

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
 $\text{Mole of solvent (H}_2\text{O)} = \frac{1000\text{g}}{18\text{g}} = \frac{1000}{18}$
 $\text{Mole fraction of solute} = \frac{1}{1 + \frac{1000}{18}} = \frac{18}{1018} = 0.0176 \approx 0.018$
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) No. of moles of $\text{NaCl} = \frac{M \times V}{1000} = \frac{0.25 \times 100}{1000} = 0.025$
 $\text{NaCl} \rightarrow \text{Na}^+ + \text{Cl}^-$
 $\text{No. of moles of Na}^+ \text{ ions} = 0.025$
 $\text{No. of Na}^+ \text{ ions} = 0.025 \times 6.023 \times 10^{23} = 1.505 \times 10^{22}$
5. (b) Molarity is intrinsic property. Hence, it is independent of amount of solution.
6. (c) $N_1 V_1 + N_2 V_2 = NV$
 $4x + 10(1-x) = 6 \times 1; -6x = -4; x = 0.67$
 $\text{Thus } 0.67 \text{ litre of } 4\text{N HCl and}$
 $1-x = 1-0.67 = 0.33 \text{ litre of } 10 \text{ N HCl}$
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) $\Delta T_f = K_f \times m = 1.86 \times 0.5 = 0.93^\circ\text{C}; T_f = -0.93^\circ\text{C}$
9. (c) $\text{Na}_2\text{SO}_4 \longrightarrow 2\text{Na}^+ + \text{SO}_4^{2-}$

Initial	1	0	0
Final	$1-\alpha$	2α	α

 $i = 1 - \alpha + 2\alpha + \alpha = 1 + 2\alpha$
10. (d) $2\text{NaHCO}_3 + \text{H}_2\text{C}_2\text{O}_4 \longrightarrow \text{Na}_2\text{C}_2\text{O}_4 + 2\text{CO}_2 + 2\text{H}_2\text{O}$
 $\text{Moles of CO}_2 \text{ evolved} = \frac{0.25}{25 \times 10^3} = 10^{-5}$
 $\therefore \text{moles of NaHCO}_3 = 10^{-5}$
 $\therefore \text{mass of NaHCO}_3 = 84 \times 10^{-5} \text{ g}$
 $= 0.84 \times 10^{-3} \text{ g} = 0.84 \text{ mg}$
 $\therefore \% \text{ by weight} = \frac{0.84}{10} \times 100 = 8.4\%$
11. (b) We know that, $M_1 V_1 = M_2 V_2$
 $0.05 \times V_1 = 0.4 \times 100$
 $\therefore V_1 = \frac{0.4 \times 100}{0.05} = 800$
 $V_2 - V_1 = 800 - 100 \Rightarrow 700 \text{ mL}$
12. (b) No. of moles in 0.02 M H_2SO_4 solution
 $[\text{H}_3\text{O}^+] = 2 \times 0.02 = 0.04 \text{ moles}$
 $\therefore 2 \text{ litre solution contains } 0.08 \text{ mole of H}_3\text{O}^+.$
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) $\text{F}^- \text{ ions in ppm} = \frac{0.2}{500} \times 10^6 = 400$
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) $\Delta T_b = K_b \times m \times i = 0.52 \times 1 \times 2 = 1.04$
 $\therefore \Delta T_b = 100 + 1.04 = 101.04^\circ\text{C}$
19. (d) Osmotic pressure is a colligative property. Hence resulting osmotic pressure of the solution is given by
 $\pi_T = \pi_1 + \pi_2 + \pi_3 \dots\dots\dots$
 $\pi_T = 1.64 + 2.46 = 4.10 \text{ atm.}$
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 H₂SO₄,

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

Hence 4 N H₂SO₄ and 2N HCl solution will have same molar concentration.

22. (c) Equivalent weight of orthophosphoric acid

$$(\text{H}_3\text{PO}_4) = \frac{3 + 31 + 64}{3} = \frac{98}{3}$$

Now 100 g solution contains 70 g H₃PO₄

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

equivalent of H₃PO₄

∴ Normality of solution,

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

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

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

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

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

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 H₂O = 25 g; mass of C₂H₅OH = 25 g; mass of CH₃COOH = 50 mg

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

$$n_{\text{H}_2\text{O}} = \frac{25 \text{ g}}{18 \text{ g}}; n_{\text{C}_2\text{H}_5\text{OH}} = \frac{25 \text{ g}}{46 \text{ g}}; n_{\text{CH}_3\text{COOH}} = \frac{50 \text{ g}}{60 \text{ g}}$$

$$= 1.389 \quad = 0.54 \quad = 0.8$$

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

$$\approx 0.510$$

26. (a) Osmotic pressure (π) of isotonic solutions are equal. For solution of unknown substance C₁(concentration).

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

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

$$\text{Given, } \pi_1 = \pi_2 \quad \therefore \pi = CRT$$

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

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

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

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

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

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

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

$$\frac{\pi_1}{\pi_2} = \frac{C_1 T_1}{C_2 T_2}$$

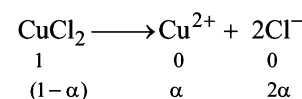
$$\pi_1 = P, \pi_2 = 2 \text{ atm}, C_1 = C, C_2 = \frac{C}{2}$$

$$T_1 = 600 \text{ K}, T_2 = 700 \text{ K}$$

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

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

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



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

Assuming 100% ionization

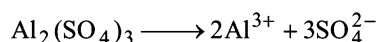
$$\text{So, } i = 1 + 2 = 3$$

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

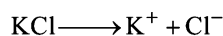
$$\Delta T_b = 3 \times 0.52 \times 0.1 = 0.156 \approx 0.16 \quad [m = \frac{13.44}{134.4} = 0.1]$$

30. (a) Depression in freezing point \propto No. of particles.
(when concentration of different solutions is equal)

$\text{Al}_2(\text{SO}_4)_3$ provides five ions on ionisation



while KCl provides two ions



$\text{C}_6\text{H}_{12}\text{O}_6$ and $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ are not ionised so they have single particle in solution.

Hence, $\text{Al}_2(\text{SO}_4)_3$ have maximum value of depression in freezing point or lowest freezing point.

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

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

$$\text{Abnormal molecular mass} = 40$$

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

$$\text{Ionisation percentage} = 0.6 \times 100 = 60\%$$

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) $P_A^\circ = 7 \times 10^3$

$$P_B^\circ = 12 \times 10^3$$

$$x_A = 0.4; x_B = 1 - 0.4$$

$$x_B = 0.6$$

$$P_{\text{total}} = P_A^\circ x_A + P_B^\circ x_B$$

$$= 7 \times 10^3 \times 0.4 + 12 \times 10^3 \times 0.6$$

$$= (7 \times 0.4 + 12 \times 0.6) \times 10^3 = 10^4$$

$$y_A = \frac{P_A^\circ x_A}{P_{\text{total}}} = \frac{7 \times 10^3 \times 0.4}{10^4}$$

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

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

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

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

For other solution of same solvent,

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

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

$$\Rightarrow \text{Mole fraction of solute} = 0.4$$

As mole fraction of solute + mole fraction of solvent = 1

Hence, mole fraction of solvent = $1 - 0.4 = 0.6$

35. (b) $P_{\text{total}} = P_A^\circ x_A + P_B^\circ x_B$

$$550 = P_A^\circ \times \frac{1}{4} + P_B^\circ \times \frac{3}{4}$$

$$P_A^\circ + 3P_B^\circ = 550 \times 4 \quad \dots(i)$$

In second case

$$P_{\text{total}} = P_A^\circ \times \frac{1}{5} + P_B^\circ \times \frac{4}{5}$$

$$P_A^\circ + 4P_B^\circ = 560 \times 5 \quad \dots(ii)$$

Subtract (i) from (ii)

$$P_B^\circ = 560 \times 5 - 550 \times 4 = 600$$

$$\therefore P_A^\circ = 400$$

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

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

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

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

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

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

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

$$\text{or } M_2 = \frac{2 \times 18}{98} \times \frac{1.013}{0.009} = 41.35 \text{ g mol}^{-1}$$

38. (a) $\Delta T_b = i K_b m$

For water $\Rightarrow 1000 \text{ mL} = 1000 \text{ g}$

$$\text{Molality of NaCl} = \frac{w/M.W.}{W(\text{solvent})} \times 1000$$

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

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

i for NaCl = 2; i for glucose = 1

$$\therefore \Delta T_b \text{ for NaCl} > \Delta T_b \text{ for glucose}$$

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

$$\pi_{\text{C}_2\text{H}_5\text{OH}} = 1 \times 0.500 \times R \times T = 0.5 RT$$

$$\pi_{\text{Mg}_3(\text{PO}_4)_2} = 5 \times 0.100 \times R \times T = 0.5 RT$$

$$\pi_{\text{KBr}} = 2 \times 0.250 \times R \times T = 0.5 RT$$

$$\pi_{\text{Na}_3\text{PO}_4} = 4 \times 0.125 \times R \times T = 0.5 RT$$

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 = $\frac{100}{342} = 0.292\text{ mole}$

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

Vapour pressure of pure water $P^\circ = 23.8\text{ mm Hg}$

According to Raoult's law,

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

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

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

At boiling point the vapour pressure of mixture,

$p_T = 1\text{ atmosphere} = 760\text{ mm Hg}$.

Using the relation,

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

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

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

$x_A + x_B = 1\}$

$760 = 520x_A + 1000 - 1000x_A$

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

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

Mole fraction of O_2

$$= \frac{\text{Moles of } \text{O}_2}{\text{Moles of } \text{O}_2 + \text{Moles of } \text{CH}_4}$$

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

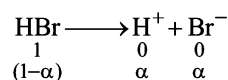
Partial pressure = Total pressure \times mole fraction

$p_{\text{O}_2} = p \times \frac{1}{3} = \frac{1}{3}p$

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

Mass of solvent = 100 g,

For HBr,



$\alpha = 90\% = 0.9$

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

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

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

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

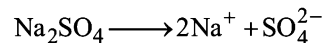
$= 1.86 \times 1 \times 1.9 = 3.534\text{K}$

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

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

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

45. (b) Sodium sulphate dissociates as



hence van't hoff factor $i = 3$

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

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

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

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

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

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

w = amount of urea = ?

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

M = Molecular mass of urea = 60

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

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

$60 = 20 \times w$

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

47. (a) Let volume of A = xL

Volume of B = (2 - x) L

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

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

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

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

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

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

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

$2 - x = 2 - 0.5 = 1.5\text{ L (Vol. of B)}$.

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

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

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

(Molecular weight of phenylacetic acid = 136)

$$K_f = 5.12$$

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

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

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

$$\frac{100}{d} \text{ mL solution contains 29 g } \text{H}_2\text{SO}_4$$

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 $\Delta T_f = K_f \cdot m$

$$\Delta T_b = K_b \cdot m$$

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

$$\text{or } \Delta T_f = \Delta T_b \frac{K_f}{K_b}$$

$$\Rightarrow [\Delta T_b = 100.18 - 100 = 0.18^\circ\text{C}]$$

$$= 0.18 \times \frac{1.86}{0.512} = 0.654^\circ\text{C}$$

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

$$p = K_H \times \text{mole fraction}$$

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

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

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

$$\text{Weight of water} = 1000 \text{ g } (\because 1000 \text{ mL} = 1000 \text{ g})$$

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

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

As $55.5 \gg x$, thus neglecting 'x' from denominator

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

$$\text{or } 55.5 \times 10^{-2} \text{ millimoles} = 0.555 \text{ millimoles}$$

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

$$w_1 = \text{mass of glucose} = 0.05 \times 180 \text{ g} = 9 \text{ g}$$

$$m_1 = \text{molecular mass of glucose} = 180 \text{ g}$$

Assuming $V_1 = V_2 = 1\text{L}$

$$w_2 = \text{mass of compound} = 6 \text{ g}$$

$$m_2 = \text{molecular mass of compound} = ?$$

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

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

$$\therefore \text{Molecular formula} = \text{C}_4\text{H}_8\text{O}_4$$

54. (d) $2\text{CH}_3\text{COOH} \rightleftharpoons (\text{CH}_3\text{COOH})_2$

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

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

here α = degree of dissociation

$$\Delta T_f = i K_f m$$

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

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

$$\alpha = 0.946$$

Degree of dissociation = 94.6%

55. (d) Using relation,

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

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 $M_2 = x$

$p^\circ = 185 \text{ torr}$

$p_s = 183 \text{ torr}$

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

$$x = 64$$

\therefore Molar mass of substance = 64