Chapter-16 : Solutions

1. **(c)** We know that, density of water = 1000 g/L

$$\therefore \text{Moles of water} = \frac{\text{Mass}}{\text{Molar mass}} = \frac{1000}{18} = 55.6$$

$$\text{Molarity} = \frac{\text{Number of moles}}{\text{Volume of solution (L)}}$$

$$\therefore \text{Molarity} = \frac{55.6}{1} = 55.6\,\text{M}$$

2. **(b)** One molal solution means one mole of solute is present in 1 kg (1000 g) solvent i.e., mole of solute = 1

$$\text{Mole of solvent } (H_2O) = \frac{1000\text{g}}{18\text{g}} = \frac{1000}{18}$$

Mole fraction of solute

$$= \frac{1}{\left(1+\frac{1000}{18}\right)} = \frac{18}{1018} = 0.0176 \approx 0.018$$

3. **(d)** As the solubility increases, value of Henry's law constant decreases. Since CO_2 is most soluble in water among the given set of gases. Therefore CO_2 has the lowest value of Henry's law constant.

4. **(b)** $\text{No. of moles of NaCl} = \frac{M \times V}{1000} = \frac{0.25 \times 100}{1000} = 0.025$

$NaCl \rightarrow Na^+ + Cl^-$

No. of moles of Na^+ ions = 0.025

No. of Na^+ ions = $0.025 \times 6.023 \times 10^{23} = 1.505 \times 10^{22}$

5. **(b)** Molarity is intrinsic property. Hence, it is independent of amount of solution.

6. **(c)** $N_1V_1 + N_2V_2 = NV$

$4x + 10(1-x) = 6 \times 1; -6x = -4; x = 0.67$

Thus 0.67 litre of 4N HCl and

$1 - x = 1 - 0.67 = 0.33$ litre of 10 N HCl

7. **(d)** Vapour pressure of a solution containing non-volatile solute is less than that of the pure solvent. The decrease in vapour pressure depends upon the quantity of non-volatile solute present in it. Hence, vapour pressure of A > C > B.

8. **(b)** $\Delta T_f = K_f \times m = 1.86 \times 0.5 = 0.93\,°C;\ T_f = -0.93°C$

9. **(c)**

	$Na_2SO_4 \longrightarrow$	$2Na^+ +$	SO_4^{2-}
Initial	1	0	0
Final	$1-\alpha$	2α	1α

$i = 1 - \alpha + 2\alpha + \alpha = 1 + 2\alpha$

10. **(d)** $2NaHCO_3 + H_2C_2O_4 \longrightarrow Na_2C_2O_4 + 2CO_2 + 2H_2O$

$$\text{Moles of } CO_2 \text{ evolved} = \frac{0.25}{25 \times 10^3} = 10^{-5}$$

$\therefore$ moles of $NaHCO_3 = 10^{-5}$

$\therefore$ mass of $NaHCO_3 = 84 \times 10^{-5}$ g

$= 0.84 \times 10^{-3}$ g $= 0.84$ mg

$$\therefore \%\text{ by weight} = \frac{0.84}{10} \times 100$$

$= 8.4\%$

11. **(b)** We know that, $M_1V_1 = M_2V_2$

$0.05 \times V_1 = 0.4 \times 100$

$$\therefore V_1 = \frac{0.4 \times 100}{0.05} = 800$$

$V_2 - V_1 = 800 - 100 \Rightarrow 700$ mL

12. **(b)** No. of moles in 0.02 M H_2SO_4 solution

$[H_3O^+] = 2 \times 0.02 = 0.04$ moles

$\therefore$ 2 litre solution contains 0.08 mole of H_3O^+.

13. **(d)** $W = \frac{N \times E_W \times V(\text{mL})}{1000} = \frac{0.05 \times 49.04 \times 100}{1000} = 0.2452\text{g}$

14. **(c)** $\pi_{\text{glucose}} = \pi_{\text{unknown solute}}$

$$\therefore \frac{m_1}{M_1} = \frac{m_2}{M_2} \text{ or } \frac{5}{180} = \frac{2}{M_2} \Rightarrow M_2 = 72$$

15. **(c)** F^- ions in ppm $= \frac{0.2}{500} \times 10^6 = 400$

16. **(b)** Acetone solution has vapour pressure less than pure water.

17. **(d)** 1000 mL of 1 N oxalic solution = 63 g

500 mL of 0.2 N oxalic acid solution

$$= \frac{63}{1000} \times 500 \times 0.2 = 6.3\text{ g}$$

18. **(b)** $\Delta T_b = K_b \times m \times i = 0.52 \times 1 \times 2 = 1.04$

$\therefore \Delta T_b = 100 + 1.04 = 101.04\,°C$

19. **(d)** Osmotic pressure is a colligative property. Hence resulting osmotic pressure of the solution is given by

$\pi_T = \pi_1 + \pi_2 + \pi_3 \ldots\ldots\ldots$

$\pi_T = 1.64 + 2.46 = 4.10$ atm.

20. **(a)** As both the solutions are isotonic hence there is no net movement of the solvent occurs through the semipermeable membrane between two solutions.

21. **(a)** From the relation

$$\frac{\text{Normality}}{\text{Molarity}} = \frac{\text{Molecular mass}}{\text{Equivalent mass}} = n$$

For 2N HCl,

$$\text{Molarity} = \frac{\text{Normality} \times \text{Equivalent weight}}{\text{Molecular weight}} = \frac{2 \times 36.5}{36.5} = 2$$

For 4 N H_2SO_4,

$$\text{Molarity} = \frac{\text{Normality} \times \text{Equivalent weight}}{\text{Molecular weight}} = \frac{4 \times 49}{98} = 2$$

Hence 4 N H_2SO_4 and 2N HCl solution will have same molar concentration.

22. **(c)** Equivalent weight of orthophosphoric acid

$$(H_3PO_4) = \frac{3+31+64}{3} = \frac{98}{3}$$

Now 100 g solution contains 70 g H_3PO_4

$\frac{100}{1000 \times 1.54}$ litre of solution contains $\frac{70}{98/3}$ g equivalent of H_3PO_4

∴ Normality of solution,

$$= \frac{\frac{70 \times 3}{98}}{\frac{1}{10 \times 1.54}} = \frac{70 \times 3}{98} \times 10 \times 1.54 = 33 \text{ N}$$

23. **(c)** Mass of solute taken = 120 g

Molecular mass of solute = 60 u

Mass of solvent = 1000 g

Density of solution = 1.15 g mL^{-1}

Total mass of solution = 1000 + 120 = 1120 g

$$\text{Volume of solution} = \frac{\text{Mass}}{\text{Density}} = \frac{1120}{1.15} \text{mL}$$

$$\text{Molarity} = \frac{\frac{\text{Mass of solute}}{\text{Molecular mass of solute}}}{\text{Volume of solution}} \times 1000$$

$$= \frac{120/60}{1120/1.15} \times 1000 = \frac{2 \times 1000 \times 1.15}{1120} = 2.05\text{M}$$

24. **(b)** For this solution intermolecular interactions between *n*-heptane and ethanol are weaker than *n*-heptane-*n*-heptane & ethanol-ethanol interactions hence the solution of *n*-heptane and ethanol is non-ideal and shows positive deviation from Raoult's law.

25. **(c)** Let the mass of solution is 100 g. Then according to question, mass of H_2O = 25 g; mass of C_2H_5OH = 25 g; mass of CH_3COOH = 50 mg

$$\text{Number of moles} = \frac{\text{mass}}{\text{Molar mass}}$$

$$n_{H_2O} = \frac{25\text{ g}}{18\text{ g}}; \ n_{C_2H_5OH} = \frac{25\text{ g}}{46\text{ g}}; \ n_{CH_3COOH} = \frac{50\text{ g}}{60\text{ g}}$$

$$= 1.389 \qquad = 0.54 \qquad = 0.8$$

$$\therefore \text{mole fraction of water} = \frac{1.389}{1.389 + 0.54 + 0.8} = 0.5089$$

$$\approx 0.510$$

26. **(a)** Osmotic pressure (π) of isotonic solutions are equal. For solution of unknown substance C_1(concentration).

$C_1 = \frac{5.25/M}{V}$ where M represents molar mass.

For solution of urea, C_2 (concentration) $= \frac{1.5/60}{V}$

Given, $\pi_1 = \pi_2$ $\quad\because \pi = CRT$

$\therefore \ C_1RT = C_2RT$ or $C_1 = C_2$ or $\frac{5.25/M}{V} = \frac{1.5/60}{V}$

$\therefore$ M = 210 g mol^{-1}

27. **(c)** $\frac{P^o - P_s}{P^o} = \frac{w_B}{m_B} \times \frac{m_A}{w_A}; \ \frac{w_B}{m_B \times w_A} \times 1000 \times \frac{m_A}{1000}$

$$\Rightarrow \frac{P^o - P_s}{P^o} = \text{molality} \times \frac{m_A}{1000}$$

$$\frac{25 - 24.5}{25} = m \times \frac{18}{1000}$$

$$m = \frac{0.02 \times 1000}{18} = 1.11$$

28. **(b)** $\pi V = CRT$

$$\frac{\pi_1}{\pi_2} = \frac{C_1T_1}{C_2T_2}$$

$$\pi_1 = P, \pi_2 = 2\ atm, C_1 = C, C_2 = \frac{C}{2}$$

$T_1 = 600\,K$, $T_2 = 700\,K$

$$\frac{P}{2} = \frac{2 \times C \times 600}{C \times 700}$$

$$P = \frac{24}{7}.$$

29. **(a)** $i = \frac{\text{No. of particles after ionisation}}{\text{No. of particles before ionisation}}$

$$\begin{array}{cccc} CuCl_2 & \longrightarrow Cu^{2+} & + & 2Cl^- \\ 1 & 0 & & 0 \\ (1-\alpha) & \alpha & & 2\alpha \end{array}$$

$$i = \frac{1+2\alpha}{1}, \ i = 1 + 2\alpha$$

Assuming 100% ionization

So, $i = 1 + 2 = 3$

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

$$\Delta T_b = 3 \times 0.52 \times 0.1 = 0.156 \approx 0.16 \quad [m = \frac{13.44}{134.4} = 0.1]$$

30. **(a)** Depression in freezing point ∝ No. of particles.
(when concentration of different solutions is equal)
$Al_2(SO_4)_3$ provides five ions on ionisation

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

while KCl provides two ions

$$KCl \longrightarrow K^+ + Cl^-$$

$C_6H_{12}O_6$ and $C_{12}H_{22}O_{11}$ are not ionised so they have single particle in solution.
Hence, $Al_2(SO_4)_3$ have maximum value of depression in freezing point or lowest freezing point.

31. **(b)** In case of ionization of binary electrolyte

$$\text{Abnormal molecular mass} = \frac{M_{\text{normal}}}{1+\alpha}$$

Abnormal molecular mass = 40

$$\text{So, } 25 = \frac{40}{1+\alpha} \Rightarrow \alpha = 0.6$$

Ionisation percentage = 0.6 × 100 = 60%

32. **(c)** $P = \frac{w}{mv}RT$ since w and T are constant thus $P \propto \frac{1}{m}$

$P_2 > P_1 > P_3$

33. **(b)** $P^\circ_A = 7\times10^3$

$P^\circ_B = 12\times10^3$

$x_A = 0.4; x_B = 1 - 0.4$

$x_B = 0.6$

$P_{\text{total}} = P^\circ_A x_A + P^\circ_B x_B$

$= 7\times10^3\times0.4 + 12\times10^3\times0.6$

$= (7\times0.4 + 12\times0.6)\times10^3 = 10^4$

$$y_A = \frac{P^\circ_A x_A}{P_{\text{total}}} = \frac{7\times10^3\times0.4}{10^4}$$

$\therefore y_A = 0.28, y_B = 1 - 0.28 = 0.72$

34. **(b)** According to Raoult's law,

$$\frac{\Delta p}{p^\circ} = \frac{n}{n+N} \text{ (mole fraction of solute)}$$

$$\frac{10}{p^\circ} = 0.2 \therefore p^\circ = 50 \text{ mm of Hg}$$

For other solution of same solvent,

$$\frac{20}{p^\circ} = \frac{n}{n+N} \text{ (Mole fraction of solute)}$$

$\Rightarrow \frac{20}{50}$ = Mole fraction of solute

⇒ Mole fraction of solute = 0.4
As mole fraction of solute + mole fraction of solvent = 1
Hence, mole fraction of solvent = 1 – 0.4 = 0.6

35. **(b)** $p_{\text{total}} = p_A^\circ x_A + p_B^\circ x_B$

$550 = p_A^\circ \times \frac{1}{4} + p_B^\circ \times \frac{3}{4}$

$p_A^\circ + 3p_B^\circ = 550\times4$ ...(i)

In second case

$p_{\text{total}} = p_A^\circ \times \frac{1}{5} + p_B^\circ \times \frac{4}{5}$

$p_A^\circ + 4p_B^\circ = 560\times5$ ...(ii)

Subtract (i) from (ii)

$p_B^\circ = 560\times5 - 550\times4 = 600$

$\therefore p_A^\circ = 400$

36. **(a)** The vapour pressure of a solution of glucose in water can be calculated using the relation

$$\frac{p^\circ - p_s}{p_s} = \frac{\text{Moles of glucose in solution}}{\text{Moles of water in solution}}$$

$$\text{or } \frac{17.5 - p_s}{p_s} = \frac{18/180}{178.2/18} \qquad [\because p^\circ = 17.5]$$

or $17.5 - p_s = \frac{0.1\times p_s}{9.9}$ or $p_s = 17.325$ mm Hg.

37. **(b)** Vapour pressure of pure water at boiling point
= 1 atm = 1.013 bar
Vapour pressure of solution (p_s) = 1.004 bar
Let mass of solution = 100 g
Mass of solute = 2 g
Mass of solvent = 100 – 2 = 98 g

$$\frac{P^\circ - P_s}{P^\circ} = \frac{n_2}{n_1+n_2} = \frac{W_2/M_2}{W_1/M_1}$$

$$\frac{1.013-1.004}{1.013} = \frac{2}{M_2}\times\frac{18}{98}$$

$$\text{or} \quad M_2 = \frac{2\times18}{98}\times\frac{1.013}{0.009} = 41.35 \text{ g mol}^{-1}$$

38. **(a)** $\Delta T_b = i K_b m$
For water ⇒ 1000 mL = 1000 g

$$\text{Molality of NaCl} = \frac{w/M.W.}{W(\text{solvent})}\times1000$$

$$= \frac{58.5/58.5}{1000}\times1000 = 1\text{ m}$$

$$\text{Molality of glucose} = \frac{180/180}{1000}\times1000 = 1\text{ m}$$

i for NaCl = 2; *i* for glucose = 1

∴ ΔT_b for NaCl > ΔT_b for glucose

39. **(a)** $\pi = i\,CRT$

$\pi_{C_2H_5OH} = 1\times0.500\times R\times T = 0.5\,RT$

$\pi_{Mg_3(PO_4)_2} = 5\times0.100\times R\times T = 0.5\,RT$

$\pi_{KBr} = 2\times0.250\times R\times T = 0.5\,RT$

$\pi_{Na_3PO_4} = 4\times0.125\times RT = 0.5\,RT$

Since the osmotic pressure of all the given solutions is equal. Hence all are isotonic solution.

40. (c) Concentration of NaCl solution $= \frac{7}{58.5} = 0.119\text{M}$

Concentration of $MgCl_2$ solution $= \frac{7}{95} = 0.074\text{M}$

Hence, particles (ions) in NaCl solution

$(B) = 2 \times 0.119 = 0.238\text{ M}$

Hence, particles (ions) in $MgCl_2$ solution

$(A) = 3 \times 0.074 = 0.222\text{ M}$

As the particles in solution B(NaCl) is more than in solution A, the osmotic pressure of solution B will be greater than that of solution A.

41. (b) Given molecular mass of sucrose = 342 g

Moles of sucrose $= \frac{100}{342} = 0.292$ mole

Moles of water $N = \frac{1000}{18} = 55.5$ moles and

Vapour pressure of pure water $P^o = 23.8$ mm Hg

According to Raoult's law,

$$\frac{\Delta P}{P^o} = \frac{n}{n+N} \Rightarrow \frac{\Delta P}{23.8} = \frac{0.292}{0.292 + 55.5}$$

$$\Delta P = \frac{23.8 \times 0.292}{55.792} = 0.125 \text{ mm Hg}$$

42. (d) At 1 atmospheric pressure the boiling point of mixture is 80°C.

At boiling point the vapour pressure of mixture, $p_T = 1$ atmosphere = 760 mm Hg.

Using the relation,

$p_T = p_A^{\circ} x_A + p_B^{\circ} x_B$, we get

$$p_T = 520x_A + 1000(1 - x_A)$$

$$\{\because p_A^o = 520 \text{ mm Hg},\ p_B^{\circ} = 1000 \text{ mm Hg},$$

$$x_A + x_B = 1\}$$

$$760 = 520x_A + 1000 - 1000x_A$$

$$x_A = \frac{240}{480} = \frac{1}{2} \text{ or 50 mol percent}$$

43. (d) Let the mass of methane and oxygen = m g.

Mole fraction of O_2

$$= \frac{\text{Moles of } O_2}{\text{Moles of } O_2 + \text{Moles of } CH_4}$$

$$= \frac{m/32}{m/32 + m/16} = \frac{m/32}{3m/32} = \frac{1}{3}$$

Partial pressure = Total pressure × mole fraction

$$p_{O_2} = p \times \frac{1}{3} = \frac{1}{3}p$$

44. (b) Given mass of solute = 8.1 g

Mass of solvent = 100 g,

For HBr,

$$\begin{array}{lll} HBr & \longrightarrow H^+ & + Br^- \\ 1 & 0 & 0 \\ (1-\alpha) & \alpha & \alpha \end{array}$$

$\alpha = 90\% = 0.9$

$i = 1 + \alpha = 1 + 0.9 = 1.9$

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

$$= 1.86 \times \frac{\text{moles of solute}}{\text{mass of solvent in kg}} \times 1.9$$

$$= 1.86 \times \frac{8.1/81}{100/1000} \times 1.9$$

$$= 1.86 \times 1 \times 1.9 = 3.534\text{K}$$

$$T_f = T_f^{\circ} - \Delta T_f$$

or $T_f = 0 - 3.534\text{ K}$

$\therefore$ $T_f = -3.534\text{K}$

45. (b) Sodium sulphate dissociates as

$$Na_2SO_4 \longrightarrow 2Na^+ + SO_4^{2-}$$

hence van't hoff factor $i = 3$

Now $\Delta T_f = iK_f.m$

$$= 3 \times 1.86 \times 0.01 = 0.0558\text{ K}$$

46. (c) Given, W = 500 cc = 500 g

($\because$ density of water = 1g/cc)

As, Density $= \frac{\text{Weight}}{\text{Volume}}$

Weight = Density × Volume $= \frac{1\text{g}}{\text{cc}} \times 500\text{cc} = 500\text{g}$

w = amount of urea = ?

$K_f = 1.86°\text{C}; \Delta T_f = 0.186°\text{C}$

M = Molecular mass of urea = 60

$$\text{M} = \frac{1000 \times K_f \times \text{w}}{\text{W} \times \Delta T_f}$$

$$60 = \frac{1000 \times 1.86 \times \text{w}}{500 \times 0.186}$$

$$60 = 20 \times \text{w}$$

$$\therefore \quad \text{w} = \frac{60}{20} = 3\text{g}$$

47. (a) Let volume of A = xL

Volume of B = (2 – x) L

$\because$ $M_1V_1 + M_2V_2 = M(V_1 + V_2)$

$0.5x + 0.1(2 - x) = 0.2(x + 2 - x)$

(for HCl, 0.1M HCl = 0.1N HCl)

$\Rightarrow$ $0.5x + 0.2 - 0.1x = 0.2 \times 2$

$\Rightarrow$ $0.4x + 0.2 = 0.4 \Rightarrow 0.4x = 0.2$

$$\therefore \quad x = \frac{0.2}{0.4} = 0.5\text{L}$$

Since x = 0.5 L (vol. of A), then

2 – x = 2 – 0.5 = 1.5 L (Vol. of B).

48. **(d)** $\Delta T_f = T_0 - T_f = 5.3 - 4.47 = 0.83$

molality of solvent $m = \dfrac{W \times 1000}{M \times w}$

$= \dfrac{0.223}{136} \times \dfrac{1000}{4.4} = 0.373$

(Molecular weight of phenylacetic acid = 136)

$K_f = 5.12$

$\Delta T_f = i.K_f m$

$$i = \frac{\Delta T_f}{K_f m} = \frac{0.83}{5.12 \times 0.373} = 0.435$$

i is nearly equal to 0.56 it means phenylacetic acid undergoes dimerization in benzene.

49. **(c)** Let the density of solution be 'd'

Molarity of solution given = 3.6

i.e., 1 litre of solution contains 3.6 moles of H_2SO_4

or 1 litre of solution contains 3.6 × 98 g of H_2SO_4

Since, the solution is 29% by mass

100 g solution contains 29 g H_2SO_4

$\dfrac{100}{d}$ mL solution contains 29 g H_2SO_4

1000 mL solution contains 3.6 × 98 g H_2SO_4

$$\therefore\ 3.6 \times 98 = \frac{29 \times d}{100} \times 1000 \Rightarrow d = 1.22$$

50. **(b)** Relative lowering of vapour pressure is given by the formula :

$$\frac{p^o - p_s}{p^o} = \frac{w_A}{m_A} \times \frac{m_B}{w_B}$$

Given, relative lowering of vapour pressure

$$= \frac{P^o - P_s}{P^o} = \frac{10}{100}$$

$m_A = 60, m_B = 18, w_B = 180, w_A = x$

$$\therefore\ \frac{10}{100} = \frac{x/60}{180/18} \Rightarrow \frac{1}{10} = \frac{x/60}{10} \Rightarrow x = 60$$

Thus, 60 g of the solute must be added to 180 g of water so that the vapour pressure of water is lowered by 10%.

51. **(b)** As $\Delta T_f = K_f.\ m$

$\Delta T_b = K_b.\ m$

Hence, we have $m = \dfrac{\Delta T_f}{K_f} = \dfrac{\Delta T_b}{K_b}$

or $\Delta T_f = \Delta T_b \dfrac{K_f}{K_b}$

$\Rightarrow [\Delta T_b = 100.18 - 100 = 0.18°C]$

$= 0.18 \times \dfrac{1.86}{0.512} = 0.654°C$

As the Freezing Point of pure water is 0°C,

$\Delta T_f = 0 - T_f$

$0.654 = 0 - T_f$

$\therefore\ T_f = -0.654$

Thus the freezing point of solution will be – 0.654°C.

52. **(a)** $K_H = 100$ k bar, p = 1 bar

p = K_H × mole fraction

$$\text{Mole fraction } = \frac{p}{K_H} = \frac{1}{100 \times 10^3} = 10^{-5}$$

$$\text{Mole fraction } = \frac{\text{Mole of gas}}{\text{Total moles}}$$

$$\text{Moles of water } = \frac{\text{Weight of water}}{\text{Molecular weight of water}}$$

Weight of water = 1000 g (∵ 1000 mL = 1000 g)

$$\text{No. of moles } = \frac{1000}{18} = 55.5$$

$$\text{Mole fraction} = 10^{-5} = \frac{x}{55.5 + x}$$

As 55.5 >>> x, thus neglecting 'x' from denominator

$$10^{-5} = \frac{x}{55.5} \Rightarrow x = 55.5 \times 10^{-5} \text{ moles}$$

or 55.5 × 10^{-2} millimoles = 0.555 millimoles

53. **(d)** For isotonic solution : $\dfrac{w_1}{m_1 V_1} = \dfrac{w_2}{m_2 V_2}$

w_1 = mass of glucose = 0.05 × 180 g = 9 g

m_1 = molecular mass of glucose = 180 g

Assuming $V_1 = V_2 = 1L$

w_2 = mass of compound = 6 g

m_2 = molecular mass of compound = ?

$$\therefore\ \frac{9}{180} = \frac{6}{x} \Rightarrow x = 120\text{ g}$$

$$\therefore\ \frac{\text{Molecular mass}}{\text{Empirical mass}} = n \Rightarrow n = \frac{120}{30} = 4$$

∴ Molecular formula = $C_4H_8O_4$

54. **(d)** $2CH_3COOH \rightleftharpoons (CH_3COOH)_2$

$\quad 1 - \alpha \qquad\qquad \alpha/2$

$i = 1 - \alpha + \alpha/2 = 1 - \alpha/2$

here α = digree of dissociation

$\Delta T_f = iK_f m$

$$0.45 = \left(1 - \frac{\alpha}{2}\right)(5.12)\frac{\left(\frac{0.2}{60}\right)}{\frac{20}{1000}}$$

$$1 - \frac{\alpha}{2} = 0.527$$

$\alpha = 0.946$

Degree of dissociation = 94.6%

55. **(d)** Using relation,

$$\frac{p^\circ - p_s}{p_s} = \frac{w_2 M_1}{w_1 M_2}$$

where w_1, M_1 = mass in g and mol. mass of solvent
w_2, M_2 = mass in g and mol. mass of solute
Let $M_2 = x$
$p° = 185$ torr
$p_s = 183$ torr

$$\frac{185-183}{183} = \frac{1.2 \times 58}{100x} \text{ (Mol. mass of acetone = 58)}$$

$x = 64$
$\therefore$ Molar mass of substance = 64