

Chapter.2 Solution Class-XII Subject-Chemistry

2.1 Define the term solution. How many types of solutions are formed? Write briefly about each type with an example.

Answer 2.1

It is a homogeneous mixture of two or more than two components.

There are three types of solutions:-

a. Gaseous solution:

A solutions in which solute may be liquid, solid, or gas but solvent is a gas, forms gaseous solutions. For example, a mixture of oxygen and nitrogen gas

b. Liquid solution:

A solution in which solute may be liquid, gas or solid but solvent is a liquid forms liquid solutions. For example, a solution of ethanol in water

c. Solid solution:

A solution in which solute may be liquid, gas or solid but solvent is a solid forms solid solutions. For example, a solution of copper in gold

2.2 Give an example of solid solution in which the solute is a gas.

Answer 2.2

Solute = gas

Solvent = solid

Example- A solution of hydrogen in palladium

2.3 Define the following terms:

- a) Mole fraction
- b) Molality
- c) Molarity
- d) Mass percentage

Answer 2.3

- a) **Mole fraction:** - It is defined as the ratio of the number of moles of the component to the total number of moles of all the components in the mixture.

$$\text{Mole fraction of a component} = \frac{\text{No. of moles of the component}}{\text{Total no. of moles of all the component}}$$

In a binary mixture, mole fraction of component will be

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

- b) **Molality:** - It can be defined as the number of moles of the solute per kilogram of the solvent.

$$\text{Molality, } m = \frac{\text{moles of solute}}{\text{mass of solvent in Kg}}$$

c)

Molarity: - it can be defined as the number of moles of the solute dissolved in one Litre of the solution.

$$\text{Molarity, } M = \frac{\text{moles of solute}}{\text{volume of solution in Litre}}$$

d)

Mass percentage: - Mass percentage of a component of a solution is defined as the mass of the solute in grams present in 100 g of the solution.

$$\text{mass \% of a component} = \frac{\text{mass of component in solution}}{\text{total mass of solution}} \times 100$$

2.4 Concentrated nitric acid used in laboratory work is 68% nitric acid by mass in aqueous solution. What should be the molarity of such a sample of the acid if the density of the solution is 1.504 g mL⁻¹?

Answer 2.4

68 g of nitric acid is dissolved in = 100 g of the solution

Molar mass of nitric acid (HNO₃) = 1 × 1 + 1 × 14 + 3 × 16

$$= 63 \text{ g/mol}$$

$$\text{Number of moles of HNO}_3 = \frac{68}{63} \text{ mol}$$

$$= 1.079 \text{ mol}$$

Density of solution = 1.504 g/mL

$$\text{Volume of 100 g solution} = \frac{100}{1.504} \text{ mL}$$

$$= 66.49 \times 10^{-3} \text{ L}$$

$$\text{Molarity of solution} = \frac{1.079}{66.49 \times 10^{-3}}$$

$$= 16.23 \text{ M}$$

2.5 A solution of glucose in water is labelled as 10% w/w, what would be the molality and mole fraction of each component in the solution? If the density of solution is 1.2 g mL⁻¹, then what shall be the molarity of the solution?

Answer 2.5

10 g of glucose is present in = 100 g of the solution

10 g of glucose is present in = (100 – 10) g

= 90 g of water

$$M_{C_6H_{12}O_6} = 180 \text{ g / mol}$$

$$n_{C_6H_{12}O_6} = \frac{10}{180} = 0.056 \text{ mol}$$

Thus,

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

$$n_{H_2O} = \frac{90}{18} = 5 \text{ mol}$$

$$\chi_{glucose} = \frac{0.056}{0.056 + 5} = 0.011$$

$$\begin{aligned}\chi_{water} &= 1 - \chi_{glucose} \\ &= 1 - 0.011 = 0.989\end{aligned}$$

$$\text{density} = 1.2 \text{ g / mL}$$

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

Now,

$$\text{Molarity of solution} = \frac{0.056}{83.33 \times 10^{-3}} = 0.67 \text{ M}$$

2.6 How many mL of 0.1 M HCl are required to react completely with 1 g mixture of Na₂CO₃ and NaHCO₃ containing equimolar amounts of both?

Answer 2.6

Let the amount of Na_2CO_3 in the solution = x g

Amount of NaHCO_3 in the solution = $(1 - x)$ g

Molar mass of $\text{Na}_2\text{CO}_3 = 2 \times 23 + 1 \times 12 + 3 \times 16$

$$= 106 \text{ g/mol}$$

$$\text{Number of moles } \text{Na}_2\text{CO}_3 = \frac{x}{106} \text{ mol}$$

Molar mass of $\text{NaHCO}_3 = 1 \times 23 + 1 \times 1 \times 12 + 3 \times 16$

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

$$\text{Number of moles of } \text{NaHCO}_3 = \frac{1 - x}{84} \text{ mol}$$

Now,

$$\frac{x}{106} = \frac{1 - x}{84}$$

$$84x = 106 - 106x$$

$$190x = 106$$

$$x = 0.5579$$

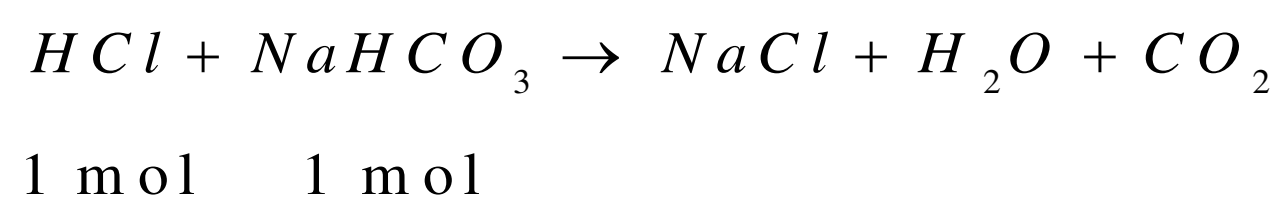
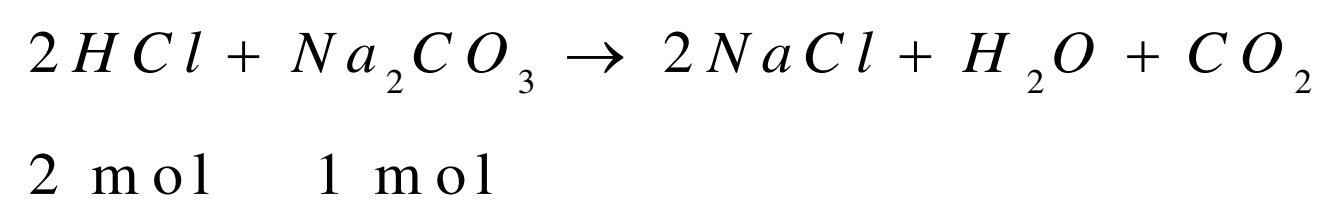
$$\text{Hence, number of moles of } \text{Na}_2\text{CO}_3 = \frac{0.5579}{106} \text{ mol}$$

$$= 0.0053 \text{ mol}$$

$$\text{And, number of moles of } \text{NaHCO}_3 = \frac{1 - 0.5579}{84} \text{ mol}$$

$$= 0.0053 \text{ mol}$$

As per the equations,



1 mol of Na_2CO_3 reacts with = 2 mol of HCl

So, 0.0053 mol of Na_2CO_3 reacts with = 2×0.0053 mol

= 0.0106 mol

Also,

1 mol of $NaHCO_3$ reacts with = 1 mol of HCl

0.0053 Mol of $NaHCO_3$ reacts with = 0.0053 mol of HCl

Total moles of HCl required = $(0.0106 + 0.0053) = 0.0159$ mol

0.1 Mol of HCl is present in = 1000 mL of the solution

Thus, 0.0159 mol of HCl is present in = $\frac{1000 \times 0.0159}{0.1}$ mol

= 159 mL

2.7 A solution is obtained by mixing 300 g of 25% solution and 400 g of 40% solution by mass. Calculate the mass percentage of the resulting solution.

Answer 2.7

$$\begin{aligned} \text{Total amount of solute present in the solution} &= \frac{300 \times 25}{100} + \frac{400 \times 40}{100} \\ &= 235 \text{ g} \end{aligned}$$

Total amount of solution = $300 + 400 = 700$ g

$$\text{Mass \% of the solute in the resulting solution} = \frac{235 \times 100\%}{700}$$

$$= 33.57\%$$

$$\text{Mass \% of the solvent in the resulting solution} = (100 - 33.57)\%$$

$$= 66.43\%$$

2.8 An antifreeze solution is prepared from 222.6 g of ethylene glycol ($\text{C}_2\text{H}_6\text{O}_2$) and 200 g of water. Calculate the molality of the solution. If the density of the solution is 1.072 g mL^{-1} , then what shall be the molarity of the solution?

Answer 2.8

$$\text{Density of the solution} = 1.072 \text{ g/mL}$$

$$\text{Volume of the solution} = \frac{422.6}{1.072} = 394.22 \text{ mL}$$

$$= 0.3942 \times 10^{-3} \text{ L}$$

$$M_{\text{ethylene glycol}} = 62 \text{ g/mol}$$

$$n_{\text{ethylene glycol}} = \frac{222.6}{62} = 3.59 \text{ mol}$$

$$\text{Therefore, molality of the solution} = \frac{3.59}{0.200} = 17.95 \text{ m}$$

$$\text{Total mass of the solution} = (222.6 + 200) \text{ g} = 422.6 \text{ g}$$

$$\text{Molarity of the solution} = \frac{3.59}{0.3942 \times 10^{-3}} = 9.11 \text{ M}$$

2.9 A sample of drinking water was found to be severely contaminated with chloroform (CHCl₃) supposed to be a carcinogen. The level of contamination was 15 ppm (by mass):

- i. Express this in percent by mass**
- ii. Determine the molality of chloroform in the water sample.**

Answer 2.9

- i.** Amount of chloroform = 15 ppm

$$\text{Hence, percent by mass} = \frac{15}{10^6} \times 100\% = 1.5 \times 10^{-5}\%$$

- ii.**

$$\begin{aligned}\text{Molar mass of chloroform} &= 1 \times 12 + 1 \times 1 + 3 \times 35.5 \\ &= 119.5 \text{ g/mol}\end{aligned}$$

Now,

15 g of chloroform is present in = 10^6 g of the solution

15 g of chloroform is present in = $(10^6 - 15) \approx 10^6$ g of water

Thus,

$$\begin{aligned}\text{Molality of the solution} &= \frac{15}{119.5 \times 10^6 \times 10^{-3}} \\ &= 1.26 \times 10^{-4} \text{ m}\end{aligned}$$

2.10 What role does the molecular interaction play in a solution of alcohol and water?

Answer 2.10

Strong hydrogen bonding is present in pure alcohol and water. Thus, in these solutions molecules are held together tightly. When both of these solutions are

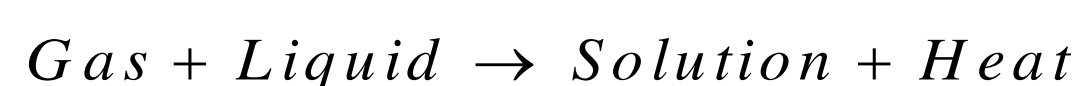
combined, new interactions are formed between the molecules of alcohol and water. These interactions are weaker than alcohol–alcohol and water–water interactions. As a result, the molecules can easily escape from the solution. This results in increases in the vapour pressure of the solution, which in turn lowers the boiling point of the resulting solution.

2.11 Why do gases always tend to be less soluble in liquids as the temperature is raised?

Answer 2.11

Solubility of gases in liquids decreases with an increase in temperature.

This happens because dissolution of gases in liquids is an exothermic process.



When we increase the temperature, heat is supplied and according to the given equation, equilibrium shifts backwards, & thus decreasing the solubility of gases.

2.12 State Henry's law and mention some important applications?

Answer 2.12

Henry's law: - It states that partial pressure of a gas in the vapour phase is proportional to the mole fraction of the gas in the solution.

$$p = K_H \cdot \chi$$

Where p = partial pressure of the gas in the vapour phase

χ = mole fraction of the gas

K_H = Henry's law constant

Applications of Henry's law are:

- To increase the solubility of CO₂ in soft drinks and soda water, bottles are sealed under high pressure.
- Oxygen tanks filled with air are used by scuba divers and are diluted with helium to avoid bends.
- The concentration of oxygen is low in the blood and tissues of people which live at high altitudes. This happens because at high altitudes, partial pressure of oxygen is less than that at ground level. Because of this reason, low-blood oxygen causes climbers to become weak and disables them from thinking clearly.

2.13 The partial pressure of ethane over a solution containing 6.56×10^{-3} g of ethane is 1 bar. If the solution contains 5.00×10^{-2} g of ethane, then what shall be the partial pressure of the gas?

Answer 2.13

$$\text{Molar mass of ethane} = 2 \times 12 + 6 \times 1$$

$$= 30 \text{ g/mol}$$

$$\text{No. of moles of ethane} = \frac{6.56 \times 10^{-2}}{30}$$

$$= 2.187 \times 10^{-3} \text{ mol}$$

Let the number of moles of the solvent = x

Now,

$$p = K_H \cdot \chi$$

$$1 = K_H \cdot \frac{2.187 \times 10^{-3}}{(2.187 \times 10^{-3}) + x}$$

$$1 = K_H \cdot \frac{2.187 \times 10^{-3}}{x}$$

$$K_H = \frac{x}{2.187 \times 10^{-3}}$$

(Since $x \gg 2.187 \times 10^{-3}$)

$$\begin{aligned}\text{Number of moles present in } 5.00 \times 10^{-2} \text{ g of ethane} &= \frac{5 \times 10^{-2}}{30} \text{ mol} \\ &= 1.67 \times 10^{-3} \text{ mol}\end{aligned}$$

Again,

$$\begin{aligned}p &= K_H \cdot \chi \\ &= \frac{x}{2.187 \times 10^{-3}} \times \frac{1.67 \times 10^{-3}}{(1.67 \times 10^{-3}) + x} \\ &= \frac{x}{2.187 \times 10^{-3}} \times \frac{1.67 \times 10^{-3}}{x} \\ &= 0.764 \text{ bar}\end{aligned}$$

Thus, partial pressure of the gas is 0.764 bar.

2.14 What is meant by positive and negative deviations from Raoult's law and how is the sign of $\Delta_{sol}H$ related to positive and negative deviations from Raoult's law?

Answer 2.14

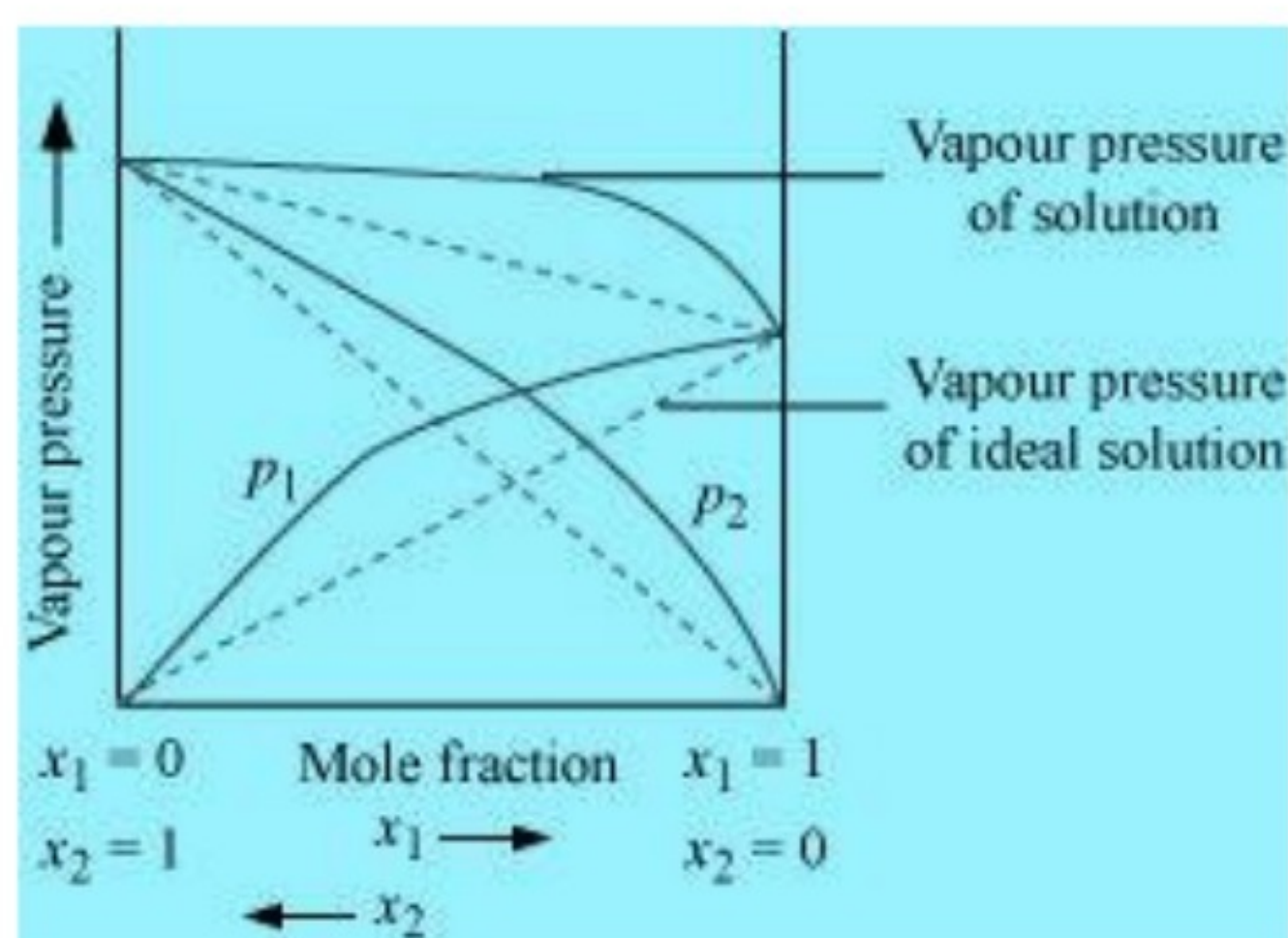
Raoult's law; - According to this law, the partial vapour pressure of each volatile component in any solution is directly proportional to its mole fraction. The solutions which obey Raoult's law over the entire range of concentration are known as ideal solutions. The solutions that do not obey Raoult's law have vapour pressures either higher or lower than that predicted by Raoult's law.

For ideal solution,

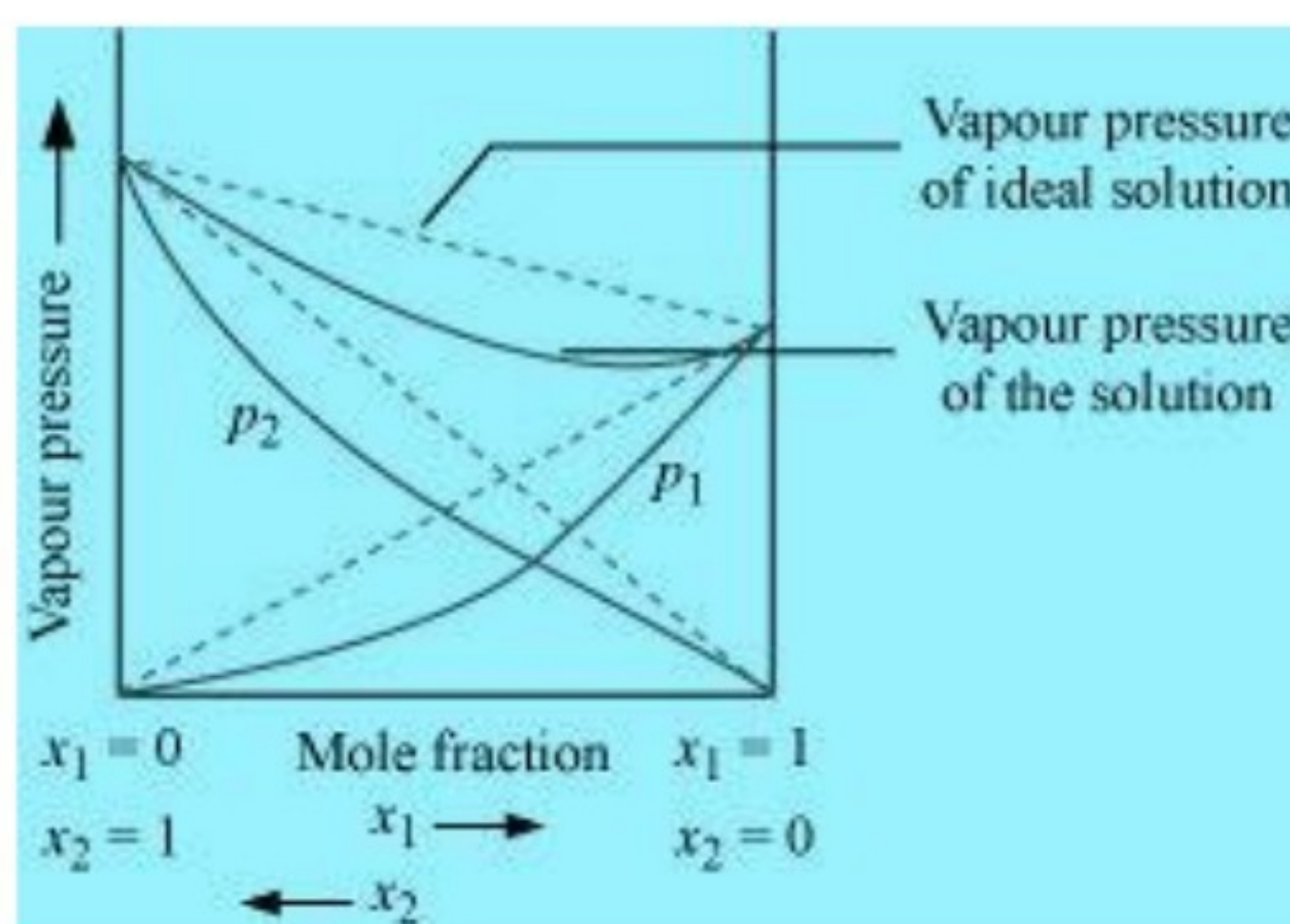
$$\Delta_{sol}H = 0$$

The solution which have higher vapour pressure show positive deviation and those which have lower vapour pressure show negative deviation from Raoult's law.

Positive deviation from Raoult's law



Negative deviation from Raoult's law



Solutions which show positive deviations, absorption of heat takes place. In such cases,

$$\Delta_{sol}H = \text{Positive}$$

Solutions which show negative deviations, evolution of heat takes place. In such cases,

$$\Delta_{sol}H = \text{Negative}$$

2.15 An aqueous solution of 2% non-volatile solute exerts a pressure of 1.004 bar at the normal boiling point of the solvent. What is the molar mass of the solute?

Answer 2.15

Vapour pressure of the solution, (p_1) = 1.004 bar

Vapour pressure of pure water, p_1^0 = 1.013 bar

Mass of solute, (w_2) = 2 g

Mass of solvent, (w_1) = 98 g

Molar mass of solvent, (M_1) = 18 g/mol

As per Raoult's law,

$$\frac{p_1^0 - p_1}{p_1^0} = \frac{w_2 \times M_1}{M_2 \times w_1}$$
$$\frac{1.013 - 1.004}{1.013} = \frac{2 \times 18}{M_2 \times 98}$$
$$\frac{0.009}{1.013} = \frac{36}{M_2 \times 98}$$

$$M_2 = 41.35 \text{ g / mol}$$

Thus, the molar mass of the solute is 41.35 g/mol.

2.16 Heptane and octane form an 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 g of octane?

Answer 2.16

Vapour pressure of heptane, p_1^0 = 105.2 kPa

Vapour pressure of octane, p_2^0 = 46.8 kPa

Molar mass of heptane (C_7H_{16}) = $7 \times 12 + 16 \times 1$

$$= 100 \text{ g/mol}$$

$$\text{Number of moles of heptane} = \frac{26}{100} \text{ mol}$$

$$= 0.26 \text{ mol}$$

$$\text{Molar mass of octane (C}_8\text{H}_{18}) = 8 \times 12 + 18 \times 1$$

$$= 114 \text{ g/mol}$$

$$\text{Number of moles of octane} = \frac{35}{114} \text{ mol}$$

$$= 0.31 \text{ mol}$$

$$\text{Mole fraction of heptane, } \chi_1 = \frac{0.26}{0.26 + 0.31}$$

$$= 0.456$$

$$\text{And, mole fraction of octane, } \chi_2 = 1 - 0.456$$

$$= 0.544$$

$$\text{Partial pressure of heptane, } p_1 = \chi_1 p_1^0$$

$$= 0.456 \times 105.2 = 47.97 \text{ kPa}$$

$$\text{Partial pressure of octane, } p_2 = \chi_2 p_2^0$$

$$= 0.544 \times 46.8 = 25.46 \text{ kPa}$$

Therefore,

$$\text{Vapour pressure of solution, } p_{\text{total}} = p_1 + p_2$$

$$= 47.97 + 25.46 = 73.43 \text{ kPa}$$

2.17 The vapour pressure of water is 12.3 kPa at 300 K. Calculate vapour pressure of 1 molal solution of a non-volatile solute in it.

Answer 2.17

1 mol of the solute is present in = 1000 g of the water

Molar mass of water = 18 g/mol

$$\text{No. of moles present in 1000 g of water} = \frac{1000}{18}$$

$$= 55.56 \text{ mol}$$

Therefore, mole fraction of the solute in the solution is

$$\chi_2 = \frac{1}{1 + 55.56} = 0.0177$$

We know that,

Vapour pressure of water, $p_1^0 = 12.3 \text{ kPa}$

Now,

$$\frac{p_1^0 - p_1}{p_1^0} = \chi_2$$

$$\frac{12.3 - p_1}{12.3} = 0.0177$$

$$12.3 - p_1 = 0.2177$$

$$p_1 = 12.0823$$

$$= 12.08 \text{ kPa}$$

Hence, the vapour pressure of the solution is 12.08 kPa.

2.18 Calculate the mass of a non-volatile solute (molar mass 40 g/mol) which should be dissolved in 114 g octane to reduce its vapour pressure to 80%.

Answer 2.15

Let the vapour pressure of pure octane = p_1^0

Then, the vapour pressure of the resulting solution is $= \frac{80 p_1^0}{100} = 0.8 p_1^0$

Molar mass of solute, $M_2 = 40 \text{ g/mol}$

Mass of octane, $w_1 = 114 \text{ g}$

Molar mass of octane, $M_1 = 8 \times 12 + 18 \times 1$

$$= 114 \text{ g/mol}$$

Now,

$$\frac{p_1^0 - p_1}{p_1^0} = \frac{w_2 \times M_1}{M_2 \times w_1}$$

$$\frac{p_1^0 - 0.8 p_1^0}{p_1^0} = \frac{w_2 \times 114}{40 \times 114}$$

$$\frac{0.2 p_1^0}{p_1^0} = \frac{w_2}{40}$$

$$w_2 = 8 \text{ g}$$

Thus, the mass of the solute is 8 g.

2.19 A solution containing 30 g of non-volatile solute exactly in 90 g of water has a vapour pressure of 2.8 kPa at 298 K. Further, 18 g of water is then added to the solution and the new vapour pressure becomes 2.9 kPa at 298 K. Calculate:

- molar mass of the solute
- Vapour pressure of water at 298 K.

Answer 2.19

- Let molar mass of the solute = M

$$\text{No. of moles of solvent, } n_1 = \frac{90}{18} = 5 \text{ mol}$$

$$\text{No. of moles of solute, } n_2 = \frac{30}{M} \text{ mol}$$

$$p_1 = 2.8 \text{ kPa}$$

Now,

$$\frac{p_1^0 - p_1}{p_1^0} = \frac{n_2}{n_1 + n_2}$$

$$\frac{p_1^0 - 2.8}{p_1^0} = \frac{\frac{30}{M}}{5 + \frac{30}{M}}$$

$$1 - \frac{2.8}{p_1^0} = \frac{\frac{30}{M}}{\frac{5M + 30}{M}}$$

$$1 - \frac{2.8}{p_1^0} = \frac{30}{5M + 30}$$

$$1 - \frac{30}{5M + 30} = \frac{2.8}{p_1^0}$$

$$\frac{5M + 30 - 30}{5M + 30} = \frac{2.8}{p_1^0}$$

$$\frac{5M}{5M + 30} = \frac{2.8}{p_1^0}$$

$$\frac{p_1^0}{2.8} = \frac{5M + 30}{5M} \dots\dots\dots(1)$$

After the addition of 18 g of water,

$$n_1 = \frac{90 + 18}{18} = 6 \text{ mol}$$

$$p_1 = 2.9 \text{ kPa}$$

Again,

$$\frac{p_1^0 - p_1}{p_1^0} = \frac{n_2}{n_1 + n_2}$$

$$\frac{p_1^0 - 2.9}{p_1^0} = \frac{\frac{30}{M}}{6 + \frac{30}{M}}$$

$$1 - \frac{2.9}{p_1^0} = \frac{\frac{30}{M}}{\frac{6M + 30}{M}}$$

$$1 - \frac{2.9}{p_1^0} = \frac{30}{6M + 30}$$

$$\frac{2.9}{p_1^0} = 1 - \frac{30}{6M + 30}$$

$$\frac{2.9}{p_1^0} = \frac{6M}{6M + 30}$$

$$\frac{p_1^0}{2.9} = \frac{6M + 30}{6M} \dots\dots\dots (2)$$

Dividing equations (1) by (2),

$$\frac{2.9}{2.8} = \frac{5M + 30}{6M + 30} \times \frac{6}{5}$$

$$2.9 \times 5 \times (6M + 30) = 2.8 \times 6 \times (5M + 30)$$

$$87M + 435 = 84M + 504$$

$$3M = 69$$

$$M = 23$$

Thus, the molar mass of the solute = 23 g/mol

ii. Putting the value of 'M' in equation (1),

$$\frac{p_1^0}{2.8} = \frac{5 \times 23 + 30}{5 \times 23}$$

$$\frac{p_1^0}{2.8} = \frac{145}{115}$$

$$p_1^0 = 3.53$$

Hence, the vapour pressure of water = 3.53 kPa

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

$$\begin{aligned}\Delta T_f &= (273.15 - 271) \text{ K} \\ &= 2.15 \text{ K}\end{aligned}$$

$$M_{C_{12}H_{22}O_{11}} = 342 \text{ g / mol}$$

5 g of cane sugar is present in = (100 – 5)g = 95 g of water

$$\begin{aligned}n_{\text{cane sugar}} &= \frac{5}{342} \\ &= 0.0146 \text{ mol}\end{aligned}$$

$$\begin{aligned}\text{Molality of solution, } m &= \frac{0.0146}{0.095} \\ &= 0.1537 \text{ mol / kg}\end{aligned}$$

Now,

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

$$\begin{aligned}K_f &= \frac{\Delta T_f}{m} \\ &= \frac{2.15}{0.1537}\end{aligned}$$

$$= 13.99 \text{ K kg mol}^{-1}$$

$$M_{\text{glucose}} = 180 \text{ g / mol}$$

5% glucose in water means 5 g of glucose is present in $(100 - 5) \text{ g} = 95 \text{ g}$ of water.

$$n_{\text{glucose}} = \frac{5}{180} = 0.0278 \text{ mol}$$

$$\text{Molality of the solution, } m = \frac{0.0278}{0.095}$$

$$= 0.2926 \text{ mol/kg}$$

Now,

$$\begin{aligned}\Delta T_f &= K_f \cdot m \\ &= 13.99 \times 0.2926 \\ &= 4.09 \text{ K}\end{aligned}$$

Hence, the freezing point of 5% glucose solution = $(273.15 - 4.09) \text{ K}$

$$= 269.06 \text{ K}$$

2.21 Two elements A and B form compounds having formula AB_2 and AB_4 . When dissolved in 20 g of benzene (C_6H_6), 1 g of AB_2 lowers the freezing point by 2.3 K whereas 1.0 g of AB_4 lowers it by 1.3 K. The molar depression constant for benzene is $5.1 \text{ K kg mol}^{-1}$. Calculate atomic masses of A and B.

Answer 2.21

$$M_2 = \frac{1000 \times w_2 \times k_f}{\Delta T_f \times w_1}$$

$$M_{\text{AB}_2} = \frac{1000 \times 1 \times 5.1}{2.3 \times 20}$$

$$= 110.87 \text{ g/mol}$$

$$M_{\text{AB}_4} = \frac{1000 \times 1 \times 5.1}{1.3 \times 20}$$

$$= 196.15 \text{ g/mol}$$

Let the atomic mass of A = x

And, atomic mass of B = y

Now, we can say

$$x + 2y = 110.87$$

$$x + 4y = 196.15$$

On solving above equations, we get

$$x = 25.59, \quad y = 42.64$$

Hence, the atomic masses of A and B are 25.59 u and 42.64 u respectively.

2.22 At 300 K, 36 g of glucose present in a litre of its solution has an osmotic pressure of 4.98 bar. If the osmotic pressure of the solution is 1.52 bars at the same temperature, what would be its concentration?

Answer 2.22

$$\pi = 1.52 \text{ bar}$$

$$T = 300 \text{ K}$$

$$R = 0.083 \text{ bar L/ K mol}$$

We know that,

$$\pi = CRT$$

$$C = \frac{\pi}{RT}$$

$$= \frac{1.52}{0.083 \times 300} = 0.061 \text{ mol}$$

Volume of the solution = 1 L

Therefore, concentration of the solution = 0.061 M

2.23 Suggest the most important type of intermolecular attractive interaction in the following pairs.

- i. n-hexane and n-octane
- ii. I_2 and CCl_4
- iii. $NaClO_4$ and water
- iv. Methanol and acetone
- v. acetonitrile (CH_3CN) and acetone (C_3H_6O)

Answer 2.23

- i. Vander Wall's forces of attraction
- ii. Vander Wall's forces of attraction
- iii. Ion-dipole interaction
- iv. Dipole-dipole interaction
- v. Dipole-dipole interaction

2.24 Based on solute-solvent interactions, arrange the following in order of increasing solubility in n-octane and explain. Cyclohexane, KCl, CH_3OH , CH_3CN

Answer 2.24

n-octane is a non-polar solvent. Thus, the solubility of a non-polar solute is more than that of a polar solute in the *n*-octane.

Order of increasing solubility is:



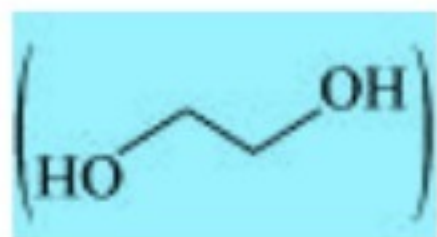
2.25 Amongst the following compounds, identify which are insoluble, partially soluble and highly soluble in water?

- i. Phenol
- ii. Toluene
- iii. Formic acid
- iv. Ethylene glycol
- v. Chloroform

vi. Pentanol

Answer 2.25

- i. Molecular formula of phenol is C_6H_5OH which has a polar group $-OH$ and non-polar group $-C_6H_5$. Thus, it is partially soluble in water.
- ii. Molecular formula of toluene is $C_6H_5-CH_3$ that means it doesn't has polar groups. Thus, it is insoluble in water.
- iii. Molecular formula of formic acid is $HCOOH$ that means it has the polar group $-OH$ which can form Hydrogen-bond with water. Thus, it is highly soluble in water.
- iv. Molecular formula of ethylene glycol is



that means it has polar $-OH$ group which can form Hydrogen-bond. Thus, it is highly soluble in water.

- v. Molecular formula of chloroform is $CHCl_3$ that means it doesn't have any polar group. Thus, it is insoluble in water.
- vi. Molecular formula of Pentanol is $C_5H_{11}OH$ that means it has polar $-OH$ group, & a very bulky non-polar $-C_5H_{11}$ group. Thus, it is partially soluble in water.

2.26 If the density of some lake water is 1.25 g mL^{-1} and contains 92 g of Na^+ ions per kg of water, calculate the molality of Na^+ ions in the lake.

Answer 2.26

$$\text{Number of moles present in 92 g of } Na^+ \text{ ions} = \frac{92}{23} = 4 \text{ mol}$$

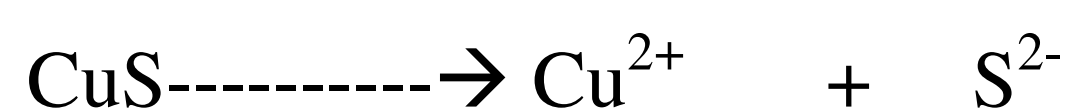
$$\text{Therefore, molality of } Na^+ \text{ ions in the lake} = \frac{4}{1} = 4 \text{ m}$$

2.27 If the solubility product of CuS is 6×10^{-16} , calculate the maximum molarity of CuS in aqueous solution.

Answer 2.27

$$\text{Solubility product of } CuS, K_{sp} = 6 \times 10^{-16}$$

Let s = the solubility of CuS



$$K_{sp} = [Cu^{2+}][S^{2-}]$$

$$= s \times s$$

$$= s^2$$

$$K_{sp} = s^2$$

$$s^2 = 6 \times 10^{-16}$$

$$s = \sqrt{6 \times 10^{-16}}$$

$$= 2.45 \times 10^{-8} \text{ mol / L}$$

Hence, the molarity of CuS in an aqueous solution is $2.45 \times 10^{-8} \text{ mol / L}$.

2.28 Calculate the mass percentage of aspirin ($C_9H_8O_4$) in acetonitrile (CH_3CN) when 6.5 g of $C_9H_8O_4$ is dissolved in 450 g of CH_3C

Answer 2.28

6.5 g of $C_9H_8O_4$ is dissolved in = 450 g of CH_3CN

Total mass of the solution = (6.5 + 450) g

$$= 456.5 \text{ g}$$

$$\text{Hence, mass percentage of } C_9H_8O_4 = \frac{6.5}{456.5} \times 100\%$$

$$= 1.424\%$$

2.29 Nalorphene ($C_{19}H_{21}NO_3$), similar to morphine, is used to combat withdrawal symptoms in narcotic users. Dose of nalorphene generally given is 1.5 mg. Calculate the mass of $1.5 \times 10^{-3} \text{ m}$ aqueous solution required for the above dose.

Answer 2.29

Molar mass of nalorphene = 311 g/mol

In aqueous solution of nalorphene,

One kg of water contains = 1.5×10^{-3} mol

$$= 1.5 \times 10^{-3} \times 311 \text{ g}$$

$$= 0.4665 \text{ g}$$

Thus, total mass of the solution = $(1000 + 0.4665) = 1000.4665 \text{ g}$

Now,

$$\begin{aligned} \text{Mass of the solution containing 1.5 mg of nalorphene is:} &= \frac{1000.4665 \times 1.5 \times 10^{-3}}{0.4665} \text{ g} \\ &= 3.22 \text{ g} \end{aligned}$$

Hence, the mass of aqueous solution is 3.22 g.

2.30 Calculate the amount of benzoic acid ($\text{C}_6\text{H}_5\text{COOH}$) required for preparing 250 mL of 0.15 M solution in methanol.

Answer 2.30

1000 mL of solution contains = 0.15 mol benzoic acid

$$250 \text{ mL of solution contains} = \frac{0.15 \times 250}{1000} \text{ mol of benzoic acid}$$

$$= 0.0375 \text{ mol of benzoic acid}$$

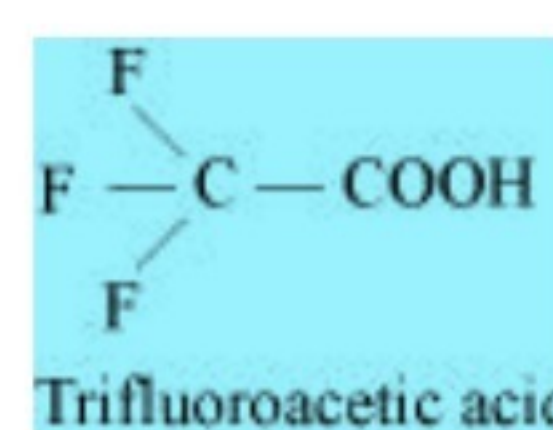
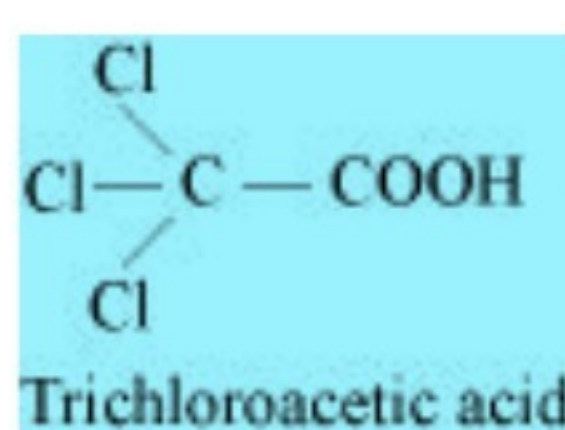
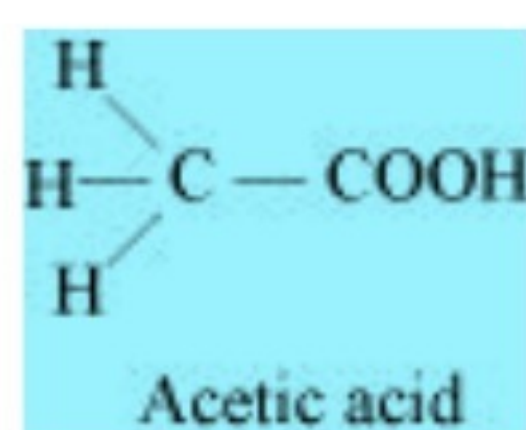
$$M_{\text{C}_6\text{H}_5\text{COOH}} = 122 \text{ g/mol}$$

Hence, required benzoic acid = $0.0375 \text{ mol} \times 122 \text{ g/mol}$

$$= 4.575 \text{ g}$$

2.31 The depression in freezing point of water observed for the same amount of acetic acid, trichloroacetic acid and trifluoroacetic acid increases in the order given above. Explain briefly.

Answer 2.31



Hydrogen is least electronegative & Fluorine is most electronegative atom. Thus, F can withdraw electrons towards itself more than Cl and H. Hence, trifluoroacetic acid can easily lose H^+ ions. Thus, trifluoroacetic acid ionizes to the largest extent. The acid which produces more ions will have great depression in the freezing point. Thus, the depression in the freezing point increases in the following order:

Acetic acid < trichloroacetic acid < trifluoroacetic acid

2.32 Calculate the depression in the freezing point of water when 10 g of $\text{CH}_3\text{CH}_2\text{CHClCOOH}$ is added to 250 g of water. $K_a = 1.4 \times 10^{-3}$, $K_f = 1.86 \text{ K kg mol}^{-1}$.

Answer 2.32

Molar mass of $\text{CH}_3\text{CH}_2\text{CHClCOOH} = 122.5 \text{ g/mol}$

$$\text{No. of moles present in 10 g of } \text{CH}_3\text{CH}_2\text{CHClCOOH} = \frac{10}{122.5} = 0.0816 \text{ mol}$$

$$\begin{aligned} \text{Molality of the solution} &= \frac{0.0816}{0.250} \times 1000 \\ &= 0.3264 \text{ mol/kg} \end{aligned}$$

Let α = degree of dissociation of $CH_3CH_2CHClCOOH$



Initial conc.	C	0	0
At equilibrium	$C(1 - \alpha)$	$C\alpha$	$C\alpha$

$$K_a = \frac{C\alpha \cdot C\alpha}{C(1 - \alpha)}$$

$$= \frac{C\alpha^2}{1 - \alpha}$$

Since α is very small, $1 - \alpha \approx 1$

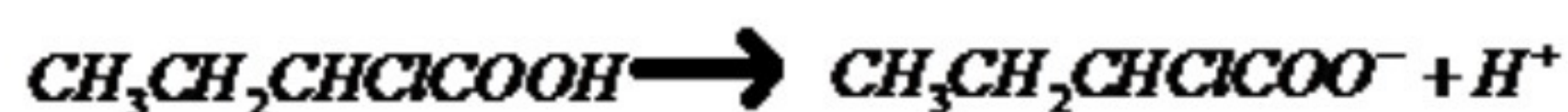
$$K_a = \frac{C\alpha^2}{1} = C\alpha^2$$

$$\alpha = \sqrt{\frac{K_a}{C}}$$

$$K_a = 1.4 \times 10^{-3}$$

$$= \sqrt{\frac{1.4 \times 10^{-3}}{0.3264}} = 0.0655$$

Again,



Initial moles	1	0	0
At equilibrium	$1 - \alpha$	α	α

Total moles of equilibrium = $1 - \alpha + \alpha + \alpha = 1 + \alpha$

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

$$\begin{aligned} &= 1 + \alpha \\ &= 1 + 0.0655 \\ &= 1.0655 \end{aligned}$$

Thus, we can calculate the depression in the freezing point of water

$$\begin{aligned} \Delta T_f &= i \cdot K_f \cdot m \\ &= 1.0655 \times 1.86 \times 0.3264 \\ &= 0.65 \text{ K} \end{aligned}$$

2.33 19.5 g of CH_2FCOOH is dissolved in 500 g of water. The depression in the freezing point of water observed is 1.0°C . Calculate the van't Hoff factor and dissociation constant of fluoroacetic acid.

Answer 2.33

$$\begin{aligned} w_1 &= 500 \text{ g} \\ w_2 &= 19.5 \text{ g} \\ \Delta T_f &= 1 \text{ K} \\ K_f &= 1.86 \text{ K Kg / mol} \end{aligned}$$

Now,

$$\begin{aligned} M_2 &= \frac{K_f \times w_2 \times 1000}{\Delta T_f \times w_1} \\ &= \frac{1.86 \times 19.5 \times 1000}{500 \times 1} = 72.54 \text{ g / mol} \end{aligned}$$

$$(M_2)_{\text{observed}} = 72.54 \text{ g / mol}$$

$$(M_2)_{\text{calculated}} = 78 \text{ g / mol}$$

$$\text{Van't Hoff factor, } i = \frac{(M_2)_{\text{calculated}}}{(M_2)_{\text{observed}}} = \frac{78}{72.54} = 1.0753$$

Let α = degree of dissociation of CH_2FCOOH



Initial conc.	C	0	0
At equilibrium	$C(1-\alpha)$	$C\alpha$	$C\alpha$

$$i = \frac{C(1+\alpha)}{C}$$

$$i = 1 + \alpha$$

$$\alpha = i - 1$$

$$= 1.0753 - 1 = 0.0753$$

Now,

$$\begin{aligned}
 K_a &= \frac{[CH_2FCOO^-][H^+]}{[CH_2FCOOH]} \\
 &= \frac{C\alpha \cdot C\alpha}{C(1-\alpha)} \\
 &= \frac{C\alpha^2}{1-\alpha}
 \end{aligned}$$

Volume = 500 mL

$$\begin{aligned}
 C &= \frac{19.5}{78 \times 500} \times 1000 \text{ M} \\
 &= 0.5 \text{ M}
 \end{aligned}$$

Thus,

$$\begin{aligned}
 K_a &= \frac{C\alpha^2}{1-\alpha} \\
 &= \frac{0.5 \times (0.0753)^2}{0.9247} = 3.07 \times 10^{-3}
 \end{aligned}$$

2.34 Vapour pressure of water at 293 K is 17.535 mm Hg. Calculate the vapour pressure of water at 293 K when 25 g of glucose is dissolved in 450 g of water.

Answer 2.34

Vapour pressure of water, $p_1^0 = 17.535 \text{ mm of Hg}$

Mass of glucose, $w_2 = 25 \text{ g}$

Mass of water, $w_1 = 450 \text{ g}$

Molar mass of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$), $M_2 = 180 \text{ g/mol}$

Molar mass of water, $M_1 = 18 \text{ g/mol}$

$$\therefore n_2 = \frac{25}{180} = 0.139 \text{ mol}$$

$$\therefore n_1 = \frac{450}{18} = 25 \text{ mol}$$

Now,

$$\frac{p_1^0 - p_1}{p_1^0} = \frac{n_2}{n_1 + n_2}$$
$$\frac{17.535 - p_1}{17.535} = \frac{0.139}{0.139 + 25}$$
$$p_1 = 17.44 \text{ mm of Hg}$$

Hence, the vapour pressure of water is 17.44 mm of Hg.

2.35 Henry's law constant for the molality of methane in benzene at 298 K is $4.27 \times 10^5 \text{ mm Hg}$. Calculate the solubility of methane in benzene at 298 K under 760 mm Hg.

Answer 2.35

$$p = 760 \text{ mm Hg}$$

$$K_H = 4.27 \times 10^5 \text{ mm Hg}$$

According to Henry's law,

$$p = K_H \cdot \chi$$

$$\chi = \frac{p}{K_H}$$

$$= \frac{760}{4.27 \times 10^5} = 178 \times 10^{-5}$$

Hence, the mole fraction of methane in benzene is 178×10^{-5} .

2.36 100 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} .

Answer 2.36

$$n_A = \frac{100}{140} \text{ mol} = 0.714 \text{ mol}$$

$$n_B = \frac{1000}{180} \text{ mol} = 5.556 \text{ mol}$$

$$\begin{aligned}\chi_A &= \frac{n_A}{n_A + n_B} \\ &= \frac{0.714}{0.714 + 5.556} = 0.114\end{aligned}$$

$$\begin{aligned}\chi_B &= 1 - \chi_A \\ &= 1 - 0.114 = 0.886\end{aligned}$$

Now,

$$p_B^0 = 500 \text{ torr}$$

$$\begin{aligned}p_B &= p_B^0 \cdot \chi_B \\ &= 500 \times 0.886 = 443 \text{ torr}\end{aligned}$$

$$p_{\text{total}} = 475 \text{ torr}$$

$$\begin{aligned}p_A &= p_{\text{total}} - p_B \\ &= 475 - 443 = 32 \text{ torr}\end{aligned}$$

Now, we can calculate p_A^0

$$\begin{aligned}
 p_A &= p_A^0 \cdot \chi_A \\
 p_A^0 &= \frac{p_A}{\chi_A} \\
 &= \frac{32}{0.114} = 280.7 \text{ torr}
 \end{aligned}$$

Hence, the vapour pressure of pure liquid A is 280.7 torr .

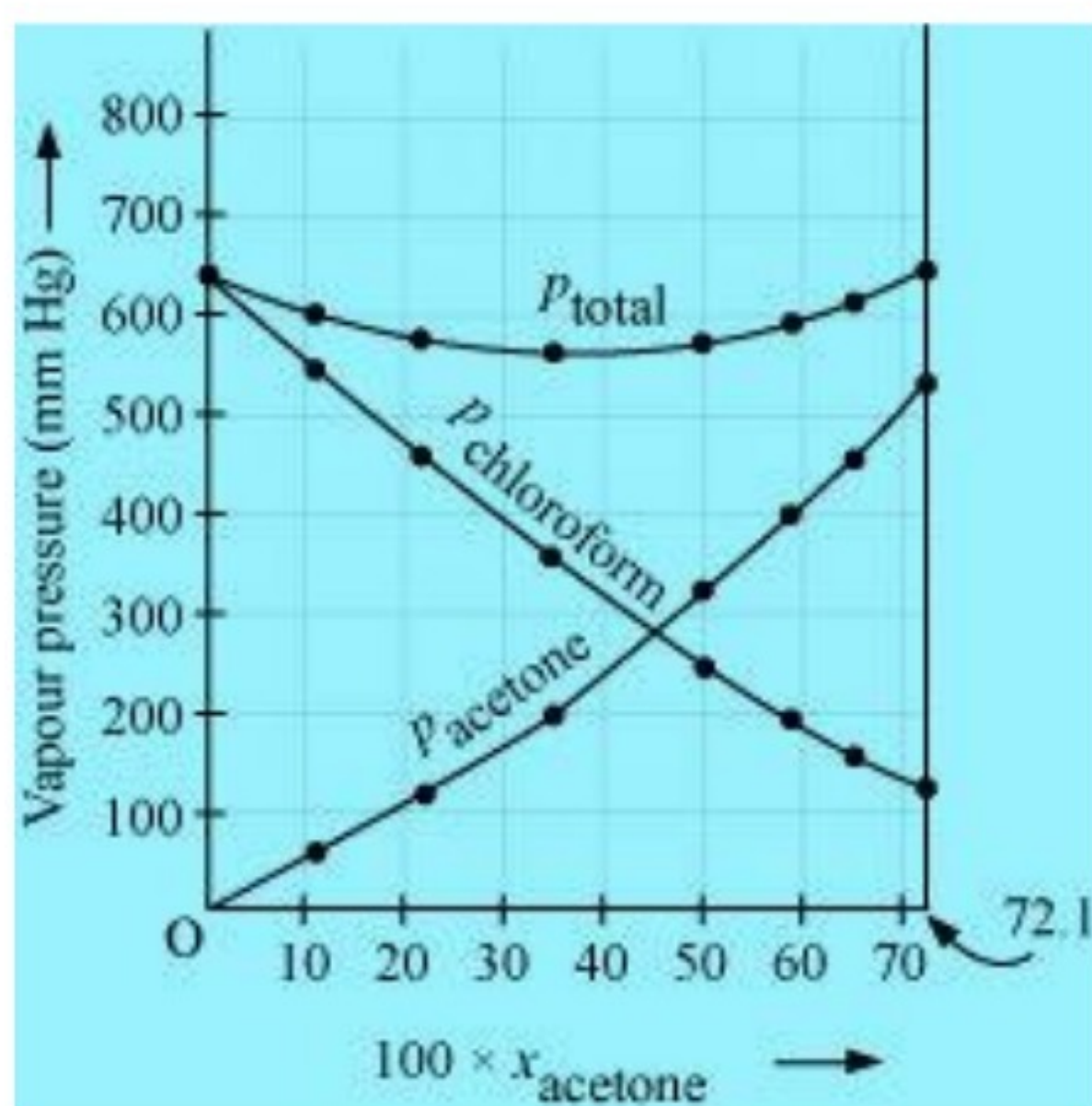
2.37 Vapour pressure of pure acetone and chloroform at 328 K are 741.8 mm Hg and 632.8 mm Hg respectively. Assuming that they form ideal solution over the entire range of composition, plot p_{total} , $p_{\text{chloroform}}$ and p_{acetone} as a function of x_{acetone} . The experimental data observed for different compositions of mixture is.

$100 \times x_{\text{acetone}}$	0	11.8	23.4	36.0	50.8	58.2	64.5	72.1
$p_{\text{acetone}}/\text{mm Hg}$	0	54.9	110.1	202.4	322.7	405.9	454.1	521.1
$p_{\text{chloroform}}/\text{mm Hg}$	632.8	548.1	469.4	359.7	257.7	193.6	161.2	120.7

Answer 2.37

We have the following data:

$100 \times x_{\text{acetone}}$	0	11.8	23.4	36.0	50.8	58.2	64.5	72.1
$p_{\text{acetone}}/\text{mm Hg}$	0	54.9	110.1	202.4	322.7	405.9	454.1	521.1
$p_{\text{chloroform}}/\text{mm Hg}$	632.8	548.1	469.4	359.7	257.7	193.6	161.2	120.7
$p_{\text{total}}(\text{mm Hg})$	632.8	603.0	579.5	562.1	580.4	599.5	615.3	641.8



From the graph, we can conclude that the solution shows negative deviation from the ideal behaviour.

2.38 Benzene and toluene form ideal solution over the entire range of composition.

The vapour pressure of pure benzene and naphthalene at 300 K are 50.71 mm Hg and 32.06 mm Hg respectively. Calculate the mole fraction of benzene in vapour phase if 80 g of benzene is mixed with 100 g of toluene.

Answer 2.38

$$M_{C_6H_6} = 78 \text{ g / mol}$$

$$M_{C_6H_5CH_3} = 92 \text{ g / mol}$$

$$\text{Now, no. of moles present in 80 g of benzene} = \frac{80}{78} = 1.026 \text{ mol}$$

$$\text{And, no. of moles present in 100 g of toluene} = \frac{100}{92} = 1.087 \text{ mol}$$

$$\chi_{\text{benzene}} = \frac{1.026}{1.026 + 1.087} = 0.486$$

$$\chi_{\text{toluene}} = 1 - 0.486 = 0.514$$

$$p_b^0 = 50.71 \text{ mm of Hg}$$

$$p_t^0 = 32.06 \text{ mm of Hg}$$

Now,

$$\begin{aligned}
 p_b &= \chi_b \cdot p_b^0 \\
 &= 0.486 \times 50.71 = 24.645 \text{ mm of Hg}
 \end{aligned}$$

$$\begin{aligned}
 p_t &= \chi_t \cdot p_t^0 \\
 &= 0.514 \times 32.06 = 16.479 \text{ mm of Hg}
 \end{aligned}$$

$$\begin{aligned}
 \text{Therefore, } \chi_{benzene} \text{ in vapour phase} &= \frac{p_b}{p_b + p_t} \\
 &= \frac{24.645}{24.645 + 16.479} = \frac{24.645}{41.124} = 0.6
 \end{aligned}$$

2.39 The air is a mixture of a number of gases. The major components are oxygen and nitrogen with approximate proportion of 20% is to 79% by volume at 298 K. The water is in equilibrium with air at a pressure of 10 atm. At 298 K. if the Henry's law constants 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.

Answer 2.39

% of O₂ in air = 20 %

% of N₂ in air = 79%

Also, we know that water is in equilibrium with air at a pressure of 10 atm = 7600 mm Hg

Thus,

$$\begin{aligned}
 p_{O_2} &= \frac{20}{100} \times 7600 \text{ mm of Hg} \\
 &= 1520 \text{ mm of Hg}
 \end{aligned}$$

$$\begin{aligned}
 p_{N_2} &= \frac{79}{100} \times 7600 \text{ mm of Hg} \\
 &= 6004 \text{ mm of Hg}
 \end{aligned}$$

$$p = K_H \cdot \chi$$

$$p_{O_2} = K_H \cdot \chi_{O_2}$$

$$K_H = 3.3 \times 10^7 \text{ mm of Hg}$$

$$\chi_{O_2} = \frac{p_{O_2}}{K_H} = \frac{1520}{3.3 \times 10^7} = 4.61 \times 10^{-5}$$

$$p_{N_2} = K_H \cdot \chi_{N_2}$$

$$K_H = 6.51 \times 10^7$$

$$\chi_{N_2} = \frac{p_{N_2}}{K_H} = \frac{6004}{6.51 \times 10^7} = 9.22 \times 10^{-5}$$

Thus, the mole fractions of oxygen and nitrogen in water are 4.61×10^{-5} and 9.22×10^{-5} respectively.

2.40 Determine the amount of CaCl_2 ($i = 2.47$) dissolved in 2.5 litre of water such that its osmotic pressure is 0.75 atm at 27°C .

Answer 2.40

$$\pi = 0.75 \text{ atm}$$

$$V = 2.5 \text{ L}$$

$$i = 2.47$$

$$T = 300 \text{ K}$$

$$R = 0.0821 \text{ L atm / K mol}$$

$$M = 111 \text{ g / mol}$$

Now,

$$\pi = \frac{i n R T}{V} = \frac{i w R T}{M V}$$

$$w = \frac{\pi M V}{i R T}$$

$$= \frac{0.75 \times 111 \times 2.5}{2.47 \times 0.0821 \times 300} = 3.42 \text{ g}$$

Hence, the required amount of CaCl_2 is 3.42 g.

2.41 Determine the osmotic pressure of a solution prepared by dissolving 25 mg of K_2SO_4 in 2 litre of water at $25^\circ C$, assuming that it is completely dissociated.

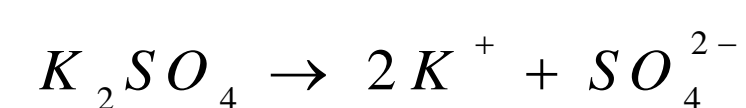
Answer 2.41

$$w = 25 \text{ mg} = 0.025 \text{ g}$$

$$V = 2 \text{ L}$$

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

When K_2SO_4 is dissolved in water, following reaction will take place:



Total number of ions = 3

$$i = 3$$

$$R = 0.0821 \text{ L atm / K mol}$$

$$M = (2 \times 39) + (1 \times 32) + (4 \times 16) = 174 \text{ g/mol}$$

We know that,

$$\begin{aligned} \pi &= \frac{inRT}{v} = \frac{iwRT}{Mv} \\ &= \frac{3 \times 0.025 \times 0.0821 \times 298}{174 \times 2} = 5.27 \times 10^{-3} \text{ atm} \end{aligned}$$