{ }^\circ\text{C}$$

IIT level Questions

6. The density of 2.0 M solution of acetic acid in water is 1.02 g ml^{-1} . Calculate the mole fraction of acetic acid.

Sol. Mass of acetic acid in 2.0 M acetic acid solution = $60 \times 2 = 120 \text{ gm per dm}^3$
 Mass of the solution = $V \times D = 1 \times 1.02 = 1.02 \text{ kg}$
 Mass of solvent = $1020 - 120 = 900 \text{ g}$
 Mole fraction of acetic acid = $\frac{2}{\frac{900}{18} + 2} = \frac{2}{50 + 2}$
 $= 0.038$

7. Liquid A and B form an ideal solution obeying Raoult's law. At 50°C , the total vapour pressure of a solution containing 1 mole of A and 2 mole of B is 300 torr. When 1 more mole of A is added to the solution, the vapour pressure increases to 400 torr. Calculate the vapour pressure of pure components.

Sol. $x_{A_1} = \frac{1}{1+2} = \frac{1}{3}$, $x_{B_1} = 1 - \frac{1}{3} = \frac{2}{3}$ (initially)
 $x_{A_2} = \frac{2}{2+2} = \frac{1}{2}$, $x_{B_2} = 1 - \frac{1}{2} = \frac{1}{2}$ (after addition of 1 mole of A)
 given, $P = x_A P_A^0 + x_B P_B^0$
 $300 = \frac{1}{3} P_A^0 + \frac{2}{3} P_B^0$ (initially)
 $400 = \frac{1}{2} P_A^0 + \frac{1}{2} P_B^0$ (after addition of 1 mole of A)
 Now from these equations
 $P_A^0 = 700 \text{ torr}$, $P_B^0 = 100 \text{ torr}$

8. A solution of A and B with 30 mole percent of A is in equilibrium with its vapour which contains 60 mole percent of A. Assuming ideality of the solution and the vapour. Calculate the ratio of the vapour pressure of pure A to that of pure B.

Sol. Given, $x_A = 0.3$ $x_B = 0.7$ in liquid or in solution.
 In the vapour phase, $X_A \text{ vapour} = 0.6$ and $x_B \text{ vapour} = 0.4$
 Using Dalton's law of partial pressures and Raoult's law,
 $x_{A, \text{vap}} = 0.6 = \frac{P_A}{P} = \frac{P_A}{P_A + P_B} = \frac{0.3P_A^0}{(0.3P_A^0 + 0.7P_B^0)}$
 $x_{B, \text{vap}} = 0.40 = \frac{P_B}{P} = \frac{P_B}{P_A + P_B} = \frac{0.7P_B^0}{(0.3P_A^0 + 0.7P_B^0)}$
 $\frac{x_{A, \text{vap}}}{x_{B, \text{vap}}} = \frac{0.60}{0.40} = \frac{0.30P_A^0}{0.70P_B^0}$

$$\frac{P_A^0}{P_B^0} = \frac{0.60 \times 0.70}{0.40 \times 0.30} = \frac{7}{2} = 3.5$$

9. What would be the vapour pressure of 0.5 molal solution of a non volatile solute in benzene at 30°C? The vapour pressure of pure benzene at 30°C is 119.6 torr.

Sol.
$$\frac{P_1^0 - P}{P_1^0} = \frac{mM}{1000} \text{ gm} \quad \dots(i)$$

Given that, $m = 0.5 \text{ mol}$, $M_1 = 78 \text{ gm/mol}$

$$P_1^0 = 119.6 \text{ torr}$$

Putting these values in above equation

$$\frac{119.6 - P_1}{119.6} = \frac{0.5 \times 78}{1000}$$

$$P_1 = 114.94 \text{ torr}$$

10. The molar heat of vaporization of water at 100°C is 40.585 KJ/mol. At what temperature will a solution containing 5.6 gm of glucose per 1000 gm of water boil?

Sol.
$$K_b = \frac{RT_b^2 M_1}{\Delta H_{\text{vap}}}$$

$$= \frac{8.314 \times (373)^2 \times 0.018}{40585} = 0.513 \text{ K.Kg mol}^{-1}$$

$$\Delta T_b = \frac{K_b W_2}{W_1 M_2} = \frac{0.513 \times 5.6 \times 10^{-3}}{1 \times 18 \times 10^{-3}} = 0.159 \text{ K}$$

$$\text{Boiling point of solution} = 373 \text{ K} + 0.16 \text{ K} = 373.16 \text{ K}$$

$$= 100.16 \text{ }^\circ\text{C}$$

11. An aqueous solution of non - volatile solute boils at 100.17°C. At what temperature would it freeze? For water, $K_b = 0.52 \text{ K Kg mol}^{-1}$ and $K_f = 1.86 \text{ K Kg mol}^{-1}$

Sol.
$$\Delta T_b = 373.17 - 373 = 0.17 \text{ K}$$

$$\Delta T_b = K_b m$$

$$\text{or } m = \frac{\Delta T_b}{K_b} = \frac{0.17}{0.52} = 0.327 \text{ mol/kg}$$

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

$$1.86 \times 0.327 = 0.608 \text{ K}$$

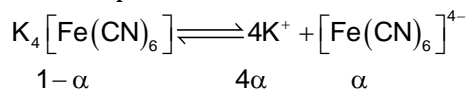
$$\text{Freezing point, } T_f = 273 - 0.608 \text{ K} = 272.39 \text{ K} = -0.608 \text{ }^\circ\text{C}$$

12. The complex compound $K_4[\text{Fe}(\text{CN})_6]$ is 45% dissociated in 0.1M aqueous solution of the complex at 27°C. What would be the osmotic pressure of the solution?

Sol. If the complex $K_4[Fe(CN)_6]$ had not been dissociated, its normal osmotic pressure would have been given by that Van't Hoff equation,

$$\pi = CRT = 0.1 \times 0.0821 \times 300 = 2.462 \text{ atm}$$

The complex, however is dissociated in solution



$$\text{No. of moles after dissociation} = 1 - \alpha + 4\alpha + \alpha = 1 + 4\alpha$$

$$\frac{\pi_{\text{observed}}}{\pi_{\text{normal}}} = \frac{1 + 4\alpha}{1}$$

$\alpha = 0.45$ hence

$$\pi_{\text{observed}} = \frac{2.462 \times (1 + 4 \times 0.45)}{1} = 6.894 \text{ atm}$$

13. The formula for low molecular weight starch is $(C_6H_{10}O_5)_n$ where n averages 2.00×10^2 . When 0.798 gm of starch is dissolved in 100 ml of water solution, what is the osmotic pressure at 25°C ?

Sol. No. of mole of starch = $\frac{0.789}{32,400} = 2.44 \times 10^{-5}$

[As M of starch = $162 \times 2 \times 10^2 = 32400$]

The molarity of the solution is = $\frac{2.44 \times 10^{-5}}{0.1} = 2.44 \times 10^{-4} \text{ M}$

Osmotic pressure $\pi = CRT$
 $= 2.44 \times 10^{-4} \times 0.0821 \times 298$
 $= 59.69 \times 10^{-4} \text{ atm}$

14. Phenol associates in benzene to a certain extent to form a dimer. A solution containing $20 \times 10^{-3} \text{ kg}$ of phenol in 1.0 kg of benzene has its f. pt depressed by 0.69 K. Calculate the fraction of phenol dimerised. K_f for $C_6H_6 = 5.120 \text{ mol}^{-1}\text{kg}$.

Sol. We know that, $\Delta T = \frac{1000 \times k_f' \times W_2}{M \times W_1}$

$$0.69 = \frac{1000 \times 5.12 \times 20 \times 10^{-3}}{m \times 1}$$

$$m_{\text{exp}} = 148.41$$

$$m_{\text{normal}} = 94$$

$$\frac{m_1}{m_{\text{normal}}} = \frac{94}{148.41} = 1 - \alpha + \frac{\alpha}{2}$$

$$\frac{1 - \alpha / 2}{1} = 0.635$$

or $\alpha = 0.734$ or 73.4 %

15. Calculate the amount of ice that will separate out on cooling a solution containing 50 gm of ethylene glycol in 200 gm water to -9.3°C .

$T_f = -9.3^{\circ}\text{C}$. (k_f for water = $1.86 \text{ K mol}^{-1} \text{ kg}$)

Sol. Given $\Delta T = 9.3$, $W_2 = 50 \text{ gm}$, $k_f = 1.86 \text{ K mol}^{-1} \text{ kg}$

$M_{\text{glycol}} = 62$

$$\Delta T = \frac{100k_f W_2}{MW_1}$$

$$9.3 = \frac{1000 \times 1.86 \times 50}{62 \times W}$$

$$\therefore W = 161.29 \text{ grams}$$

$$W_{\text{water}} = 161.29 \text{ gm}$$

$$\begin{aligned} \text{Thus weight of ice separated} &= 200 - 161.29 \\ &= 38.71 \text{ gm} \end{aligned}$$

16. What is vant hoff factor?

Sol. It is the ratio of normal molecular mass to observed molecular mass. It is denoted as "i"

$i = \text{normal m.m.} / \text{observed m.m}$

= no. of particles after association or dissociation / no. of particles before association or dissociation

17. Calculate the volume of 75% of H_2SO_4 by weight ($d=1.8\text{gm/ml}$) required to prepare 1 L of 0.2 M H_2SO_4

Sol: $M_1 = P \times D \times 10/98$

$$M_1 V_1 = M_2 V_2$$

$$14.5\text{ml}$$

18. An aqueous solution of 2% non-volatile exerts a pressure of 1.004 bar at the normal boiling point of the solvent. What is the molar mass of the solute.

Sol: $P_A^0 - P_A^0/P_A^0 = W_B X M_A / M_B X W_A$
 $1.013 - 1.004 / 1.013 = 2X18 / m_B X 98$
 $m_g = 41.35 \text{ gm/mol}$

19. Calculate the mass of ascorbic acid (vitamin C, $\text{C}_6\text{H}_8\text{O}_6$) to be dissolved in 75 g of acetic acid to lower its melting point by 1.5° $K_f = 3.9 \text{ K kg mol}^{-1}$

Sol: Mass of acetic acid, $w_1 = 75 \text{ g}$

$$\begin{aligned} \text{Molar mass of ascorbic acid } (C_6H_8O_6), M_2 &= 6 \times 12 + 8 \times 1 + 6 \times 16 \\ &= 176 \text{ g mol}^{-1} \end{aligned}$$

Lowering of melting point, $\Delta T_f = 1.5 \text{ K}$

We know that:

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

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

$$= \frac{1.5 \times 176 \times 75}{3.9 \times 1000}$$

$$= 5.08 \text{ g (approx.)}$$

Hence, 5.08 g of ascorbic acid is needed to be dissolved.

- 20** Calculate the mass of a non-volatile solute (molar mass 40 g mol^{-1}) which should be dissolved in 114 g octane to reduce its vapour pressure to 80 %

Sol: Let the vapour pressure of pure octane be P_1^0 .
 Then, the vapour pressure of the octane after dissolving the non-volatile solute is
 $\frac{80}{100} p_1^0 = 0.8 p_1^0$

Molar mass of solute, $M_2 = 40 \text{ g mol}^{-1}$

Mass of octane, $w_1 = 114 \text{ g}$

Molar mass of octane. (C_8H_{18}), $M_1 = 8 \times 12 + 18 \times 1$
 $= 114 \text{ g mol}^{-1}$

Applying the relation,

$$\frac{p_1^0 - p_1}{p_1^0} = \frac{w_2 \times M_1}{M_2 \times 114} = \frac{w_2 \times 114}{40 \times 114}$$

$$\Rightarrow \frac{p_1^0 - 0.8 p_1^0}{p_1^0} = \frac{w_2 \times 114}{40 \times 114}$$

$$\Rightarrow \frac{0.2 p_1^0}{p_1^0} = \frac{w_2}{40}$$

$$\Rightarrow 0.2 = \frac{w_2}{40}$$

$$\Rightarrow w_2 = 8 \text{ g}$$

Hence, the required mass of the solute is 8 g.

- 21.** 1.00g of a non-electrolyte solute dissolved in 50 g of benzene lowered the freezing point of benzene by 0.40 K. The freezing point depression constant of benzene is $5.12 \text{ k kg mol}^{-1}$. Find the molar mass of the solute.

Sol: Substituting the values of various terms involved in equation (2.36) we get,

$$M_2 = \frac{5.12 \text{ kg mol}^{-1} \times 1.00 \text{ g} \times 1000 \text{ g kg}^{-1}}{0.40 \times 50 \text{ g}} = 256 \text{ g mol}^{-1}$$

Thus, molar mass of the solute = 256 g mol^{-1}

22. 45g 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.

Sol: Depression in freezing point is related to the molality. Therefore, The molality of the solution with respect to ethylene glycol = $\frac{\text{moles of ethylene glycol}}{\text{mass of water in kilogram}}$

$$\text{Moles of ethylene glycol} = \frac{45 \text{ g}}{62 \text{ g mol}^{-1}} = 0.73 \text{ mol}$$

$$\text{Mass of water in kg} = \frac{600 \text{ g}}{1000 \text{ g kg}^{-1}} = 0.6 \text{ kg}$$

$$\text{Hence molality of ethylene glycol} = \frac{0.73 \text{ mol}}{0.60 \text{ kg}} = 1.2 \text{ mol kg}^{-1} \text{ Therefore freezing point}$$

$$\text{depression, } \Delta T = 1.86 \text{ K kg mol}^{-1} \times 1.2 \text{ mol kg}^{-1} = 2.2 \text{ K}$$

$$\text{Freezing point of the aqueous solution} = 273.15 \text{ K} - 2.2 \text{ K} = 270.95 \text{ K}$$

23. 200 cm^3 of an aqueous solution of a protein contains 1.26 g of the protein. The osmotic pressure of such a solution at 300 K is found to be 2.57×10^{-3} bar. Calculate the molar mass of the protein.

Sol: The various quantities known to us are as follows: $\pi = 2.57 \times 10^{-3}$ bar,

$$V = 200 \text{ cm}^3 = 0.200 \text{ litre } T = 300 \text{ K}$$

$$R = 0.083 \text{ L bar mol}^{-1} \text{ K}^{-1}$$

Substituting these values in equation $\pi = C \cdot R \cdot T / V$

we get

$$M_2 = \frac{1.26 \text{ g} \times 0.083 \text{ l bar K}^{-1} \text{ mol}^{-1} \times 300 \text{ K}}{2.57 \times 10^{-3} \text{ bar} \times 0.200 \text{ L}}$$

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