Unit 2 Solution : Numericals By- Shailendra Shukla Sir

Q.1. 4 g NaOH is dissolved in 500 ml of its aqueous solution. Find the normality of solution. (Gram equivalent Mass of NaOH = 40g).

Sol^{n}- Given that,

Mass of solute (NaOH) = 4 g

Gram Equivalent Mass of solute (NaOH) = 40 g

Volume of solution = 500 ml

Normality of solution, N = ?

$\because$ Normality, $N = \frac{Mass\ of\ solute\ (g) \times 1000}{\text{GEM of Solute} \times \text{Volume of Solution (ml)}}$

$\therefore$ Normality, $N = \frac{4\ g \times 1000}{40\ g \times 500\ ml}$

$= 0.2\ N$

Q.2. 1.8 g glucose ($C_6H_{12}O_6$) is dissolved in 500 ml of its aqueous solution. Find the molarity of solution. (Gram Molecular Mass of $C_6H_{12}O_6$= 180g).

Sol^{n}- Given that,

Mass of solute $C_6H_{12}O_6$ = 1.8 g

Gram Molecular Mass of solute ($C_6H_{12}O_6$) = 180g

Volume of solution = 500 ml

Molarity of solution, M = ?

$\because$ Molarity, $M = \frac{Mass\ of\ solute\ (g) \times 1000}{\text{GMM of Solute} \times \text{Volume of Solution (ml)}}$

$\therefore$ Molarity, $M = \frac{1.8\ g \times 1000}{180\ g \times 500\ ml}$

$= 0.02\ M$

Q.3. 3.42 g sucrose ($C_{12}H_{22}O_{11}$) is dissolved in 500 g of water. Find the molality of solution. (Gram Molecular Mass of $C_{12}H_{22}O_{11}$= 342 g).

Sol^{n}- Given that,

Mass of solute $C_{12}H_{22}O_{11}$ = 3.42 g

Gram Molecular Mass of solute ($C_{12}H_{22}O_{11}$) = 342 g

Mass of solvent = 500 g

Molality of solution, m = ?

$\because$ Molality, $m = \frac{Mass\ of\ solute\ (g) \times 1000}{\text{GMM of Solute} \times \text{Mass of solvent (g)}}$

$\therefore$ Molality, $m = \frac{3.42\ g \times 1000}{342\ g \times 500\ g}$

$= 0.02\ m$

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Q.4. A solution of glucose ($C_6H_{12}O_6$) in water is lebelled as 10% by weight. If density of this solution is 1.2 g/ml, then find the molarity of solution. (Molar Mass of Glucose = 180 g mol^{-1}).

Sol^n- Given that,

Conc. of glucose solution = 10% by weight

Density of glucose solution, d = 1.2 g/ml

Molar Mass of Glucose, M_B = 180 g mol^{-1}

Molarity of solution, M = ?

$$\because \quad \text{Molarity, } M = \frac{Mass\ of\ solute\ (g) \times 1000}{\text{GMM of Solute} \times \text{Volume of Solution (ml)}}$$

Here, 10% solution of glucose by weight means 10 g of glucose is present in 100 g of solution.

$$\because \quad \text{Density, } d = \frac{Mass}{Volume}$$

$$\therefore \quad \text{Volume of 100 g glucose solution} = \frac{Mass\ of\ Solution}{Density\ of\ solution}$$

$$= \frac{100\ g}{1.2\ g/ml}$$

$$= 83.33\ ml$$

$$\therefore \quad \text{Molarity, } M = \frac{10\ g \times 1000}{180\ g \times 83.33\ ml}$$

$$= 0.67\ M$$

Q.5. 18 g of glucose ($C_6H_{12}O_6$) is dissolved in 100 g of water. Find the mole fraction of glucose in solution. (Molar Mass of Glucose = 180 g mol^{-1}).

Sol^n- Given that,

Mass of glucose = 18 g

Molar Mass of Glucose, $M_{glucose}$ = 180 g mol^{-1}

Mass of water = 100 g

Molar Mass of water, M_{water} = 18 g mol^{-1}

Mole fraction of glucose, $X_{glucose}$ = ?

$$\because \quad \text{Mole Fraction of solute, } X_{glucose} = \frac{\text{No.of moles of glucose}}{\text{Total number of moles in solution}}$$

$$= \frac{n_{glucose}}{n_{glucose} + n_{water}}$$

$$= \frac{W_{glucose}/M_{glucose}}{W_{glucose}/M_{glucose} + W_{water}/M_{water}}$$

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$$= \frac{18\ g/180\ g}{18\ g/180\ g + 100g/18\ g}$$

$$= 0.018$$

Q.6. **Vapour Pressure of pure water is 40 mm Hg. On adding a nonvolatile solute, it reduces by 4 mm Hg. Find the molality of solution.**

Sol^n- Given that,

Vapour Pressure of pure water $(p^0_A) = 40$ mm Hg

Vapour Pressure of solution $(p_A) = (40 - 4)$ mm Hg

$= 36$ mm Hg

Molality of solution, $m = ?$

From Raoult's law,

$$P_A = p^0_A X_A$$

$$\therefore \quad 36 = 40 \times X_A$$

$$\text{or} \quad X_A = \frac{36}{40}$$

$$= 0.9$$

$$\therefore \quad X_B = (1 - X_A)$$

$$= (1 - 0.9)$$

$$= 0.1$$

Again since, molality,

$$m = \frac{X_B \times 1000}{X_A \times M_A}$$

$$= \frac{0.1 \times 1000}{0.9 \times 18}$$

$$= 6.17\ m$$

Q.7. **The vapour pressure of water at 296 K is 19.8 mm Hg. 0.1 mole of glucose is dissolved in 178.2 g of water. Calculate the vapour pressure of resultant solution.**

Sol^n- Given that,

Vapour Pressure of pure water $(p^0_A) = 19.8$ mm Hg

No. of moles of solute, $(n_B) = 0.1$ moles

Mass of solvent (water), $(W_A) = 178.2$ g

Vapour Pressure of solution, $(p_A) = ?$

From Raoult's law,

$$P_A = p^0_A X_A$$

$$= 19.8 \times \frac{n_A}{n_B + n_A}$$

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Here, $n_A = \frac{178.2}{18}$

$= 9.9$ moles

$\therefore P_A = 19.8 \times \frac{9.9}{0.1 + 9.9}$

$= 19.6$ mm Hg

Q.8. At a given temperature, vapour pressure of ethanol and methanol are 44.5 and 88.7 mm Hg respectively. An ideal solution is prepared by mixing 60 g of ethanol with 40 g of methanol at same temperature. Calculate the total vapour pressure of the solution and mole fraction of methanol in the vapour.

Sol^n- Given that,

Vapour pressure of ethanol $(p^0_A) = 44.5$ mm Hg

Vapour pressure of methanol $(p^0_B) = 88.7$ mm Hg

Mass of ethanol $(W_A) = 60$ g

Mass of methanol $(W_B) = 40$ g

Total vapour pressure of solution, $(p_{Total}) = ?$

Mole fraction of methanol in the vapour $= ?$

From Raoult's law, for an ideal solution,

$P_{Total} = p^0_A X_A + p^0_B X_B$

$= 44.5 \times \frac{n_A}{n_B + n_A} + 88.7 \times \frac{n_B}{n_B + n_A}$

Here, $n_A = \frac{60}{46}$

$= 1.304$ moles

And $n_B = \frac{40}{32}$

$= 1.25$ moles

$\therefore P_{Total} = 44.5 \times \frac{1.304}{1.25 + 1.304} + 88.7 \times \frac{1.25}{1.25 + 1.304}$

$= (22.73 + 43.40)$ mm Hg

$= 66.13$ mm

$\because$ Mole fraction of a component in the vapour $= \frac{Partial\ pressure\ of\ component}{Total\ vapour\ pressure}$

$\therefore$ Mole fraction of a methanol in the vapour $= \frac{Partial\ pressure\ of\ methanol}{Total\ vapour\ pressure}$

$= \frac{43.40}{66.13}$

$= 0.6563$

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Q.9. **At a given temperature, vapour pressure of two liquids A and B are 120 and 180 mm Hg respectively. An ideal solution is prepared by mixing 2 moles of A and 3 moles of B. Calculate the vapour pressure of the solution at same temperature.**

Sol^{n}- **Given that,**

Vapour pressure of liquid A $(p^0_A) = 120$ **mm Hg**

Vapour pressure of liquid B $(p^0_B) = 180$ **mm Hg**

No. of moles of A $(n_A) = 2$ **moles**

No. of moles of B $(n_B) = 3$ **moles**

Total vapour pressure of solution, $(p_{Total}) = ?$

From Raoult's law, for an ideal solution,

$$P_{Total} = p^0_A X_A + p^0_B X_B$$

$$= 120 \times \frac{n_A}{n_B + n_A} + 180 \times \frac{n_B}{n_B + n_A}$$

$$\therefore \quad P_{Total} = 120 \times \frac{2}{2+3} + 180 \times \frac{3}{2+3}$$

$$= (48 + 108)\ \text{mm Hg}$$

$$= 156\ \text{mm Hg}$$

Q.10. **At 20^0C temperature, vapour pressure of two liquids A and B are 22 and 75 mm Hg respectively. A solution is prepared by mixing equal moles of A and B. Assuming the solution to be ideal, calculate the vapour pressure of the solution at same temperature.**

Sol^{n}- **Given that,**

Vapour pressure of liquid A $(p^0_A) = 22$ **mm Hg**

Vapour pressure of liquid B $(p^0_B) = 75$ **mm Hg**

No. of moles of A (n_A) = **No. of moles of B** (n_B)

Total vapour pressure of solution, $(p_{Total}) = ?$

Here since,

$$n_A = n_B$$

$$\therefore \quad X_A = X_B$$

$$= \frac{1}{2}$$

From Raoult's law, for an ideal solution,

$$P_{Total} = p^0_A X_A + p^0_B X_B$$

$$\therefore \quad P_{Total} = 22 \times \frac{1}{2} + 75 \times \frac{1}{2}$$

$$= (11 + 37.5)\ \text{mm Hg}$$

$$= 48.5\ \text{mm Hg}$$

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Q.11. 50 g of sucrose (molar mass = 342) is added to 500 g of water. If vapour pressure of pure water at 25^0C temperature is 23.8 mm Hg, find the lowering in vapour pressure of the solution.

Sol^n- Given that,

Vapour pressure of pure water $(p^0_A) = 23.8$ mm Hg

Mass of solute (sucrose), $W_B = 50$ g

Gram molecular mass of solute (sucrose), $M_A = 342$

Mass of solvent (water), $W_A = 500$ g

Lowering in vapour pressure, $(p^0_A - p_A) = ?$

From Raoult's law about relative lowering in vapour pressure,

$$\frac{p^0_A - p_A}{p^0_A} = X_B$$

$$\therefore \quad \frac{p^0_A - p_A}{p^0_A} = \frac{n_B}{n_B + n_A}$$

$$\text{or} \quad \frac{p^0_A - p_A}{p^0_A} = \frac{W_B/M_B}{W_B/M_B + W_A/M_A}$$

$$\therefore \quad (p^0_A - p_A) = p^0_A \times \frac{W_B/M_B}{W_B/M_B + W_A/M_A}$$

$$= 23.8 \times \frac{\frac{50\ g}{342\ g}}{\frac{50\ g}{342\ g} + \frac{500\ g}{18\ g}}$$

$$= 23.8 \times \frac{0.146}{0.146 + 27.78}$$

$$= 0.124 \text{ mm Hg}$$

Q.12. 50 g of sucrose (molar mass = 342) is added to 500 g of water. If vapour pressure of pure water at 25^0C temperature is 23.8 mm Hg, find the relative lowering in vapour pressure of the solution.

Sol^n- Given that,

Vapour pressure of pure water $(p^0_A) = 23.8$ mm Hg

Mass of solute (sucrose), $W_B = 50$ g

Gram molecular mass of solute (sucrose), $M_A = 342$

Mass of solvent (water), $W_A = 500$ g

Relative lowering in vapour pressure, $\frac{p^0_A - p_A}{p^0_A} = ?$

From Raoult's law about relative lowering in vapour pressure,

$$\frac{p^0_A - p_A}{p^0_A} = X_B$$

$$\therefore \quad \frac{p^0_A - p_A}{p^0_A} = \frac{n_B}{n_B + n_A}$$

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or $$\frac{p0_A - p_A}{p0_A} = \frac{W_B/M_B}{W_B/M_B + W_A/M_A}$$

$$= \frac{\frac{50\ g}{342\ g}}{\frac{50\ g}{342\ g} + \frac{500\ g}{18\ g}}$$

$$= \frac{0.146}{0.146 + 27.78}$$

$$= 0.0052 \text{ mm Hg}$$

Q.13. The vapour pressure of 5 % aqueous solution of a non-volatile solute at 373 K temperature is 745 mm Hg. If vapour pressure of pure water at 373 K temperature is 760 mm Hg, find the molecular mass of the solute.

Sol^{n}- Given that,

Vapour pressure of pure water at 373 K, $p^0_A = 760$ mm Hg

Vapour pressure of solution, at 373 K, $p_A = 745$ mm Hg

Mass of solute, $W_B = 5$ g

Mass of solvent (water), $W_A = (100 - 5)$ g

$= 95$ g

Gram molecular mass of solute, $M_B = ?$

From Raoult's law about relative lowering in vapour pressure,

$$\frac{p0_A - p_A}{p0_A} = \frac{W_B \times M_A}{M_B \times W_A}$$

$$\therefore \quad M_B = \frac{p^0_A}{p^0_A - p_A} \times \frac{W_B \times M_A}{W_A}$$

$$= \frac{760}{760 - 745} \times \frac{5\ g \times 18}{95}$$

$$= 47.05$$

Q.14. The vapour pressure of a solution containing 30 g of a non-volatile solute in 90 g of water at 298 K temperature is 21.85 mm Hg. When 18 g of water is added to this solution the vapour pressure of solution becomes 22.15 mm Hg. Find the molecular mass of the solute and the vapour pressure of water at 298 K temperature.

Sol^{n}- Given that,

Vapour pressure of first solution at 298 K = 21.85 mm Hg

Vapour pressure of second solution at 298 K = 22.15 mm Hg

Mass of solute, $W_B = 30$ g

Mass of water in first solution = 90 g

Mass of water in second solution = (90 + 18) g

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= 108 g

Molecular mass of solute, M_B = ?

Vapour pressure of pure water at 298 K, p^0_A = ?

From Raoult's law about relative lowering in vapour pressure,

$$\frac{p0_A - p_A}{p0_A} = \frac{W_B \times M_A}{M_B \times W_A}$$

Here, for first solution, $\frac{p0_A - 21.85}{p0_A} = \frac{30 \times 18}{M_B \times 90}$ ------- (1)

For second solution, $\frac{p0_A - 22.15}{p0_A} = \frac{30 \times 18}{M_B \times 108}$ ------- (2)

Dividing eq. (1) by eq. (2) we get,

$$\frac{p0_A - 21.85}{p0_A} \times \frac{p0_A}{p0_A - 22.15} = \frac{30 \times 18}{M_B \times 90} \times \frac{M_B \times 108}{30 \times 18}$$

or $$\frac{p0_A - 21.85}{p0_A - 22.15} = \frac{108}{90}$$

$\therefore$ p^0_A = 23.65 mm Hg

Putting this value of p^0_A in eq. (1) we get,

$$\frac{23.65 - 21.85}{23.65} = \frac{30 \times 18}{M_B \times 90}$$

$\therefore$ M_B = 78.83

Q.15. 0.1 m aqueous solution of glucose boils at 100.052 ^{0}C temperature. Find molal boiling point elevation constant for water.

Sol^n- Given that,

Molality of solution, m = 0.1

Boiling point of solution, T_b = 100.052 ^{0}C

Molal boiling point elevation constant for water = ?

$\because$ Elevation in boiling point,

$\Delta T_b = K_b \times m$

$\therefore$ $K_b = \frac{\Delta T_b}{m}$

Here, $\Delta T_b = (T_b - T^0_b)$

$= (100.052 - 100)^0$C

$= 0.052\ ^0$C

$\therefore$ $K_b = \frac{0.052}{0.1}$

$= 0.52\ Km^{-1}$

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Q.16. 5 g of a nonvolatile solute is dissolved to 50 g of benzene. If elevation in boiling point of solution is 1^0C and molal boiling point elevation constant for benzene is 2.53 Km^{-1}, find the molecular mass of solute.

Sol^n- Given that,

Mass of solute, W_B = 5 g

Mass of solvent (benzene), W_A = 50 g

Elevation in boiling point, ΔT_b = 1^0C

Molal boiling point elevation constant for benzene = 2.53 Km^{-1}

Gram molecular mass of solute, M_B = ?

$\because$ Elevation in boiling point,

$$\Delta T_b = \frac{1000 K_b . W_B\,(g)}{M_B . W_A\,(g)}$$

$$\therefore \quad M_B = \frac{1000 K_b . W_B\,(g)}{\Delta Tb . W_A\,(g)}$$

$$= \frac{1000 \times 2.53 \times 5}{1 \times 50}$$

$$= 253 \text{ g mol}^{-1}$$

Q.17. 0.36 g of glucose is dissolved to 100 g of water. If molal boiling point elevation constant for water is 0.52 Km^{-1}, find the boiling point of the solution.

Sol^n- Given that,

Mass of solute (glucose), W_B = 0.36 g

Mass of solvent (water), W_A = 100 g

Molal boiling point elevation constant for water = 0.52 Km^{-1}

Boiling point of the solution, (T_b) = ?

Here, elevation in boiling point,

$$\Delta T_b = \frac{1000 K_b . W_B\,(g)}{M_B . W_A\,(g)}$$

$$= \frac{1000 \times 0.52 \times 0.36}{180 \times 100}$$

$$= 0.01 \text{ K}$$

$\because$ $\Delta T_b = (T_b - T^0_b)$

$\therefore$ $T_b = T^0_b + \Delta T_b$

$= 373 \text{ K} + 0.01 \text{ K}$

$= 373.01 \text{ K}$

Q.18. Molar enthalpy of vaporisation of water at 373 K temperature is 40.585 kJ mol^{-1}. Find molal boiling point elevation constant of water.

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Sol^{n}- Given that,

Molar enthalpy of vaporisation of water at 373 K = 40.585 kJ mol^{-1}

= 40585 J mol^{-1}

Molal boiling point elevation constant for water = ?

∵ Molal boiling point elevation constant,

$$K_b = \frac{R \times (T^0_b)^2}{1000 \times \Delta H_V}$$

Where, ΔH_V = Latent heat of vaporization

Here, Latent heat of vaporization of water (per gram),

$$\Delta H_V = \frac{Molar\ Latent\ heat}{\text{Molar mass}}$$

$$= \frac{40585\ J}{18}$$

$$= 2254\ J\ g^{-1}$$

$$\therefore \quad K_b = \frac{8.314 \times (373)^2}{1000 \times 2254}$$

$$= 0.513\ K\ kg\ mol^{-1}$$

Q.19. Find the freezing point of a solution containing 0.5 g of glucose (molar mass = 180) dissolved in 80 g of water. K_f for water is 1.86 Km^{-1}.

Sol^{n}- Given that,

Mass of solute (glucose), W_B = 0.5 g

Mass of solvent (water), W_A = 80 g

Molal freezing point depression constant for water, K_f = 1.86 Km^{-1}

Freezing point of a solution, T_f = ?

∵ Depression in freezing point,

$$\Delta T_f = \frac{1000 K_f . W_B\ (g)}{M_B . W_A\ (g)}$$

$$= \frac{1000 \times 1.86 \times 0.5}{180 \times 80}$$

$$= 0.065\ K$$

Again since, $\Delta T_f = T^0_f - T_f$

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

$= (273 - 0.065)\ K$

$= 272.935\ K$

Q.20. find the mass of ethylene glycol (molar mass = 62) added to 5 Kg of water to lower its freezing point from 0^0C to -10^0C. K_f for water is 1.86 Km^{-1}.

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Sol^{n}- Given that,

Mass of solvent (water), W_A = 5 Kg = 5000 g

Molal freezing point depression constant for water, K_f = 1.86 Km^{-1}

Freezing point of a solution, T_f = -10^0C

Mass of solute (Ethylene glycol), W_B = ?

Here, depression in freezing point,

$$\Delta T_f = T^0_f - T_f$$
$$= [\,0 - (-10)\,]\ ^0C$$
$$= 10\ ^0C$$

Now since, depression in freezing point,

$$\Delta T_f = \frac{1000 K_f . W_B\,(g)}{M_B . W_A\,(g)}$$

$$\therefore \quad W_B = \frac{\Delta T_f . M_B . W_A\,(g)}{1000 K_f}$$
$$= \frac{10 \times 62 \times 5000}{1000 \times 1.86}$$
$$= 1666.67\ g$$
$$= 1.67\ Kg$$

Q.21. 5 g of a nonvolatile solute is dissolved to 50 g of water. If depression in freezing point of solution is 1^0C and molal freezing point depression constant for water is 1.86 Km^{-1}, find the molecular mass of solute.

Sol^{n}- Given that,

Mass of solute, W_B = 5 g

Mass of solvent (water), W_A = 50 g

Depression in Freezing point, ΔT_f = 1^0C

Molal freezing point depression constant for water, K_f = 1.86 Km^{-1}

Gram molecular mass of solute, M_B = ?

$\because$ Depression in freezing point,

$$\Delta T_f = \frac{1000 K_f . W_B\,(g)}{M_B . W_A\,(g)}$$

$$\therefore \quad M_B = \frac{1000 K_f . W_B\,(g)}{\Delta T_f . W_A\,(g)}$$
$$= \frac{1000 \times 1.86 \times 5}{1 \times 50}$$
$$= 186\ g\ mol^{-1}$$

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Q.22. **Elevation in boiling point of an aqueous solution of urea is 0.128 K. Find the freezing point of the solution. K_f and K_b for water are 1.86 Km^{-1} and 0.512 Km^{-1} respectively.**

Sol^n- **Given that,**

Elevation in boiling point, $\Delta T_b = 0.128$ K

Molal freezing point depression constant for water, $K_f = 1.86\ Km^{-1}$

Molal boiling point elevation constant for water, $K_b = 0.512\ Km^{-1}$

Freezing point of the solution, T_f = ?

∵ Elevation in boiling point,

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

$$\therefore \quad 0.128 = 0.512 \times m$$

$$\text{or} \quad m = \frac{0.128}{0.512}$$

$$= 0.25\ m$$

Again since,

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

$$= 1.86 \times 0.25$$

$$= 0.465\ K$$

∵ Depression in freezing point,

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

$$\therefore \quad T_f = T^0_f - \Delta T_f$$

$$= [273.15 - (0.465)]\ K$$

$$= 272.685\ K$$

Q.23. **An aqueous solution freezes at –0.186 ^{0}C. Find the elevation in boiling point of the solution. K_f and K_b for water are 1.86 Km^{-1} and 0.512 Km^{-1} respectively.**

Sol^n- **Given that,**

Freezing point of aqueous solution, $T_f = -0.186\ ^0C$

Molal freezing point depression constant for water, $K_f = 1.86\ Km^{-1}$

Molal boiling point elevation constant for water, $K_b = 0.512\ Km^{-1}$

Elevation in boiling point, ΔT_b = ?

∵ Depression in freezing point,

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

$$= [0 - (-0.186)]\ ^0C$$

$$= +0.186\ ^0C$$

Again since,

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

$$\therefore \quad 0.186\ ^0C = 1.86 \times m$$

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$$\text{or} \qquad m = \frac{0.186}{1.86}$$

$$= 0.1m$$

Now since, elevation in boiling point,

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

$$= 0.512 \times 0.1$$

$$= 0.0512\ ^0C$$

Q.24. 100 cm^3 of an aqueous solution contains 5.0 g of an unknown solute. If the osmotic pressure of the solution at 298 K temperature is 2.5 atm, then find the molecular mass of the solute.

Sol^n- Given that,

Mass of solute, $W_B = 5.0$ g

Volume of solution, $V = 100\ cm^3 = 0.1$ L

Absolute temperature, $T = 298$ K

Molecular mass of solute, $M_B = ?$

$\because$ Osmotic pressure of solution,

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

Where, S = Solution constant

$= 0.0821$ L-atm $K^{-1}mol^{-1}$

$\therefore$ Molecular mass of solute,

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

$$= \frac{5.0}{2.5 \times 0.1} \times 0.0821 \times 298$$

$$= 489.3\ g\ mol^{-1}$$

Q.25. A 5% aqueous solution (m/v) cane sugar is isotonic with 0.877% aqueous solution of urea (m/v). If the molecular mass of cane sugar is 342, then find the molecular mass of urea.

Sol^n- Given that,

Conc. of sugar solution = 5 % (m/v)

Conc. of urea solution = 0.877 % (m/v)

Molecular mass of cane sugar, $M_{sugar} = 342$

Molecular mass of urea, $M_{urea} = ?$

Here, for isotonic solutions,

π_1 (cane sugar) = π_2 (urea)

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Again since, $\pi = \frac{W_B}{M_B V} ST$

Where, S = Solution constant

= 0.0821 L-atm $K^{-1}mol^{-1}$

$\therefore \quad \frac{W_1}{M_1 V_1} ST = \frac{W_2}{M_2 V_2} ST$

or $\quad \frac{W_1}{M_1 V_1} = \frac{W_2}{M_2 V_2}$

or $\quad \frac{5}{342 \times 0.1} = \frac{0.877}{M_2 \times 0.1}$

or $\quad M_2 = \frac{342 \times 0.1 \times 0.877}{5 \times 0.1}$

$= 59.987 \text{ g mol}^{-1}$

Q.26. Osmotic pressure of blood is 8.21 atm at 37^0C. How much glucose (molar mass = 180 g mol^{-1}) should be added per litre for preparing an intravenous injection which is isotonic to the blood.

Sol^n- Given that,

Osmotic pressure of blood = 8.21 atm

Volume of solution, V = 1 L

Temperature = 37^0C = (37 + 273) K = 310 K

Molecular mass of solute (glucose), M_B = 180 g mol^{-1}

Mass of solute (glucose), W_B = ?

Here, since intravenous injection is isotonic to the blood. Hence,

π (Injection) = π (Blood)

= 8.21 atm

Again since, $\pi = \frac{W_B}{M_B V} ST$

Where, S = Solution constant

= 0.0821 L-atm $K^{-1}mol^{-1}$

$\therefore$ Mass of solute,

$W_B = \frac{\pi \times M_B \times V}{ST}$

$= \frac{8.21 \times 180 \times 1}{0.0821 \times 310}$

= 58.06 g

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Q.27. 100 ml 1.8% aqueous solution of glucose (molar mass = 180 g mol^{-1}) is mixed with 100 ml of 0.6 % aqueous solution of urea (molar mass = 60 g mol^{-1}) at at 27^0C temperature. Find the osmotic pressure of resultant solution.

Sol^n- Given that,

Conc. of glucose solution = 1.8 % (m/v)

Conc. of urea solution = 0.6 % (m/v)

Molecular mass of glucose, $M_{glucose}$ = 180 g mol^{-1}

Molecular mass of urea, M_{urea} = 60 g mol^{-1}

Temperature, T = (27 + 273) K = 300 K

Osmotic pressure of resultant solution (π_R) = ?

Here, Osmotic pressure of resultant solution,

$$\pi_R = \frac{\pi_1 V_1 + \pi_2 V_2}{V_1 + V_2}$$

Again since, osmotic pressure,

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

Where, S = Solution constant

= 0.0821 L-atm $K^{-1}mol^{-1}$

$$\therefore \quad \pi_1 = \frac{1.8}{180 \times 0.1} \times 0.0821 \times 300$$

$$= 2.463 \text{ atm}$$

$$\text{And} \quad \pi_2 = \frac{0.6}{60 \times 0.1} \times 0.0821 \times 300$$

$$= 2.463 \text{ atm}$$

$$\therefore \quad \pi_R = \frac{(2.463 \times 0.1) + (2.463 \times 0.1)}{0.1 + 0,1}$$

$$= 2.463 \text{ atm}$$

Q.28. The solubility of pure N_2 gas at 25^0C and 1 atm is 6.8×10^{-4} mol L^{-1}. Find the concentration of N_2 in water under atmospheric condition. The partial pressure of N_2 in atmosphere is 0.78 atm.

Sol^n- Given that,

Solubility of N_2 gas at 25^0C and 1 atm, S = 6.8×10^{-4} mol L^{-1}

Partial pressure of N_2 in atmosphere p_{N_2} = 0.78 atm

Conc. of N_2 in water under atmospheric condition = ?

From Henry's law, S α p

$$\therefore \quad \frac{S_1}{S_2} = \frac{p_1}{p_2}$$

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$$\therefore \quad s_2 = s_1 \times \frac{p_2}{p_1}$$

$$= 6.8\times10^{-4} \text{ mol } L^{-1} \times \frac{0.78}{1}$$

$$= 5.304\times10^{-4} \text{ mol } L^{-1}$$

Q.29. 27 g acetylene is dissolved in one liter of acetone at 1 atm pressure. If the partial pressure of acetylene is increased to 12 atm, then find the solubility of acetylene in acetone.

Sol^n- Given that,

Solubility of acetylene at 1 atm, $S_1 = 27 \text{ g } L^{-1}$

Solubility of acetylene at 12 atm, $S_2 = ?$

From Henry's law, $S \ \alpha \ p$

$$\therefore \quad \frac{S_1}{S_2} = \frac{p_1}{p_2}$$

$$\therefore \quad s_2 = s_1 \times \frac{p_2}{p_1}$$

$$= 27 \text{ gL}^{-1} \times \frac{12}{1}$$

$$= 324 \text{ gL}^{-1}$$

Q.30. What concentration of N_2 should be present in a glass of water at room temperature? Assume a temperature of 25^0C, a total pressure of 1 atm and mole fraction of N_2 in air to be 0.78. K_H of $N_2 = 8.42\times10^{-7}$ M/mm Hg.

Sol^n- Given that,

Total pressure = 1 atm

= 760 mm Hg

K_H of N_2 = 8.42×10^{-7} mm Hg

Mole fraction of N_2 in air, $x_{N_2} = 0.78$

Conc. of N_2 in a glass of water at room temperature, $x_{N_2} = ?$

From Henry's law,

Conc. of N_2 = $K_H \times p_{N_2}$

= $K_H \times x_{N_2} \times p_{Total}$

= 8.42×10^{-7} M/mm Hg × 0.78 × 760 mm Hg

= 4.99×10^{-4} mol L^{-1}

Q.31. O_2 is bubbled in water at 298 K temperature. If partial pressure of O_2 is 0.98 bar, then find the solubility of O_2 in gL^{-1}. Henry's law constant for O_2 is 34.84 K bar at 298 K temperature.

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Sol^{n}- Given that,

K_H of O_2 = 34.84 K bar

= 34.84×10^3 bar

p_{O_2} = 0.98 bar

Solubility of O_2 in gL^{-1} = ?

From Henry's law,

$$p_{O_2} = K_H \times x_{O_2}$$

$$\therefore \quad x_{O_2} = \frac{p_{O_2}}{K_H}$$

$$= \frac{0.98 \text{ bar}}{34.84 \times 10^3 \text{ bar}}$$

$$= 2.8 \times 10^{-5}$$

Again since, $x_{O_2} = \frac{n_{O_2}}{n_{O_2} + n_{H_2O}}$

For dilute solution, $x_{O_2} \approx \frac{n_{O_2}}{n_{H_2O}}$

$$\therefore \quad n_{O_2} = x_{O_2} \times n_{H_2O}$$

$$= 2.8 \times 10^{-5} \times 55.55$$

$$= 1.56 \times 10^{-3} \text{ Mole}$$

$\therefore$ Conc. of O_2 in gL^{-1} = n_{O_2} × Molar mass of O_2

$$= 1.56 \times 10^{-3} \times 32$$

$$= 0.05 \text{ gL}^{-1}$$

Q.32. Henry's law constant for CO_2 dissolved in water is 1.67×10^8 Pa at 298 K. Calculate the quantity of CO_2 in in 1 L of soda water when packed under 2.5 atm CO_2 pressure at 298 K.

Sol^{n}- Given that,

K_H of CO_2 = 1.67×10^8 Pa

p_{CO_2} = 2.5 atm

= $2.5 \times 1.01 \times 10^5$ Pa

Quantity of CO_2 in in 1 L of soda water = ?

From Henry's law,

$$p_{CO_2} = K_H \times x_{CO_2}$$

$$\therefore \quad x_{CO_2} = \frac{p_{CO_2}}{K_H}$$

$$= \frac{2.5 \times 1.01 \times 10^5 \text{ Pa}}{1.67 \times 10^8 \text{ Pa}}$$

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$= 1.517 \times 10^{-3}$

Again since, $x_{CO_2} = \frac{n_{CO_2}}{n_{CO_2} + n_{H_2O}}$

For dilute solution, $x_{CO_2} \approx \frac{n_{CO_2}}{n_{H_2O}}$

$\therefore \quad n_{CO_2} = x_{CO_2} \times n_{H_2O}$

$= 1.517 \times 10^{-3} \times 55.55$

$= 8.426 \times 10^{-2}$ Mole

$\therefore$ Mass of CO_2 = n_{CO_2} × Molar mass of CO_2

$= 8.426 \times 10^{-2} \times 44$

$= 3.707$ g

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