

Solutions-solutions

SUBJECTIVE PROBLEMS:

Sol 1.

(i) Molarity = Moles of solute/Volume of solution in L

(ii) $N_1V_1 = N_2V_2$

A 13% solution (by weight) contains 13 g of solute (i.e H_2SO_4) per 100 gm of solution

Moles of solute = Mass of H_2SO_4 /M. wt. of $H_2SO_4 = 13/98 = 0.1326$

Volume of solution in L

= Mass of solution/density of solution * 1000 = $100/1.02 * 1000 = 0.0980$ Litre

∴ Molarity of solution = $0.1326/0.0980 = 1.35$ M

Again, Molality = Moles of solute/Mass of solvent in kg

Mass of solute in 100 ml of solution = 13 g [13% solution]

Mass of solvent = Mass of solution – Mass of solute

= $100 - 13 = 87$ g

∴ Molality = $13/98/87/100 = 1.57$ m

Normality = Molarity * Mol. wt/Eq. wt. or $1.35 * 98/49 = 2.70$ N

$N_1 = 2.70$, $V_1 = 100$ ml, $N_2 = 1.5$, $V_2 = ?$ [∴ Eq. wt = $98/2$ $2H_2SO_4 = 49$]

$N_1V_1 = N_2V_2$

$2.70 * 100 = 1.5 * v_2$

Or $V_2 = 2.70 * 100/1.5 = 180$ ml. ∴ 100 ml of this acid should be diluted to 180 ml to prepare 1.5 solution.

Sol 2.

Molarity = Moles of solute/Volume of solution in L

= Mass of solute/M. wt. of solute/Mass of solution/density of solution * 1000

$$M = 86/98/100/1.787 * 1000 = 0.8775/55.5 * 1000 = 15.81 \text{ M}$$

$$M_1V_1 = M_2V_2$$

$$M_1 = 15.81, V_1 = ?$$

$$M_2 = 0.2, V_2 = 1 \text{ L} = 1000 \text{ ml}$$

$$\therefore 15.81 * V_1 = 0.2 * 1000$$

$$\text{Or } V_1 = 0.2 * 1000 / 15.81 = 12.65 \text{ ml}$$

\therefore Amount of acid to be used to make 1 L of 0.2 M H_2SO_4 – 12.65.

Sol 3.

$$N_1 = 1, V_1 = ?, N_2 = 26.7, V_2 = 0.4$$

$$N_1V_1 = N_2V_2$$

$$1 * V_1 = 26.7 * 0.4$$

$$V_1 = 26.7 * 0.4 / 1 = 10.68$$

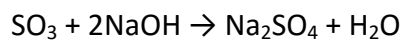
49g (\because eqwt of $\text{H}_2\text{SO}_4 = 49$) of H_2SO_4 will be neutralized by = 1N 1000 ml NaOH

\therefore 0.5g of H_2SO_4 will be neutralized

$$= 1000/49 * 0.5 = 10.20 \text{ ml } 1\text{N NaOH}$$

Volume of 1 N NaOH used by dissolved

$$SO_3 = 10.68 - 10.20 = 0.48 \text{ ml}$$



$$\therefore \text{Eqwt of } \text{SO}_3 = \text{Molwt}/2 = 80/2 = 40$$

Wt of SO_3 in 0.48 ml of 1 M solution

$$= 40/1000 * 0.48 = 0.0192 \text{ g}$$

$$\% \text{ of } \text{SO}_3 = 0.0192/0.5 * 100 = 3.84 \%$$

Sol 4.

$$p^\circ - p/p^\circ = n/N \text{ [Raoult's Equation]}$$

Let the molality of the solution = m

Now the solution contain 'm' moles of solute per 1000 gm of benzene

Vapour pressure of benzene, $p^\circ = 639.7 \text{ mm}$

Vapour pressure of solution, $p = 63.9 \text{ mm}$

Moles of benzene (Mol. wt. 78), $N = 1000/78$

Moles of solute, $n = ?$

Substitute these values in the Raoult's equation

$$p^\circ - p/p^\circ = n/N \text{ or } 639.7 - 631.9/639.7 = n * 78/1000$$

$$\text{or } 7.8/639.7 = 78 n/1000$$

$$\therefore n = 1000 * 7.8/78 * 639.7 = 0.156$$

Hence, molality of solution = 0.156 m

Sol 5.

$$P_T = p_1^\circ + p_2^\circ x_2$$

At 300 K, the vapour pressure of the solution containing 1 mole of A & 3 moles of B = 550 mm of Hg and vapour pressure of the solution containing 1 mole of A & 4 moles of B at 300 K = 560 mm Hg.

Let the vapour pressure of pure A = p_1°

And the vapour pressure of pure B = p_2°

Further, let x_1 and x_2 be the mole fractions of A and B in the solutions. Then the total vapour pressure of solution

$$P_{\text{total}} = p_1^0 x_1 + p_2^0 x_2 \quad \dots(i)$$

In solution 1, $p_{\text{total}} = 550$ mm,

$$\therefore x_1 = 1/1 - 3, x_2 = 3/1 + 3 \quad [\because \text{moles of A} = 1 \text{ moles of B} = 3]$$

Substituting these values in (i) we get

$$550 = p_1^0 * \frac{1}{4} + p_2^0 * \frac{3}{4}$$

$$\text{Or } 550 = p_1^0/4 + 3p_2^0/4$$

$$2200 = p_1^0 + 3p_2^0 \quad \dots(ii)$$

In solution 2, $p_{\text{total}} = 560$ mm,

$$x_1 = 1/1 + 4, x_2 = 4/1 + 4 \quad [\because \text{moles of A} = 1 \text{ moles of B} = 4]$$

Substituting the various values, in equ. (i) we get

$$560 = p_1^0 * \frac{1}{5} + p_2^0 * \frac{4}{5}$$

$$\text{Or } 560 p_1^0/5 + 4p_2^0/5$$

$$2800 = p_1^0 + 4p_2^0 \quad \dots(iii)$$

Solving equation (ii) and (iii), we get

$$P_2^0 = 6400 \text{ mm of Hg}$$

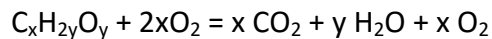
$$p_1^0 = 400 \text{ mm of Hg}$$

\therefore Vapour pressure of pure A = 400 mm of Hg

And vapour pressure of pure B = 600 mm of Hg

Sol 6.

The chemical equation for the combustion of organic compound $C_xH_{2y}O_y$ can be represented as:



The gases obtained after cooling = $x + x = 2x$

$$\therefore 2x = 2.24 \text{ litres} \quad [\because H_2O \text{ is in liquid state}]$$

$$\text{Or } x = 2.24/2 = 1.12 \text{ litres}$$

$$\text{Number of moles of } CO_2 = 1.12 \text{ litres}/22.4 \text{ litres mole} \quad [\because 22.4 \text{ L at NTP} = 1 \text{ mole}]$$

$$= 1/2 \text{ mole} = 0.05 \text{ mole}$$

The empirical formula of the organic compound is $C(H_2O)$... (i)

The mole fraction of the solute (A)

= relative decrease in vapour pressure of the solvent (B)

$$p^0 - p/p^0 = W_A/M_A / W_A/M_A + W_B/M_B$$

$$\text{Or } 0.104/17.5 = 50/M_A / 50/M_A + 1000/18 \quad [M_A = \text{mol. wt. g A}]$$

$$\text{Or } 0.104/17.5 = 50/M_A(50 * 18 + 100M_A/18M_A)$$

$$\text{Or } 104/17500 = 50 * 18/900 + 1000/M_A$$

$$\text{Or } M_A = 150.6$$

Molecular wt. of the organic compound

$$(CH_2O)_n = 150$$

$$\text{Molecular wt. } CH_2O = 12 + 2 + 16 = 30$$

$$\therefore 30 * n = 150 \quad [\because (CH_2O)_n = \text{mol, formula}]$$

$$\text{Or } n = 150/30 = 5$$

\therefore Molecular formula of the given organic compound is **$(CH_2O)_5$** or **$C_5H_{10}O_5$** .

Sol 7.

If they form an ideal solution which obeys 'Raoult's Law and for which

$$\Delta H_{\text{mixing}} = 0 \text{ and } \Delta V_{\text{mixing}} = 0$$

Thus we can separate two volatile and miscible liquids by fractional distillation if, they should not form azeotropic solutions.

Sol 8.

$$P_{\text{total}} = p_A + p_B$$

$$\text{Molecular weight of CH}_3\text{OH} = 12 + 3 + 16 + 1 = 32$$

$$\text{Molecular weight of C}_2\text{H}_5\text{OH} = 24 + 5 + 16 + 1 = 46$$

According to Raoult's law

$$P_{\text{total}} = p_1 + p_2$$

Where P_{total} = Total vapour pressure of the solution

P_1 = Partial vapour pressure of one component

P_2 = Partial vapour pressure of other component

Again, p_1 = Vapour pressure (p_1^0) * mole fraction

Similarly, p_2 = Vapour pressure (p_2^0) * mole fraction

$$\text{Mole fraction of CH}_3\text{OH} = 40/32 / 40/32 + 60/46 = 0.49$$

$$\text{Mole fraction of ethanol} = 60/46 / 60/46 + 40/32 = 0.51$$

NOTE THIS STEP : Thus now let us first calculate the partial vapour pressures, i.e., p_1 and p_2 of the two component.

Partial vapour pressure of $\text{CH}_3\text{OH}(p_1)$

$$= 44.5 * 0.51 = 22.69 \text{ mm } \therefore \text{ Total vapour pressure of the solution}$$

$$= 43.48 + 22.69 \text{ mm} = \mathbf{66.17 \text{ mm}}$$
 Mole fraction of CH_3OH in vapour = $43.48/66.17 = 0.65$

Sol 9.

Molality, M = moles of solute/kg of solvent

Mole fraction, $x_A = n_A/n_A+n_B$, $x_B = n_B/n_A+n_B$

$$p_1 = x_1 p_1^0$$

$$\therefore x_1 = p_1/p_1^0 = 750/760 = 0.9868$$

$$x_2(\text{solute}) = 1 - 0.9868 = 0.0132$$

$$\text{molality, } m = x_2/x_1 M_1 * 1000 = 0.0132 * 1000/0.9868 * 18 = \mathbf{0.7503 \text{ mol kg}^{-1}}$$

ALTERNATIVE SOLUTION :

Given that :

Temperature = 273 K

boiling point of H₂O = 373 K

\therefore vapour pressure H₂O = 76 cm

We have,

$$P^0 - P_s/P_s = w^*M/w^*M$$

\therefore molality

$$= w/w^*M * 1000 = P^0 - P_s/P_s * 1/M * 1000$$

$$= 760 - 750/750 * 1/18 * 1000$$

$$= 0.741 \text{ mol/kg of solvent}$$

Also we have,

$$= 1000 * 1.72 * 20/50 * 2$$

$$= 344$$

van't Hoff factor (i) = actual mol. wt./calculate mol. wt = 172/344 = 0.5

$$P^0 - P_s / P_s = n/n + N$$

$$\therefore \text{mole fraction} = P^0 - P_s / P^0 = 760 - 750 / 760$$

$$= 10 / 760 = 0.013$$

Sol 10.

According to Rault's law,

$$P^0 - P / P^0 = w/m / w/m + W/M$$

$$\text{Here, } P^0 = 640 \text{ mm} \quad p = 600 \text{ mm}$$

$$W = 2.175 \text{ g} \quad W = 39.0$$

$$M = ? \quad M = 78$$

Substituting the various values in the above equation for Rault's law :

$$640 - 600 / 640 = 2.175/m / 2.175/m + 39/78$$

$$1/16 = 2.175/2.175 + 0.5 m$$

$$m = 65.25$$

Sol 11.

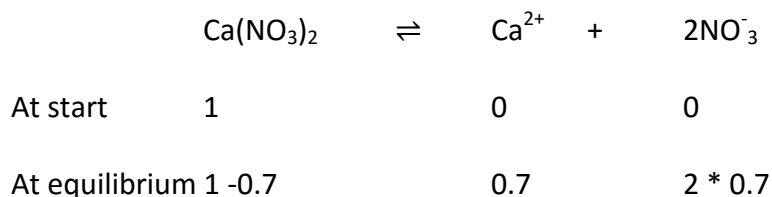
First find moles of $\text{Ca}(\text{NO}_3)_2$ and water. Then use the expression

$$P^0 - P / P^0 = n/n + N \text{ to find vapour pressure of solution}$$

Let initially 1 mole of $\text{Ca}(\text{NO}_3)_2$ is taken

$$\text{Degree of dissociation of } \text{Ca}(\text{NO}_3)_2 = 70/100 = 0.7$$

Ionization of $\text{Ca}(\text{NO}_3)_2$ can be represent as



∴ Total number of moles in the solution at equilibrium

$$= (1 - 0.7) + 0.7 + 2 * 0.7 = 2.4$$

No. of moles when the solution contains 1 gm of calcium nitrate instead of 1 mole of the salt

$$= 2.4/164 \text{ (164 is the mol. wt. of Cal. nitrate)}$$

∴ No. of moles of the solute in the solution containing 7 g of salt, i.e.,

$$n = 2.4/164 * 7 = 0.102$$

No. of moles of water (N) = Wt. of water/Mol. wt. of water = 100/18 = 5.55

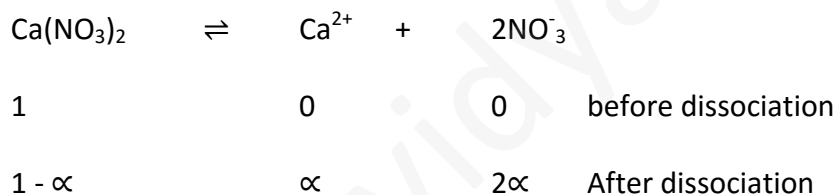
Applying Raoult's law, $P^0 - P/P^0 = n/n + N$

$$760 - p/760 = 0.102/0.102 + 5.55$$

$$760 - p/760 = 0.0180$$

$$\Rightarrow p = 760 - (760 * 0.0180) = 746.3 \text{ mm Hg}$$

ALTERNATIVE SOLUTION :



∴ Total moles at equilibrium = $(1 + 2\alpha)$

$$= 1 + 2 * 0.7 \quad (\because \alpha = 0.7)$$

$$= 2.4$$

For $\text{Ca}(\text{NO}_3)_2$: $m_{\text{ob}}/m_{\text{exp}} = 1 + 2\alpha$

$$\therefore m_{\text{exp}} = m_{\text{ob}}/1 + 2 * 0.7 = 164/2.4 = 68.33$$

Also at 100° $P_{\text{H}_2\text{O}}^0 = 760 \text{ mm}$, $w = 7\text{g}$

$$W = 100 \text{ g}$$

$$\text{Now, } P^{\circ} - P_s / P_s = 7 * 18 / 68.33 * 100 = 0.0184$$

$$\text{Or } P^{\circ} / P^s - 1 = 0.0184$$

$$\therefore P_s = 760 / 1.0184 = 74.26 \text{ mm}$$

Sol 12.

Give Wt. of benzene (solvent),

$$W = \text{Volume} * \text{density} = 50 * 0.789 = 43.95 \text{ g}$$

Wt. of compound (solute), $w = 0.643 \text{ g}$

Mol. wt of benzene, $M = 78$

Mol. wt. of solute, $m = ?$

Depression in freezing point, $\Delta T_f = 5.51 - 5.03 = 0.48$

Molal freezing constant, $K_f = 5.12$

Now we know that,

$$M - 1000 * K_f * w / W * \Delta T_f = 1000 * 5.12 * 0.643 / 43.95 * 0.48 = 156.056$$

Sol 13.

$$P^{\circ} - P / P^{\circ} = w / 3 / w / m + W / M$$

Here, w and m are wt. and molecular wt. of solute, W and M are wt. and molecular weight of solvent

p = pressure of solution;

p° = Normal vapour pressure

Let the initial (normal) pressure (p°) = p

$$\therefore \text{Pressure of solution} = 75/100 * p = 3/4 p$$

$M = 60, m = 118, W = 100 \text{ gm}$

$$\therefore p - 3/4 p / p = w / 60 / w / 60 + 100 / 18$$

$$1/4 = w/60 / (w/60) + 5.55$$

$$\text{Or } 4w/60 = w/60 + 5.55$$

$$3w/60 = w/20 = 5.55$$

$$\text{Or } w = 111 \text{ g}$$

$$\text{Molality} = \text{No. of moles of solute/Wt. of solvent} * 1000$$

$$= 111 * 1000 / 60 * 100 = \mathbf{18.52 \text{ m}}$$

Sol 14.

$$(i) \text{Volume} = \text{No. of moles} * \text{molar mass/density}$$

$$(ii) PV = nRT \text{ or } P = nRT/V$$

$$\text{Volume of 1 mole of liq. Benzene} = 78/0.877$$

$$\text{Volume of 1 mole of toluene. } 92/0.867$$

In vapour phase,

At 20°C, for 1 mole of benzene,

$$\text{Volume} = 1 * 78 * 2750 / 0.877 = 244583.80 \text{ mL}$$

$$= 244.58 \text{ L}$$

Similarly for 1 mole of toluene,

$$\text{Volume} = 1 * 92 / 0.867 * 7720$$

$$= 819192.61 \text{ mL} = 819.199 \text{ L}$$

As we know that, $PV = nrt$

$$\text{For benzene, } P_B^0 = nRT/V = 1 * 0.0821 * 293 / 244.58 \text{ atm}$$

$$= 0.098 \text{ atm}$$

$$\text{For toluene, } P_T^0 = nRT/V = 1 * 0.0825 * 293 / 819.19 \text{ atm} = 0.029 \text{ atm}$$

$$P = P_B^0 \cdot X_B + P_T^0 \cdot X_T$$

$$\therefore X_B + X_T = 1$$

$$\therefore X_T = 1 - X_B$$

$$P = P_B^0 \cdot X_B + P_T^0 (1 - X_B)$$

Total vapour pressure = 46 torr = 46/760 = 0.060 atm

$$\text{Thus, } 0.060 = 0.098 X_B + 0.029 (1 - X_B)$$

$$\Rightarrow 0.060 = 0.098 X_B + 0.029 - 0.029 X_B$$

$$\Rightarrow 0.031 = 0.069 X_B$$

$$\therefore X_B = 0.031/0.069 = 0.45 \text{ (in liquid phase)}$$

$$X_B + X_T = 1$$

$$X_T = 1 - 0.45 = 0.55 \text{ (in liquid phase)}$$

$$\text{Also, } P'_B = P_B^0 \cdot X_B = P X'_B$$

$$\text{So, } 0.098 * 0.045 = 0.060 * X'_B$$

$$X'_B = 0.098 * 0.45/0.060 = \mathbf{0.735 \text{ (in gas phase)}}$$

Sol 15.

According to Rault's law

$$P^0 - p/p^0 = w/m * M/W ; \Delta T_f = K_f * m$$

Substituting the given values;

$$1000 - 98.88/98.88 = w * 78 * 1000/m * W * 1000$$

$$w/m * 1000/W = 1.12 * 1000/78.98.88 = 0.1452$$

\therefore Molality = **0.1452** ($\because w/m * 1000/W = \text{Molality}$) Further $\Delta T = K_f \text{ molality}$

$$0.73 = K_f * 0.1452 \quad K_f = \mathbf{5.027 \text{ K molality}^{-1}}$$

Sol 16.

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

$$P^0 - p/p^0 = \text{moles of solute/moles of solvent}$$

Depression in freezing point, $\Delta T_f = K_f m$

$$\therefore m \Delta T_f / K_f = 0.30 / 1.86 = 0.161$$

According to Raoult's law

$$P^0 - p/p^0 = \text{No. of moles of solute/No. of moles of solvent}$$

$$23.51 - p/23.51 = 0.161/1000/18 = 0.161 * 18/1000$$

$$(\because \text{No. of moles of H}_2\text{O} = 1000/18)$$

On usual calculations,

$$23.51 - p/23.51 = 0.0020898$$

$$P = 23.51 - 23.51 * 0.0020898 = 23.51 - 0.68$$

$$p = 23.44 \text{ mm Hg}$$

Sol 17.

$$\Delta T_b = k_b * m$$

Element	%	Relative no. of atoms	Simplest ratio
C	42.86	$42.86/12 = 3.57$	$3.57/1.19 = 3$
H	2.40	$2.40/1 = 2.40$	$2.40/1.19 = 2$
N	16.67	$16.67/14 = 1.19$	$1.19/1.19 = 1$
O	38.07	$38.07/16 = 2.38$	$2.38/1.19 = 2$

\therefore Empirical formula of the minor product is $\text{C}_3\text{H}_2\text{NO}_2$ Molar empirical formula mass of the minor product = $3 * 12 + 2 * 1 + 14 + 2 * 16 = 84 \text{ g mol}^{-2}$

Let M be the molar mass of the minor product. For 5.5 g of the minor product dissolved in 45 g benzene, the molality

$$(m) \text{ of the solution} = 5.5 \text{ g}/M/0.045 \text{ kg}$$

Substituting this in the expression of elevation of boiling point,

$$\Delta T_b = k_b m \Rightarrow 1.84 \text{ K} = (2.53 \text{ K kg mol}^{-1}) (5.5 \text{ g}/M/0.045 \text{ kg})$$

$$\text{Or } M = 168 \text{ g mol}^{-1}$$

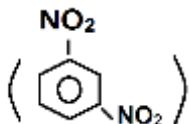
No. of unit of empirical formula in molecular formula

$$= 168 \text{ g mol}^{-1}/84 \text{ g} = 2$$

Hence the molecular formula of the minor product is



The product is *m* – dinitrobenzene



Sol 18.

$$\Delta T_f = i * k_f * m$$

$$\text{Weight of water} = 500 * 0.997 = 498.5 \text{ g (weight = volume * density)}$$

No. of moles of acetic acid

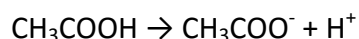
$$= \text{Wt. of CH}_3\text{COOH in g}/\text{Mol. wt. of CH}_3\text{COOH} = 3 * 10^{-3} * 10^3/60 = 0.05$$

Since 498.5 g of water has 0.05 moles of CH₃COOH

$$1000 \text{ g of water has} = 0.05 * 1000/498.5 = 0.1$$

Therefore molality of the solution = 0.1

Determination of van't Hoff factor, i



No. of moles at start	1	0	0
No. of moles at equb.	$1 - 0.23$	0.23	0.23

Therefore vant Hoff factor

= No. of particles before dissociation/No. of particles after dissociation

$$= 1 - 0.23 + 0.23 + 0.23/1 = 1.23$$

Now we know that

$$\Delta T_f = i * k_f * m = 1.23 = 1.86 * 0.1 = \mathbf{0.2228K}$$

ALTERNATIVE SOLUTION :

Density of water = 0.997 g/cm^3

Weight of water (W) = $500 * 0.997 = 498.5 \text{ g}$

Weight of acetic acid (w) = $3.0 * 10^{-3} \text{ kg} = 33.0 \text{ kg}$

$$\Delta T_f = 1000 * K_f * w/m * W$$

Given that K_f for water = $1.86 \text{ K kg}^{-1} \text{ mol}^{-1}$

mol. wt. of $\text{CH}_3\text{COOH}_{(m)} = 60$

$$(\Delta T_f)_{\text{cal}} = 1000 * 1.86 * 3.0/60 * 498.5 = 0.186$$

(Because CH_3COOH is an electrolyte and 23% dissociated)

	CH_3COOH	\rightleftharpoons	CH_3COO^-	+	H^+
At t = 0	1 mole		0		0
At equilibrium	$(1 - \alpha)$ mole		α gm-ion		α gm-ion

No. of particles after dissociation = $1 - \alpha + \alpha + \alpha = 1 + \alpha$

α for $\text{CH}_3\text{COOH} = 23/100 = 0.23$ (on 23% dissociation)

So, no. of particles after dissociation = $1 + 0.23 = 1.23$

By van't Hoff factor

$(\Delta T_f)_{\text{obs}}/(\Delta T_f)_{\text{cal}} = \text{No. of particles after dissociation/No. of particles before dissociation}$

$$(\Delta T_f)_{\text{obs}}/(\Delta T_f)_{\text{cal}} = 1.23/1$$

$$(\Delta T_f)_{\text{obs}} = 1.23 * 0.186 = 0.228 \text{ K}$$

Hence depression in freezing point $(\Delta T_f) = 0.228\text{K}$

Sol 19.

$$\Delta T_b = K_b * M$$

In first case,

$$\Delta T_b = K_b * m = K_b * \text{Wt. of solute/Mol. wt. of solute}$$

$$\text{Or } 0.17 = 1.7 * 1.22/M * 1000 * 10^{-3}$$

$$\text{Or } M = 122$$

Thus the benzoic acid exists as a monomer in acetone

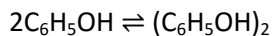
$$\text{(ii) } \Delta T_b = K_b * \text{Wt. of solute/Mol. wt. of solute}$$

$$\text{Or } 0.13 = 2.6 * 1.22/M' * 100 * 10^{-3}$$

$$M' = 224$$

NOTE : Double the expected molecular weight of benzoic acid (244) in acetone solution indicates that benzoic acid exists as a dimer in acetone.

Sol 20.



Initial no. of moles 1 0

No. of moles at equilibrium $1 - \alpha$ $\alpha/2$

Total number of moles at equilibrium = $1 - \alpha + \alpha/2 = 1 - \alpha/2$

$$\Delta T_f = iK_f * (\text{molality})$$

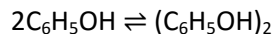
$$\Rightarrow 7 = 14 * 75.2/94 * (1 - \alpha/2) \text{ [weight of phenol} = 75.2\text{g mol. wt of phenol} = 94]$$

$$\therefore \alpha = 0.75$$

So the percentage of phenol that dimerises = **75%**.

ALTERNATIVE SOLUTION :

Phenol dimerises in the solvent (organic) as :



1 mol 0 mol (Initial)

$1 - \alpha$ mol $\alpha/2$ mol (Equilibrium)

Van't Hoff factor, $i = 1 - \alpha + \alpha/2 = 1 - \alpha/2$

Where α is the degree of dimerization.

$$\Delta T_{f(\text{obs})} = K_f * \text{molality} * i$$

$$\text{Or } 7 = 14 * m * (1 - \alpha/2)$$

NOTE : In the given problem, amount of the solvent used to dissolve 75.2 g of phenol is not mentioned. Hence molality m , can not be calculated and so also the value of α .

If amount of solution is presumed to be 1 kg (= 100g) then,

$$M = 75.2/94 = 0.8 \text{ (mol. mass of phenol = 94)}$$

$$\text{And } 7 = 14 * 75.2/94 * (1 - \alpha/2)$$

$$\text{Or } \alpha = 0.75 = 75\%$$