

SOLUTIONS

SOLVED SUBJECTIVE EXERCISE

Very Short Answer Type Questions (1 mark)

1. State the condition resulting in reverse osmosis.

Ans. Reverse osmosis occurs when a pressure larger than the osmotic pressure applied to the solution.

2. What is the significance of Henry's law constant K_H ?

Ans. Higher the value of Henry's law constant K_H at a given pressure, lower is the solubility of the gas in the liquid.

3. When would dissolution of solute in a solvent lead to the liberation of energy?

Ans. Dissolution of solute in a solvent would lead to the liberation of energy when the solute-solvent attraction, is greater than the intermolecular forces of attraction present among the molecules of solutes and among the molecules of solvent.

4. Why is the elevation in boiling point of water different in the following two solutions?

(a) 0.1 molar NaCl solution (b) 0.1 molar sugar solution.

Ans. NaCl solution is an electrolyte. It dissociates to give Na^+ and Cl^- ions in the solution. Sugar is a non-electrolyte. It does not dissociate in the solution. So, number of particles is greater in 0.1 M NaCl solution than in 0.1 M sugar solution. Thus, elevation in boiling point is greater in case of NaCl solution.

5. Explain the solubility rule "like dissolves like" in terms of intermolecular forces that exist in solutions.

Ans. A substance (solute) dissolves in a solvent if the intermolecular interactions are similar in both the components; for example, polar solutes dissolve in polar solvents and non polar solutes in non polar solvents thus we can say "like dissolves like".

6. Why is osmotic pressure of 1 M KCl higher than 1 M urea solution?

Ans. This is because KCl dissociates to give K^+ and Cl^- ions while urea being a molecular solid does not dissociate into ions in the solution.

7. Why is an increase in temperature observed on mixing chloroform with acetone?

Ans. The bonds between chloroform and acetone molecules are dipole-dipole interactions but on mixing, the chloroform and acetone molecules start forming hydrogen bonds which are stronger bonds resulting in the release of energy. This gives rise to an increase in temperature.

8. What is the van't Hoff factor for a compound which undergo dimerisation in an organic solvent?

Ans. $2X \longrightarrow X_2$

$$i = \frac{\text{Number of moles of particle after association}}{\text{Number of moles of particles before association}} = \frac{1}{2}$$

9. How does Henry's constant (K_H) of a gas in a particular solvent vary with temperature?

Ans. According to Henry's law $p_A = K_H \cdot x_A$ with increase of temperature, solubility of a gas decreases at the same partial pressure, i.e., x_A decreases. Hence, at constant p_A , K_H increases. Thus, K_H increases with increase of temperature.

10. At the same temperature, hydrogen is more soluble in water than helium. Which of them will have a higher value of K_H and why?

Ans. As $p_A = K_H \cdot x_A$ Thus, at constant temperature, for the same partial pressure of different gases, $x_A \propto 1/K_H$. In other words, solubility is inversely proportional to Henry's constant of the gas. Higher the value of K_H , lower is the solubility of the gas.

As H_2 is more soluble than helium, H_2 will have lower value K_H than that of helium.

17. 45 g 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.

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

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

Molality, $m = \frac{\text{moles of ethylene glycol}}{\text{Mass of water in kg}} = \frac{0.73\text{ mol}}{0.60\text{ kg}} = 1.2\text{ mol kg}^{-1}$

(a) Therefore, freezing point depression,

$$\Delta T_f = K_f m = 1.86\text{ K kg mol}^{-1} \times 1.2\text{ mol kg}^{-1} = \mathbf{2.2\text{ K}}$$

(b) Freezing point of the aqueous solution = $T_f^\circ - \Delta T_f = 273.15\text{ K} - 2.2\text{ K} = \mathbf{270.95\text{ K}}$

18. If N_2 gas is bubbled through water at 293 K, how many millimoles of N_2 gas would dissolve in 1 litre of water? Assume that N_2 exerts a partial pressure of 0.987 bar. Given that Henry's law constant for N_2 at 293 K is 76.48 k bar.

Ans. The solubility of gas is related to the mole fraction in aqueous solution. The mole fraction of the gas in the solution is calculated by applying Henry's law. Thus:

$$x(\text{Nitrogen}) = \frac{P(\text{nitrogen})}{K_H} = \frac{0.987\text{ bar}}{76480\text{ bar}} = 1.29 \times 10^{-5}$$

As 1 litre of water contains 55.5 mol of it, therefore if n represents the number of moles of N_2 in solution,

$$x(\text{Nitrogen}) = \frac{n\text{ mol}}{n\text{ mol} + 55.5\text{ mol}} = \frac{n}{55.5} = 1.29 \times 10^{-5} \quad (\text{n in denominator is neglected as it is } \ll 55.5)$$

Thus $n = 1.29 \times 10^{-5} \times 55.5\text{ mol} = 7.16 \times 10^{-4}\text{ mol} = \mathbf{0.716\text{ millimoles}}$.

19. 0.5 g KCl was dissolved in 100 g water and the solution originally at 20°C , froze at -0.24°C . Calculate the percentage ionization of salt. K_f per 1000 g of water = 1.86°C .

Ans. Observed mol. mass is obtained from the given data, i.e.,

$$w_2 = 0.5\text{ g}, w_1 = 100\text{ g}, M_1 = 18\text{ g mol}^{-1}\text{ for H}_2\text{O}$$

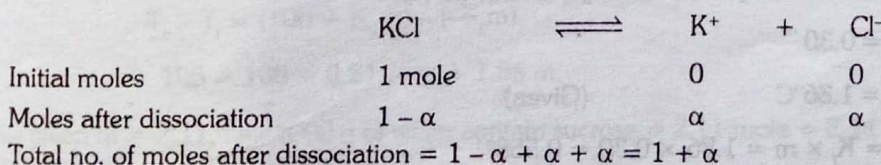
$$\Delta T_f = 0 - (-0.24) = 0.24^\circ\text{C}$$

$$M_2 = \frac{1000K_f w_2}{\Delta T_f \times w_1} = \frac{1000\text{g kg}^{-1} \times 1.86\text{K kg mol}^{-1} \times 0.5\text{g}}{0.24\text{K} \times 100\text{g}} = 38.75\text{ g mol}^{-1}$$

Calculated (theoretical) mol mass of KCl = $39 + 35.5 = 74.5\text{ g mol}^{-1}$

$$\therefore i = \frac{\text{Calculated mol mass}}{\text{Observed mol mass}} = \frac{74.5}{38.75} = 1.92$$

Now, KCl dissociates as



$$\therefore i = \frac{1 + \alpha}{1} \text{ or } \alpha = i - 1 = 1.92 - 1 = 0.92$$

Percentage ionization = $0.92 \times 100 = \mathbf{92\%}$

20. 19.5g of CH_2FCOOH is dissolved in 500g of water. The depression in the freezing point observed is 1.0°C . Calculate the Van't Hoff factor and dissociation constant of fluoroacetic acid. K_f for water is 1.86K kg mol^{-1} .

Ans. Here, $w_2 = 19.5\text{ g}$, $w_1 = 500\text{ g}$, $K_f = 1.86\text{ kg mol}^{-1}$, $(\Delta T_f)_{\text{obs}} = 1.0^\circ$

$$\therefore M_2 (\text{observed}) = \frac{1000K_f w_2}{w_1 \Delta T_f} = \frac{(1000\text{g kg}^{-1})(1.86\text{K kg mol}^{-1})(19.5\text{g})}{(500\text{g})(1.0\text{K})} = 72.54\text{ g mol}^{-1}$$

$$M_2 (\text{calculated}) \text{ for } \text{CH}_2\text{FCOOH} = 14 + 19 + 45 = 78\text{ g mol}^{-1}$$

$$\text{van't Hoff factor (i)} = \frac{(M_2)_{\text{cal}}}{(M_2)_{\text{obs}}} = \frac{78}{72.54} = \mathbf{1.0753}$$

Calculation of dissociation constant. Suppose degree of dissociation at the given concentration is α .

| | | | | |
|---------|--|-----------|-----------|-------------------------|
| Then | $\text{CH}_2\text{FCOOH} \rightleftharpoons \text{CH}_2\text{FCOO}^- + \text{H}^+$ | | | |
| Initial | $C\text{ Mol L}^{-1}$ | 0 | 0 | |
| At eqm. | $C(1 - \alpha)$ | $C\alpha$ | $C\alpha$ | Total = $C(1 + \alpha)$ |

$$\therefore i = \frac{C(1 + \alpha)}{C} = 1 + \alpha \quad \text{or} \quad \alpha = i - 1 = 1.0753 - 1 = 0.0753$$

$$K_a = \frac{[\text{CH}_2\text{FCOO}^-][\text{H}^+]}{[\text{CH}_2\text{FCOOH}]} = \frac{C\alpha \cdot C\alpha}{C(1 - \alpha)} = \frac{C\alpha^2}{(1 - \alpha)}$$

Taking volume of the solution as 500 mL,

$$C = \frac{19.5}{78} \times \frac{1}{500} \times 1000 = 0.5\text{ M}$$

$$\therefore K_a = \frac{C\alpha^2}{1 - \alpha} = \frac{(0.5)(0.0753)^2}{1 - 0.0753} = \mathbf{3.07 \times 10^{-3}}$$

21. The average osmotic pressure of human blood is 7.7 atm at 40°C .

- (a) What would be the total concentration of the various solutes in the blood?
 (b) Assuming the concentration to be essentially the same as the molality, find the freezing point of blood (K_f for water = 1.86°C).

Ans. (a) We are given that $P = 7.7\text{ atm}$, $T = 40^\circ\text{C} = 40 + 273 = 313\text{ K}$

$$R = 0.0821\text{ litre atm/degree/mole}$$

According to van't Hoff equation, $\pi = CRT$

$$C = \frac{\pi}{RT} = \frac{7.7\text{ atm}}{0.0821\text{L atm K}^{-1}\text{mol}^{-1} \times 313\text{K}} = \mathbf{0.30\text{ mole/litre}}$$

- (b) Taking the molar concentration as equal to molality (Given), we have

$$M = 0.30$$

$$K_f = 1.86^\circ\text{C} \quad (\text{Given})$$

$$\Delta T_f = K_f \times m = 1.86 \times 0.30 = 0.558^\circ$$

$$\therefore \text{Freezing point of blood} = 0^\circ\text{C} - 0.558^\circ = \mathbf{-0.558^\circ\text{C}}$$

22. Phenol associates in benzene to a certain extent to form a dimer. A solution containing 20 g of phenol in 1.0 kg of benzene has its freezing point lowered by 0.09 K. Calculate the fraction of phenol that has dimerised (Given K_f for benzene = 5.1 K m⁻¹)

Ans. Here, $n = 2$ because phenol forms dimer on association.

Given, $W_2 = 20 \text{ g}$, $W_1 = 1 \text{ kg} = 1000 \text{ g}$

$$\Delta T_f = 0.09 \text{ K}$$

$$K_f = 5.1 \text{ K m}^{-1}$$

$$\Delta T_f = \frac{K_f \times W_2 \times 1000}{M_2 W_1}$$

$$M_2 = \frac{K_f \times W_2 \times 1000}{\Delta T_f \times W_1}$$

$$M_2 = \frac{5.1 \text{ K m}^{-1} \times 20 \text{ g} \times 1000}{0.09 \text{ K} \times 1000 \text{ g}}$$

M_2 (observed) = 117.80

M_2 (calculated)

$$C_6H_5OH = 6 \times 12 + 6 \times 1 + 16 = 94 \text{ g mol}^{-1}$$

$$i = \frac{\text{normal (calculated) molar mass}}{\text{Observed molar mass}}$$

$$= \frac{94}{117.80} = 0.80$$



$$\alpha = \frac{1-i}{\left(1-\frac{1}{n}\right)} = \frac{1-0.80}{1-\frac{1}{2}} = \frac{0.20}{0.5} = 0.40$$

$\alpha = 40\%$

23. How many grams of sucrose (M wt. = 342) should be dissolved in 100g water in order to produce a solution with 105.0°C difference between the freezing point and the boiling point?

($K_b = 1.86^\circ \text{C/m}$, $K_f = 0.51^\circ \text{C/m}$)

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

B. pt. of solution (T_b) = 100 + $\Delta T_b = 100 + K_b m$

f. pt. of solution (T_f) = 0 - $\Delta T_f = 0 - K_f m$

$$T_b - T_f = (100 + K_b m) - (-K_f m)$$

$$105 = 100 + 0.51 \times m + 1.86 m$$

This gives $m = 2.11$, i.e., 1000 g of water contain sucrose = 2.11 mole = 2.11 × 342 g

\therefore Mass of sucrose to be dissolved in 100 g of water = $\frac{2.11 \times 342}{1000} \times 100 = 72 \text{ g}$

24. Heptane and octane form ideal solution. At 373 K, the vapour pressures of the two liquid components are 105.2 kPa and 46.8 kPa respectively. What will be the vapour pressure of a mixture of 26.0 g of heptane and 35.0 g of octane?

Ans. Given $P_b^\circ = 105.2 \text{ kPa}$

$$P_b^\circ = 46.8 \text{ kPa}$$

Molar mass of heptane (C_7H_{16}) = 100 g mol^{-1}

(As $C_7H_{16} = 7 \times 12 + 16 = 84 + 16 = 100$)

Molar mass of octane (C_8H_{18}) = 114 g mol^{-1}

($C_8H_{18} = 8 \times 12 + 18 = 96 + 18 = 114$)

Number of moles of 26.0 g heptane

$$= \frac{26.0 \text{ g}}{100 \text{ g mol}^{-1}} = 0.26 \text{ mol}$$

Number of moles of 35.0 g octane

$$= \frac{35.0 \text{ g}}{114 \text{ g mol}^{-1}} = 0.31 \text{ mol}$$

$$x(\text{heptane}) = \frac{0.26 \text{ g}}{0.26 + 0.31} = 0.456$$

$$x(\text{octane}) = 1 - 0.456 = 0.544$$

$$p(\text{heptane}) = 0.456 \times 105.2 \text{ kPa} = 47.97 \text{ kPa}$$

$$p(\text{octane}) = 0.544 \times 46.8 \text{ kPa} = 25.46 \text{ kPa}$$

$$P_{\text{total}} = 47.97 + 25.46 = \mathbf{73.43 \text{ kPa}}$$

Short Answer Type Questions (3 mark)

25. An aqueous solution containing 1.248 g of barium chloride (molar mass = $208.34 \text{ g mol}^{-1}$) in 100 g of water boils at 100.0832°C . Calculate the degree of dissociation of barium chloride. (K_b for water = $0.52 \text{ K kg mol}^{-1}$)

Ans. Moles of solute = $\frac{\text{Mass of solute}}{\text{Molar mass}} = \frac{1.248}{208.34} = 0.00599$

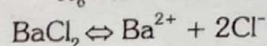
Mass of solvent = $100 \text{ g} = 0.1 \text{ kg}$

Molarity of solution, $m = \frac{0.00599}{0.1} = 0.0599 \text{ m}$

$\Delta T_b = 100.0832 - 100 = 0.0832^\circ\text{C}$

$K_b = 0.52 \text{ K kg mol}^{-1}$

$$i = \frac{\Delta T_b}{K_b \times m} = \frac{0.0832}{0.52 \times 0.0599} = 2.67$$



$n = 3$

Degree of dissociation,

$$\alpha = \frac{i-1}{n-1} = \frac{2.67-1}{3-1} = \mathbf{0.835}$$

26. A 0.1539 molal aqueous solution of cane sugar (mol. mass = 342 mol⁻¹) has a freezing point of 271 K while the freezing point of pure water is 273.15 K. What will be the freezing point of an aqueous solution containing 5 g of glucose (mol. mass = 180 g mol⁻¹) per 100 g of solution.

Ans. Molarity = 0.1539 m, $\Delta T_f, 273.15 - 271 = 2.15$ K

$$\therefore \Delta T_f = K_f \cdot m \Rightarrow m = \frac{\Delta T_f}{K_f} = \frac{2.15}{0.1539}$$

Also mass of solute, m = 5 g, molar mass of solute,

$M_{\text{solute}} = 180 \text{ g mol}^{-1}$ mass of solution, W = 100 g

$$\text{Using } \Delta T_f = \frac{K_f \times W_B \times 1000}{M_{\text{solute}} \times W_A} = \frac{2.15 \times 5 \times 1000}{0.1539 \times 180 \times 100} = \frac{10.75}{2.7702} = 3.88$$

\therefore Freezing point of solution = 273.15 - 3.88 = **269.27 K**.

27. What mass of NaCl (molar mass = 58.5 g mol⁻¹) be dissolved in 65 g of water to lower the freezing point by 7.5°C? The freezing point depression constant, K_f for water is 1.86 K kg mol⁻¹. Assume van't Hoff factor for NaCl is 1.87.

Ans. Here $M_2 = 58.5 \text{ g mol}^{-1}$, $w_1 = 65 \text{ g}$, $w_2 = ?$

$$\Delta T_f = 7.5^\circ\text{C}, i = 1.87, K_f = 1.86 \text{ K kg mol}^{-1}$$

Using the formula, $\Delta T_f = i K_f m$

$$\Delta T_f = i K_f \times \frac{w_2 \times 1000}{M_2 \times w_1} \Rightarrow 7.5 = 1.87 \times 1.86 \times \frac{w_2 \times 1000}{58.5 \times 65}$$

$$\therefore w_2 = \frac{7.5 \times 58.5 \times 65}{1.87 \times 1.86 \times 1000} = \frac{28518.75}{3478.2} = \mathbf{8.199 \text{ g}}$$

28. Assuming complete dissociation, calculate the expected freezing point of a solution prepared by dissolving 6.00 g of Glauber's salt, $\text{Na}_2\text{SO}_4 \cdot 10 \text{ H}_2\text{O}$ in 0.100 kg of H_2O .

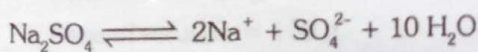
[Given for water, $K_f = 1.86 \text{ K kg mol}^{-1}$] Atomic masses : Na = 23, S = 32, O = 16, H = 1 amu]

Ans. Molar mass of $\text{Na}_2\text{SO}_4 \cdot 10 \text{ H}_2\text{O}$

$$= [2 \times 23 + 32 + 16 \times 4 + 10 \times (2 \times 1 + 16)] \text{ g/mole}$$

$$= (46 + 32 + 64 + 180) \text{ g mol}^{-1} = 322 \text{ g mole}^{-1}$$

$\text{Na}_2\text{SO}_4 \cdot 10 \text{ H}_2\text{O}$ ionises as :



$$\Rightarrow i = 3$$

$$m = \frac{n_B}{W_A} = \frac{W_B}{M_B \times W_A} = \frac{6.00 \text{ g}}{322 \text{ g mol}^{-1} \times 1 \text{ kg}}$$

$$= 0.186 \text{ mol kg}^{-1} = 0.186 \text{ m}$$

$$= 1.05 \text{ K}$$

$$\Rightarrow T_f = (273 - 1.05) \text{ K} = \mathbf{271.95 \text{ K}}$$

29. (a) What is meant by :

(i) Colligative properties (ii) Molality of a solution

(b) What concentration of nitrogen should be present in a glass of water at room temperature? Assume a temperature of 25°C, a total pressure of 1 atmosphere and mole fraction of nitrogen in air of 0.78 [K_H for nitrogen = 8.42×10^{-7} M/mm Hg]

Ans. (a) (i) **Colligative Properties.** The properties which depend only on the number of moles of solute (and not on their molecular weights or sizes) are referred to as colligative properties e.g. Lowering of vapor pressure, depression of freezing point, elevation of boiling point, osmotic pressure etc.

(ii) **Molality** : It is the number of moles of solute dissolved per kg (or per 1000 g) of the solvent. It is denoted by m .

$$\therefore \text{Molality} = \frac{\text{moles of solute}}{\text{weight of solvent in kg}} = \frac{w_2 \times 1000}{M \times w_1}$$

where w_1 = mass of solvent in grams

w_2 = mass of solute and

M_2 = Molar mass of solute

(b) Given $P_{N_2} = 1 \text{ atm}$

Applying Henry's law

$$P_{N_2} = K_H \cdot X_{N_2}$$

$$X_{N_2} = \frac{P_{N_2}}{K_H} = \frac{1}{8.42 \times 10^{-7}}$$

$$\text{Since } X_{N_2} = \frac{n(N_2)}{n(N_2) + n(H_2O)} = \frac{n(N_2)}{n(H_2O)}$$

$$n(N_2) = X_{N_2} \cdot n(H_2O) \quad [\because n(H_2O) = \frac{1000}{18} = 55.5 \text{ moles}]$$

$$= \frac{1}{8.42 \times 10^{-7}} \times 55.5 = 6.59 \times 10^7 \text{ moles.}$$

Hence concentration of nitrogen in water = $6.59 \times 10^7 \text{ mol L}^{-1}$.

30. 100 g of liquid A (molar mass 140 g mol⁻¹) was dissolved in 1000 g of liquid B (Molar mass 180 g mol⁻¹). The vapour pressure of pure liquid B was found to be 500 torr. Calculate the vapour pressure of pure liquid A and its vapour pressure in the solution if the total vapour pressure of the solution is 475 torr.

Ans. $n_A = \frac{100 \text{ g}}{140 \text{ g mol}^{-1}} = \frac{5}{7} \text{ mol}$ and $n_B = \frac{1000 \text{ g}}{180 \text{ g mol}^{-1}} = \frac{50}{9} \text{ mol}$

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

$$\therefore x_A = \frac{5/7}{5/7 + 50/9} = \frac{5/7}{395/63}$$

$$= \frac{5}{7} \times \frac{63}{395} = \frac{45}{395} = 0.114$$

$$\therefore x_B = 1 - 0.114 = 0.886$$

Also, given $P_B^0 = 500 \text{ torr}$

According to Raoult's law,

$$P_A = x_A P_A^0 = 0.114 \times P_A^0 \quad \dots(i)$$

$$P_B = x_B P_B^0 = 0.886 \times 500 = 443 \text{ torr}$$

$$P = P_A + P_B$$

$$475 = 0.114 P_A^0 + 443$$

or
$$P_A^0 = \frac{475 - 443}{0.114} = 280.7 \text{ torr}$$

Substituting this value in equation (i), we get

$$P_A = 280.7 \times 0.114 = 32 \text{ torr}$$

31. A 5% solution (by mass) of cane sugar in water has a freezing point of 271 K. Calculate the freezing point of a 5% glucose in water if the freezing point of pure water is 273.15 K.

Ans. Molality of sugar solution = $\frac{n_{C_{12}H_{22}O_{11}}}{W_{H_2O} \text{ (in grams)}} \times 1000$

$$= \frac{5}{342} \times \frac{1000}{100} = 0.146 \text{ m}$$

$$\Delta T_f \text{ for sugar solution} = 273.15 \text{ K} - 271 \text{ K} = 2.15 \text{ K}$$

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

$$\therefore K_f = \frac{\Delta T_f}{m} = \frac{2.15}{0.146} \text{ K kg mol}^{-1}$$

Molality of glucose solution = $\frac{n_{C_6H_{12}O_6}}{W_{H_2O} \text{ (in grams)}} \times 1000$

$$= \frac{5}{180} \times \frac{1000}{100} = 0.278 \text{ mol kg}^{-1}$$

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

$$\therefore \Delta T_f \text{ (Glucose)} = \frac{2.15}{0.146} \text{ K kg mol}^{-1} \times 0.278 \text{ mol kg}^{-1}$$

$$= 4.09 \text{ K}$$

$$\therefore \text{Freezing point of glucose solution} = 273.15 \text{ K} - 4.09 \text{ K} = 269.06 \text{ K}$$

32. Calculate (a) molality (b) molarity and (c) mole fraction of KI if the density of 20% (mass/mass) aqueous KI is 1.202 g mL⁻¹.

Ans. 20% (mass/mass) aqueous KI solution means that

$$\text{Mass of KI} = 20 \text{ g}$$

$$\text{Mass of solution} = 100 \text{ g}$$

$$\therefore \text{Mass of solvent (water)} = 100 - 20 = 80 \text{ g}$$

(a) Calculation of molality

$$\text{Molar mass of KI} = 39 + 127 = 166 \text{ g mol}^{-1}$$

$$\therefore n_{KI} = \frac{20 \text{ g}}{166 \text{ g mol}^{-1}} = 0.120$$

$$\text{Molality of solution} = \frac{n_{KI}}{W_{H_2O}} \times 1000 = \frac{0.120}{80} \times 1000 = 1.5 \text{ m}$$

(b) Calculation of molarity

$$\text{Density of solution} = 1.202 \text{ g mL}^{-1}$$

$$\therefore \text{Volume of solution} = \frac{100 \text{ g}}{1.202 \text{ g mL}^{-1}}$$

$$= 83.2 \text{ mL} = 0.0832 \text{ L}$$

$$\text{Molarity of solution} = \frac{n_{\text{KI}}}{\text{Volume of solution in L}}$$

$$= \frac{0.120 \text{ mol}}{0.0832 \text{ L}} = \mathbf{1.44 \text{ M}}$$

Alternatively,

$$\text{Molarity} = \frac{\% \times d \times 10}{\text{Molar mass of KI}}$$

$$= \frac{20 \times 1.202 \times 10}{166} = \frac{240.4}{166} = \mathbf{1.45 \text{ M}}$$

(c) Calculation of mole fraction of KI

$$n_{\text{KI}} = 0.120 \text{ mol}$$

$$n_{\text{H}_2\text{O}} = \frac{\text{Mass of water}}{\text{Molar mass of water}} = \frac{80 \text{ g}}{18 \text{ g mol}^{-1}} = 4.44 \text{ mol}$$

$$x_{\text{KI}} = \frac{n_{\text{KI}}}{n_{\text{H}_2\text{O}} + n_{\text{KI}}} = \frac{0.120 \text{ mol}}{4.44 \text{ mol} + 0.12 \text{ mol}}$$

$$= \frac{0.120 \text{ mol}}{4.560 \text{ mol}} = \mathbf{0.0263}$$

- 33.** Benzene and toluene form ideal solution over the entire range of composition. The vapour pressures of pure benzene and toluene at 300 K are 50.71 mm Hg and 32.06 mm Hg respectively. Calculate the mole fraction of benzene in the vapour phase, if 80 g of benzene is mixed with 100 g of toluene.

Ans. Molar mass of benzene (C_6H_6) = 78 g mol⁻¹

Molar mass of toluene (C_7H_8) = 92 g mol⁻¹

$$n_{\text{C}_6\text{H}_6} \text{ (no. of moles of benzene)} = \frac{80 \text{ g}}{78 \text{ g mol}^{-1}} = 1.026 \text{ mol}$$

$$n_{\text{C}_7\text{H}_8} \text{ (no. of moles of Toluene)} = \frac{100 \text{ g}}{92 \text{ g mol}^{-1}} = 1.087 \text{ mol}$$

$$\text{Mole fraction of benzene } (x_{\text{C}_6\text{H}_6}) = \frac{1.026 \text{ mol}}{(1.026 + 1.087) \text{ mol}} = 0.486$$

$$\text{Mole fraction of toluenes } (x_{\text{C}_7\text{H}_8}) = 1 - 0.486 = 0.514$$

Applying Raoult's law,

Partial vapour pressure of benzene in solution ($p_{\text{C}_6\text{H}_6}$)

$$= p_{\text{C}_6\text{H}_6}^0 \times x_{\text{C}_6\text{H}_6}$$

$$= (50.71 \text{ mm}) \times (0.486)$$

$$= 24.65 \text{ mm}$$

Partial vapour pressure of toluene in solution

$$p_{\text{C}_7\text{H}_8} = p_{\text{C}_7\text{H}_8}^0 \times x_{\text{C}_7\text{H}_8}$$

$$= 32.06 \text{ mm} \times 0.514$$

$$= 16.48 \text{ mm}$$

Total vapour pressure of solution

$$(p) = (24.65 + 16.48) \text{ mm}$$

$$= 41.13 \text{ mm Hg}$$

$$\begin{aligned} \text{Mole fraction of benzene in vapour phase} &= \frac{x_{C_6H_6} \times P_{C_6H_6}^{\circ}}{P_{\text{Total}}} \\ &= \frac{0.486 \times (50.71 \text{ mm})}{41.13 \text{ mm}} = \mathbf{0.6} \end{aligned}$$

- 34.** If the partial pressure of oxygen in air is 0.2 atm, calculate the concentration (in moles/litre) of dissolved oxygen in water in equilibrium with air at 25°C. (The Henry's law constant for oxygen dissolved in water is 4.34×10^4 atm at 25°C).

Ans. According to Henry's law $p = K_H \cdot x$ or $p_{O_2} = K_H x_{O_2}$

Here, $p_{O_2} = 0.2 \text{ atm}$

$$K_H = 4.34 \times 10^4 \text{ atm}$$

Then,
$$x_{O_2} = \frac{p_{O_2}}{K_H} = \frac{0.2}{4.34 \times 10^4} = 4.6 \times 10^{-6}$$

On changing mole fraction in molarity

$$\text{Moles of water } (n_{H_2O}) = \frac{1000}{18} = 55.5 \text{ mol}$$

Since, n_{O_2} is very small in comparison to n_{H_2O}

so,
$$n_{O_2} + n_{H_2O} \approx n_{H_2O}$$

$$x_{O_2} = \frac{n_{O_2}}{n_{H_2O}}$$

or
$$\begin{aligned} n_{O_2} &= x_{O_2} \times n_{H_2O} \\ &= 4.6 \times 10^{-6} \times 55.5 \\ &= 2.55 \times 10^{-4} \text{ mol} \end{aligned}$$

Since, 2.55×10^{-4} moles are present in 1000 mL of water, so molarity = $\mathbf{2.55 \times 10^{-4} \text{ mol/L or M}}$.

- 35.** The air is a mixture of a number of gases. The major components are oxygen and nitrogen with approximate proportion of 20% and 79% by volume at 298 K respectively. The water is in equilibrium with air at a pressure of 10 atm. At 298 K, if the Henry's law constant for oxygen and nitrogen are 3.30×10^7 mm and 6.51×10^7 mm respectively, calculate the composition of these gases in water.

Ans. Given that total pressure of air in equilibrium with water = 10 atm

As air contains 20% oxygen and 79% nitrogen by volume

$$\begin{aligned} \therefore \text{Partial pressure of oxygen } (p_{O_2}) &= \frac{20}{100} \times 10 \text{ atm} \\ &= 2 \text{ atm} = 2 \times 760 \text{ mm} \\ &= 1520 \text{ mm} \end{aligned}$$

$$\text{Partial pressure of nitrogen, } (p_{N_2}) = \frac{79}{100} \times 10 \text{ atm}$$

$$= 7.9 \text{ atm} = 7.9 \times 760 \text{ mm} = 6004 \text{ mm}$$

Given that, $K_H(O_2) = 3.30 \times 10^7 \text{ mm}$,
 $K_H(N_2) = 6.51 \times 10^7 \text{ mm}$,

According to Henry's law,

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

or
$$x_{O_2} = \frac{p_{O_2}}{K_H} = \frac{1520 \text{ mm}}{3.30 \times 10^7 \text{ mm}} = \mathbf{4.61 \times 10^{-5}}$$

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

or
$$x_{N_2} = \frac{p_{N_2}}{K_H} = \frac{6004 \text{ mm}}{6.51 \times 10^7 \text{ mm}} = \mathbf{9.22 \times 10^{-5}}$$

EXERCISE-1

PREVIOUS YEARS BOARD PROBLEMS

CBSE 2016

- Write the colligative property which is used to find the molecular mass of macromolecules.
 - In non-ideal solution, what type of deviation shows the formation of minimum boiling azeotropes?
- Calculate the boiling point of solution when 2 g of Na_2SO_4 ($M = 142 \text{ g mol}^{-1}$) was dissolved in 50 g of water, assuming Na_2SO_4 undergoes complete ionization. (K_b for water = $0.52 \text{ K kg mol}^{-1}$)

CBSE 2015

- Why are aquatic species more comfortable in cold water than in warm water?
 - What happens when we place the blood cell in saline water solution (hypertonic solution)? Give reason.
- Vapour pressure of water at 20°C is 17.5 mm Hg. Calculate the vapour pressure of water at 20°C when 15 g of glucose (Molar mass = 180 g mol^{-1}) is dissolved in 150 g of water.

CBSE 2014

- Define the following terms :
 - Molarity
 - Molal elevation constant (K_b)
 - A solution containing 15 g urea (molar mass = 60 g mol^{-1}) per litre of solution in water has the same osmotic pressure (isotonic) as solution of glucose (molar mass = 180 g mol^{-1}) in water. Calculate the mass of glucose present in one litre of its solution. **[45 g]**
- What type of deviation is shown by a mixture of ethanol and acetone? Give reason.
 - A solution of glucose (molar mass = 180 g mol^{-1}) in water is labelled as 10% (by mass). What would be the molality and molarity of the solution? (Density of solution = 1.2 g mL^{-1}) **[Molality - 0.61 mol kg^{-1}]**
[Molarity of the solution - 0.66 mol L^{-1}]

CBSE 2013

- State Raoult's law for a solution containing volatile components. How does Raoult's law become a special case of Henry's law?
 - 1.00 g of a non-electrolyte solute when dissolved in 50 g of benzene lowered the freezing point of benzene by 0.40 K. Find the molar mass of the solute. (K_f for benzene = $5.12 \text{ K kg mol}^{-1}$) **[256 g mol⁻¹]**
- Define the following terms
 - Ideal solution
 - Azeotrope
 - Osmotic pressure
 - A solution of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) in water is labelled as 10% by weight. What would be the molality of the solution? (Molar mass of glucose = 180 g mol^{-1}) **[0.67 M]**

CBSE 2012

- Define mole fraction.
- Explain the Henry's law about dissolution of a gas in a liquid.
- Define ideal solution.
- Explain boiling point elevation constant for a solvent or Ebullioscopic constant.

5. Define the following terms :

- (i) Mole fraction (ii) Isotonic solution (iii) Van't Hoff factor (iv) Ideal solution.

6. 15.0 g of unknown molecular material is dissolved in 450 g of water. The resulting solution freezes at -0.34°C . What is the molar mass of the material? (K_f for water = 1.86 K mol^{-1}) **[182.35 g mol⁻¹]**

7. A solution of glycerol ($\text{C}_3\text{H}_8\text{O}_3$) in water was prepared by dissolving some glycerol in 500 g of water. This solution has a boiling point of 100.42°C , what mass of glycerol was dissolved to make this solution? (K_b for water = $0.512 \text{ K kg mol}^{-1}$) **[37.73 g]**

8. (i) Explain the following :

- (a) Henry's law about dissolution of a gas in a liquid.
(b) Boiling point elevation constant for a solvent.

(ii) A solution of glycerol ($\text{C}_3\text{H}_8\text{O}_3$) in water was prepared by dissolving some glycerol in 500 g of water. This solution has a boiling point of 100.42°C . What mass of glycerol was dissolved to make this solution? (K_b for water = 0.512 K mol^{-1}). **[37.73 g]**

CBSE 2011

1. Define Raoult's law in its general form in reference to solutions.

2. What is meant by reverse osmosis?

3. The molecular masses of polymers are determined by osmotic pressure method and not by measuring other colligative properties. Give two reasons.

4. A solution prepared by dissolving 8.95 mg of a gene fragment in 35.0 mL of water has an osmotic pressure of 0.335 torr at 25°C . Assuming the gene fragment is non-electrolyte, determine its molar mass. **[1.4193 × 10⁴ g mol⁻¹]**

5. What mass of NaCl must be dissolved in 65.0 g of water to lower the freezing point of water by 7.50°C ? The freezing point depression constant (K_f) for water is 1.86°C/m . Assume van't Hoff factor for NaCl is 1.87. (Molar mass of NaCl = 58.5 g mol^{-1}). **[8.199 g]**

6. At 300 K, 36 g of glucose, $\text{C}_6\text{H}_{12}\text{O}_6$ present per litre in its solution has an osmotic pressure of 4.98 bar. If the osmotic pressure of another glucose solution is 1.52 bar at the same temperature, calculate the concentration of the other solution. **[0.0610 mol L⁻¹]**

7. Calculate the boiling point of one molar aqueous solution. Density of KBr solution is 1.06 g mL^{-1} . (K_b for $\text{H}_2\text{O} = 0.52 \text{ K kg mol}^{-1}$, atomic mass; K = 39, Br = 80.) **[T_b = 373.553 K]**

8. (i) The molecular masses of polymers are determined by osmotic pressure method and not by measuring other colligative properties. Give two reasons.

(ii) At 300 K, 36 g of glucose, $\text{C}_6\text{H}_{12}\text{O}_6$ present per litre in its solution has an osmotic pressure of 4.98 bar. If the osmotic pressure of another glucose solution is 1.52 bar at the same temperature, calculate the concentration of the other solution. **[C = 0.0610 mol L⁻¹]**

9. (i) List any four factors on which the colligative properties of a solution depend.

(ii) Calculate the boiling point of one molar aqueous solution (density 1.06 g mL^{-1}) of KBr. [Give K_b for $\text{H}_2\text{O} = 0.52 \text{ K kg mol}^{-1}$ Atomic mass : K = 39, Br = 80]. **[373.553 K]**

CBSE 2010

1. State Henry's law and mention its two important applications.
2. Define the term 'molarity of a solution'. State one disadvantage in using the molarity as the unit of concentration.
3. Define the term osmosis and osmotic pressure. What is the advantage of using osmotic pressure as compared to other colligative properties for the determination of molar masses of solutes in solutions?
4. Find the boiling point of a solution containing 0.520 g glucose ($C_6H_{12}O_6$) dissolved in 80.2 g of water. [Given, K_b for water = 0.52 K m^{-1}]. **[373.019 k]**
5. Find the freezing point of a solution containing 0.520 g glucose ($C_6H_{12}O_6$) dissolved in 80.2 g of water [Given, K_f for water = 1.86 K m^{-1}]. **[272.933 k]**
6. What mass of ethylene glycol (molar mass 62.0 g mol^{-1}) must be added to 5.50 kg of water to lower the freezing point of water from 0°C to -10.0°C ? [K_f for water = $1.86 \text{ K kg mol}^{-1}$]. **[1833.3 g]**
7. Calculate the mass of ascorbic acid ($C_6H_8O_6$) to be dissolved in 75 g of acetic acid to lower its melting point by 1.5°C . [K_f for acetic acid is $3.9 \text{ K kg mol}^{-1}$]. **[5.08 g]**

CBSE 2009

1. What is meant by molality of the solution?
2. What is meant by positive and negative deviations from Raoult's law and how is the sign of $\Delta_{\text{mix}} H$ related to positive and negative deviations from Raoult's law?
3. Define the term van't Hoff factor.
4. What is meant by colligative properties?
5. 100 mg of a protein is dissolved in just enough water to make 10.0 mL of solution. If this solution has an osmotic pressure of mass of the protein? ($R=0.0821 \text{ L atm mol}^{-1} \text{ K}^{-1}$ and $760 \text{ mm Hg} = 1 \text{ atm}$). **[13980.45 g mol⁻¹]**
6. Calculate the freezing point of a solution containing 18 g glucose, $C_6H_{12}O_6$ and 68.4 g sucrose, $C_{12}H_{22}O_{11}$ in 200 g of water. The freezing point of pure water is 273 K and K_f for water is 1.86 K m^{-1} . **[270.21 k]**

EXERCISE-1

SOLUTIONS OF PREVIOUS YEARS BOARD PROBLEMS

SOLUTION

CBSE 2016

- Sol.1** (i) Osmotic pressure
(ii) Positive deviation from Raoult's law/ Positive deviation

$$\text{Sol.2 } \Delta T_b = i \frac{K_b w_b \times 1000}{M_b \times w_a}$$

$$\Delta T_b = \frac{3 \times 0.52 \times 2 \times 1000}{142 \times 50}$$

$$= 0.439 \text{ K}$$

$$\Delta T_b = T_b - T_b^0$$

$$T_b = 0.439 + 373 = 373.439 \text{ K}$$

CBSE 2015

- Sol.1** (i) As solubility of gases decreases with increase of temperature, less oxygen is available in summer in the lakes / as cold water contains more oxygen dissolved.
(ii) They will shrink, due to osmosis.

Sol.2 $P_A^0 = 17.5 \text{ mm of Hg}$ $W_B = 15 \text{ g}$ $M_B = 180 \text{ g/mol}$
 $W_A = 150 \text{ g}$ $P_S = ?$

$$\frac{P_A^0 - P_S}{P_A^0} = \frac{W_B \times M_A}{M_B \times W_A} \therefore \frac{P_A^0 - P_S}{P_A^0} = \frac{15 \times 18}{180 \times 150} = 0.01$$

$$\frac{P_A^0 - P_S}{P_A^0} = \frac{17.5 - P_S}{17.5} = 0.01$$

$$\therefore P_S = 17.325 \text{ mm of Hg}$$

CBSE 2014

- Sol.1** (a) (i) Molarity

Molarity (M) is defined as the number of moles of the solute dissolved in one Litre of solution.

It is expressed as :

$$\text{Molarity (M)} = \frac{\text{Moles of solute}}{\text{Volume of solvent in Litre}}$$

- (ii) When molality (m) of solution is taken as unity then $K_b = T_b$
(b) For isotonic solution

Hence

$$\pi_1 = \pi_2$$

$$C_1 = C_2$$

$$\frac{W_1}{M_1} = \frac{W_2}{M_2}$$

$$\frac{15}{60} = \frac{W_2}{180}$$

$$W_2 = 45 \text{ gm}$$

Hence glucose is 45% by weight present in aqueous solution.

Sol.2 (a) In case of negative deviations from Raoult's law, the intermolecular attractive forces between A-A and B-B are weaker than those between A-B and leads to decrease in vapour pressure resulting in negative deviations. A mixture of ethanol and acetone forms a solution with negative deviation from Raoult's law. This is because ethanol molecule is able to form hydrogen bond with acetone molecule as shown.

This decreases the escaping tendency of molecules for each component and consequently the vapour pressure decreases resulting in negative deviation from Raoult's law.

(b) 10% w/w solution of glucose in water means that 10 g of glucose is present in 100 g of the solution i.e., 10 g of glucose is present in $(100 - 10) \text{ g} = 90 \text{ g}$ of water.

Molar mass of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) = $6 \times 12 + 12 \times 1 + 6 \times 16 = 180 \text{ g mol}^{-1}$

Then, number of moles of glucose = $\frac{10}{180} \text{ mol} = 0.056 \text{ mol}$

\therefore **Molality** of solution = $\frac{0.056}{0.09 \text{ kg}} = 0.62 \text{ m}$

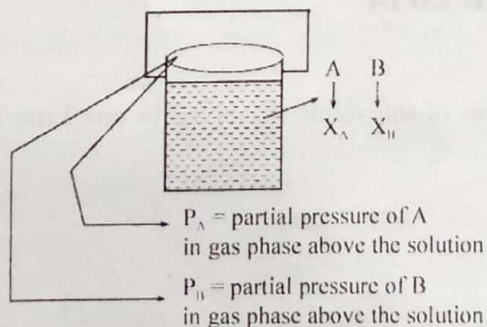
If the density of the solution is 1.2 g mL^{-1} , then the volume of the 100 g solution can be given as :

$$V = \frac{100 \text{ g}}{1.2 \text{ g mL}^{-1}} = 83.33 \text{ mL} = 83.33 \times 10^{-3} \text{ L}$$

\therefore **Molarity** of the solution = $\frac{0.056 \text{ mol}}{83.33 \times 10^{-3} \text{ L}} = 0.67 \text{ M}$

CBSE 2013

Sol.1 (a) Raoult's law states that for a binary solution containing two volatile components.



P_T = Total pressure of A & B in gas phase above the solution then

$$P_T = P_A + P_B$$

According Raoult's law

$$P_A \propto X_A$$

$$\Rightarrow P_A = P_A^\circ X_A \quad \dots(1)$$

Similarly $P_B = P_B^\circ X_B$
 So $P_T = P_A^\circ X_A + P_B^\circ X_B$

(b) Henry's law is a law dealing with solubility of gases in liquid, it states that for a solution of a gas in a liquid the mol fraction of gas in liquid phase is directly proportional to the partial pressure of gas above the liquid.

$$X \propto P$$

$$\Rightarrow X = K_H P \quad \dots(2)$$

Now compare equation 2 and 1 assuming $[P_B^\circ = 0]$

it means $P_A^\circ = \frac{1}{K_H}$ if solvent is non volatile. $\Delta T_f = K_f m$

$$0.40 = 5.12 \times \left[\frac{1}{M \times 50} \times 1000 \right]$$

$$M = \frac{5.12 \times 1 \times 1000}{50 \times 0.40} = 256 \text{ gm}$$

Sol.2 (a) Ideal solution - A solution which obeys Raoult's law under all conditions of temperature and pressure

(b) Azeotrope - A type of liquid mixture having definite composition and boiling like a pure liquid, is called a constant boiling mixture or an azeotropic mixture.

$$\text{So } P_T = P_A^\circ X_A + P_B^\circ X_B$$

it means $P_A^\circ = \frac{1}{K_H}$ if solvent is non volatile.

(C) Osmotic pressure - The minimum excess pressure that has to be applied on the solution to prevent the entry of the solvent into the solution through the semi permeable membrane is called the osmotic pressure.

(b) 10% by weight means

100 g solution has 10 gm glucose

$$\Rightarrow 90 \text{ g solvent has } \frac{10}{180} \text{ mol glucose}$$

$$1000 \text{ g solvent has } \frac{10}{180 \times 90} \times 1000 \text{ mol of solute}$$

$$\text{So molality} = 0.61728 \text{ m}$$

CBSE 2012

Sol.1 (a) (i) Mole fraction - ratio of moles of solute and sum of moles of solute and moles of solution.

Sol.2 The partial pressure of the gas in vapour phase (P) is proportional to the mole fraction of the gas (x) in the solution.

Sol.3 (ii) Ideal solution - A solution which follows Raoult's law for a large range of concentration.

Sol.4 As the elevation in boiling point produced when 1 mole of solute is dissolved in one kg. Of the solvent.

Sol.5 (i) Mole fraction \Rightarrow ratio of moles of solute and sum of moles of solute and moles of solution.

(ii) **Isotonic solution** \Rightarrow Two solutions having same osmotic pressure at a given temp. are called isotonic solutions.

- (iii) **Van't Hoff factor** \Rightarrow The van't Hoff factor is the ratio between the actual conc. Of particles produced when the substance is dissolved and the conc. Of a substance as calculated from its mass.
- (iv) **Ideal solution** \Rightarrow A solution which follows Raoult's law for a large range of concentration.

Sol.6 $\Delta T_f = K_f \times \frac{W_B \times 1000}{M_B \times W_A}$

$$0.34 = \frac{1.86 \times 15 \times 1000}{M_B \times 450}$$

$$M_B = \frac{1.86 \times 15 \times 1000}{0.34 \times 450} = 164.7 \text{ g/mol}$$

Sol.7 $\Delta T_b = K_b \times M$

$$100.42 - 100 = 0.512 \times \frac{w \times 1000}{92 \times 500}$$

$$w = \frac{92 \times 500 \times 0.42}{0.512 \times 1000} = 37.7 \text{ g.}$$

Sol.8 (i) (a) same as **Q. 2 in 2012**

(b) Same as **Q. 4 in 2012**

(ii) Same as **Q. 7 in 2012**

CBSE 2011

Sol.1 Raoult's law for solution containing volatile liquids : 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.

Sol.2 The direction of a osmosis can be reversed if a pressure larger than the osmotic pressure is applied to the solution side that is, now the pure solvent flows out of the solution through the semi permeable membrane. this phenomenon is called reverse osmosis.

Sol.3 (i) Osmotic pressure can be measured at room temp. whereas other colligative properties need either high temp or low temp conditions.

(ii) Value of the other colligative properties is very small that a small error can account to large difference.

Sol.4 $w = 8.95 \text{ mg} = 8.95 \times 10^{-4} = 0.0895 \text{ g.}$

$$\pi = 0.335, T = 25^\circ\text{C} = 25 + 273 = 298 \text{ K.}$$

$$R = 35.0 \text{ ml.}$$

$$\text{Molar Mass } M = \frac{wRT}{\pi V} = \frac{8.95 \times 35.0 \times 298}{0.335 \times 0.0895} = 311.34 \times 10^4 \text{ g mol}^{-1}.$$

Sol.5 Here $M_2 = 58.5 \text{ g mol}^{-1}$, $w_1 = 65 \text{ g}$, $w_2 = ?$

$$\Delta T_f = 7.5^\circ\text{C}, i = 1.87, K_f = 1.86 \text{ K kg mol}^{-1}$$

Using the formula, $\Delta T_f = i K_f m$

$$\Delta T_f = i K_f \times \frac{w_2 \times 1000}{M_2 \times w_1} \Rightarrow 7.5 = 1.87 \times 1.86 \times \frac{w_2 \times 1000}{58.5 \times 65}$$

$$\therefore w_2 = \frac{7.5 \times 58.5 \times 65}{1.87 \times 1.86 \times 1000} = \frac{28518.75}{3478.2} = 8.199 \text{ g}$$

7. Allen Career Institute, Kota (Rajasthan) | CBSE Solutions | Chemistry | Class 12 | Chapter 1 | Solutions | Sol. 5

Sol.6 $\pi_1 = CRT$ $\pi_2 = CRT$

$$4.98 = \frac{36 RT}{180 \times v} \quad \dots(1)$$

$$1.52 = \frac{W_B \times RT}{M_B \times V} \quad \dots(2)$$

Divide eq. (2)/(1)

$$\frac{W_B}{M_B} = \frac{1.52}{4.98} \times \frac{1}{5} = 0.0610 \text{ mol L}^{-1}$$

Sol.7 $\Delta T_b = i \times K_b \times m$
 $= 2 \times 0.52 \times 1.0627$
 $= 1.1052$
 b.P of kBr solution
 $= 373 + 1.1052$
 $= 374.10 \text{ K}$

$$\left\{ \begin{aligned} d &= \frac{M}{V} \\ 1.06 &= \frac{M}{1000} \Rightarrow M = 1.06 \times 1000 \text{ g} \\ &= 1060 \text{ g} \\ m &= \frac{\text{Moles}}{\text{Mass of solvent (kg)}} = \frac{1}{\frac{1000}{941}} \\ &= \frac{941}{1000} = 1.0627 \end{aligned} \right.$$

KBr = 39 + 80 = 119
 1060 - 119 = 941 g.

Sol.8 $\pi_1 = CRT$ $\pi_2 = CRT$

$$4.98 = \frac{36 RT}{180 \times v} \quad \dots(1)$$

$$1.52 = \frac{W_B \times RT}{M_B \times V} \quad \dots(2)$$

Divide eq (2)/(1)

$$\frac{W_B}{M_B} = \frac{1.52}{4.98} \times \frac{1}{5} = 0.0610 \text{ mol L}^{-1}$$

- Sol.9 (i) (a) Relative lowering of vapour pressure
 (b) Elevation in boiling point
 (c) Depression in freezing point
 (c) Osmotic pressure

(ii)
 $\Delta T_b = i K_b \times m$
 $= 2 \times 0.52 \times 1.0627$
 $= 1.1052$
 b.P of kBr Solution
 $= 373 + 1.1052$
 $= 374.10 \text{ K}$

$$\left\{ \begin{aligned} d &= \frac{M}{V} \\ 1.06 &= \frac{M}{1000} \Rightarrow M = 1.06 \times 1000 \text{ g} \\ &= 1060 \text{ g} \\ m &= \frac{\text{Moles}}{\text{Mass of Solvent (g)}} = \frac{1}{\frac{1000}{941}} \\ &= \frac{941}{1000} = 1.0627 \end{aligned} \right.$$

kBr = 39 + 80 = 119
 1060 - 119 = 941 g.

CBSE 2010

- Sol.1 (i) Bottles are sealed under high pressure to increase the solubility of CO₂ in soft drinks and soda water.
 (ii) The conc. of oxygen is low in the blood and tissues of people living at high altitudes such as climbers. This is because at high altitudes, partial pressure of oxygen is less than that at ground level. Low-blood oxygen causes climbers to become weak and disables them from thinking clearly. These are symptoms of anoxia.

Sol.2 Molarity is defined as the number of moles of the solute dissolved in one litre of solution.

It is temp depended

Sol.3 The phenomenon of the flow of solvent molecules through a semipermeable membrane molecules from pure solvent to the solution is known as osmosis.

→ The Osmotic pressure of a solution is the excess pressure that must be applied to a solution to prevent osmosis.

→ It's temp remains constant that is room temp

$$\text{Sol.4 } T_b - 373 = 1 \times .52 \times \frac{.52}{180} \times \frac{1000}{80.2}, T_b = 0.0184 \times 373$$

$$T_b = 373.01 \text{ K}$$

$$\text{Sol.5 } 273 - T_f = 1 \times 1.86 \times \frac{.52 \times 1000}{180 \times 80.2} - T_f = 0.669 - 273$$

$$- T_f = -272.3$$

$$T_f = 272.3 \text{ K}$$

$$\text{Sol.6 } 10 = 1.86 \times \frac{w}{62 \times 5.5} = \frac{62 \times 5.5 \times 10}{1.86} = 1833.3 \text{ g}$$

$$\text{Sol.7 } 1.5 = 3.9 \frac{w \times 1000}{176 \times 75} = \frac{176 \times 75 \times 1.5}{3.9 \times 1000} = 5.08 \text{ gm}$$

CBSE 2009

Sol.1 It is the number of moles of solute dissolved per kg (or per 1000 g) of the solvent. It is denoted by m.

$$\therefore \text{ Molality} = \frac{\text{moles of solute}}{\text{weight of solvent in kg}}$$

$$= \frac{w_2 \times 1000}{M \times w_1}$$

where w_1 = mass of solvent in grams

w_2 = mass of solute and

M_2 = Molar mass of solute

Sol.2

| | | | |
|------|--|------|---|
| (I) | The intermolecular interactions. i.e. A-B interactions are weaker than in pure components, i.e. between A-A or B-B | (I) | The intermolecular interactions, ie. A-B interactions are stronger than in pure components, i.e. between A-A or B-B |
| (II) | $\Delta_{\text{mix}} H = \text{positive}$ [endothermic] | (II) | $\Delta_{\text{mix}} H = \text{negative}$ [exothermic] |

Sol.3 (a) (i) Mole fraction : Mole fraction (X) of a component in a mixture is the ratio between the number of moles of that component to the total number of moles of all the component of the mixture.

\therefore Mole fraction of component A is given by

$$X_A = \frac{\text{moles of A}}{\text{total moles of all the components of the mixture}}$$

(ii) Van't Hoff factor : Van't Hoff factor i is a correction factor defined as

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

Sol.4 The properties which depend only on the number of moles of solute (and not on their molecular weights or sizes) are referred to as colligative properties.

e.g. Lowering of vapor pressure, depression of freezing point, elevation of boiling point, osmotic pressure etc.

Sol.5 Here $w = 100 \text{ mg} = 0.100 = 0.1 \text{ g}$

$$V = 10.0 \text{ mL} = 0.01 \text{ L}$$

$$p = 13.3 \text{ mm Hg} = \frac{13.3}{760} \text{ atm}$$

$$T = 25^\circ\text{C} = 25 + 273 = 298 \text{ K}$$

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

$$M = ?$$

$$\therefore \text{Molar mass } M = \frac{wRT}{\pi V}$$

$$M = \frac{0.1 \times 0.0821 \times 298}{\frac{13.3}{760} \times 0.01} = \frac{0.1 \times 0.0821 \times 298 \times 760}{13.3 \times 0.01} = \frac{1859.4008}{0.133}$$

$$M = 13980.4 \text{ g mol}^{-1}$$

$$\text{Sol.6 } 273 - T_f = 1.86 \times \frac{180 + \frac{68.4}{1000}}{200}$$

$$273 - T_f = 2.79$$

$$T_f = 273 - 2.73 = 270.21 \text{ K}$$

$$273 - T_f = 1.86 \times \frac{0.1 + .2}{.1}$$

$$273 - T_f = \frac{0.558}{.2}$$