

CBSE NCERT Solutions for Class 12 Chemistry Chapter 2

Back of Chapter Questions

- 2.1. Calculate the mass percentage of benzene (C_6H_6) and carbon tetrachloride (CCl_4) if 22 g of benzene is dissolved in 122 g of carbon tetrachloride.

Solution:

Given: mass of benzene (C_6H_6 /solute) = 22 gram

Mass of carbon tetrachloride (CCl_4 /solvent) = 122 gram

We Find: mass percentage of benzene (C_6H_6) = ?

Mass percentage of carbon tetrachloride (CCl_4) = ?

Formula required: Mass percentage of solute = $\frac{\text{Mass of solute (gm)}}{\text{Total mass of the solution (gm)}} \times 100\%$

$$= \frac{\text{Mass of solute (gm)}}{\text{mass of the solute (gm) + mass of the solvent (gm)}} \times 100\%$$

Calculation: Mass percentage of C_6H_6 = $\frac{\text{Mass of } C_6H_6(\text{gm})}{\text{Total mass of the solution (gm)}} \times 100\%$

$$= \frac{\text{Mass of } C_6H_6(\text{gm})}{\text{mass of the solute (gm) + mass of the solvent (gm)}} \times 100\%$$

$$= \frac{\text{Mass of } C_6H_6(\text{gm})}{\text{Mass of } C_6H_6(\text{gm}) + \text{Mass of } CCl_4} \times 100\%$$

$$= \frac{22}{22 + 122} \times 100\%$$

$$= 15.28\%$$

Mass percentage of CCl_4 = $\frac{\text{Mass of } CCl_4}{\text{Total mass of the solution}} \times 100\%$

$$= \frac{\text{Mass of } CCl_4(\text{gm})}{\text{Mass of } C_6H_6(\text{gm}) + \text{Mass of } CCl_4(\text{gm})} \times 100\%$$

$$= \frac{122}{22 + 122} \times 100\%$$

$$= 84.72\%$$

Alternatively,

$$\text{Mass percentage of } CCl_4 = (100 - 15.28)\% = 84.72\%$$

- 2.2. Calculate the mole fraction of benzene in solution containing 30% by mass in carbon tetrachloride.

Solution:

Give % by mass C_6H_6 in $CCl_4 = 30\%$

benzene in solution containing 30% by mass in carbon tetrachloride.

Assume:

Let the total mass of the solution be 100 g and the mass of benzene be 30

\therefore Mass of carbon tetrachloride = $(100 - 30)g = 70 g$

We know that Molar mass of benzene (C_6H_6) = $(6 \times 12 + 6 \times 1)g \text{ mol}^{-1} = 78 g \text{ mol}^{-1}$

Molar mass of carbon tetrachloride (CCl_4) = $1 \times 12 + 4 \times 35.5 = 154 g \text{ mol}^{-1}$

The formula used for mole fraction:

$$\frac{\text{Number of moles of solute } (n_1)}{\text{Number of moles of solute } (n_1) + \text{Number of moles of solute } (n_2)}$$

$$\therefore \text{Number of moles of } C_6H_6(n_1) = \frac{30}{78} \text{ mol} = 0.3846 \text{ mol}$$

$$\therefore \text{Number of moles of } CCl_4(n_2) = \frac{70}{154} \text{ mol} = 0.4545 \text{ mol}$$

the mole fraction of C_6H_6 is given as:

$$\begin{aligned} & \frac{\text{Number of moles of } C_6H_6(n_1)}{\text{Number of moles of } C_6H_6(n_1) + \text{Number of moles of } CCl_4(n_2)} \\ &= \frac{0.3846}{0.3846 + 0.4545} \\ &= 0.458 \end{aligned}$$

- 2.3. Calculate the molarity of each of the following solutions:

- (a). 30 g of $Co(NO_3)_2 \cdot 6H_2O$ in 4.3 L of solution
 (b). 30 mL of 0.5 M H_2SO_4 diluted to 500 mL.

Solution:

Molarity is given by:

(a) Given : $Co(NO_3)_2 \cdot 6H_2O = 30 \text{ gm}$

Volume of solution = 4.3 liter

We know : Molar mass of $\text{Co}(\text{NO}_3)_2 \cdot 6\text{H}_2\text{O} = 59 + 2(14 + 3 \times 16) + 6 \times 18$

$$= 291 \text{ g mol}^{-1}$$

$$\text{Moles of } (\text{NO}_3)_2 \cdot 6\text{H}_2\text{O} = \frac{(\text{weight})(\text{gm})}{\text{molecularweight}} = \frac{30}{291} \text{ mol}$$

$$= 0.103 \text{ mol}$$

The formula used : $\text{Molarity} = \frac{\text{Moles of solute}}{\text{Volume of solution in litre}}$

$$\text{Therefore, molarity} = \frac{0.103 \text{ mol}}{4.3 \text{ L}} = 0.023 \text{ M}$$

(b) Given : 30 mL of 0.5 M H_2SO_4

Initial Volume of H_2SO_4 solution = 30 ml

Number of moles present in 1000 mL of 0.5 M $\text{H}_2\text{SO}_4 = 0.5 \text{ mol}$

$$\text{Number of moles present in 30 mL of 0.5 M } \text{H}_2\text{SO}_4 = \frac{0.5 \times 30}{1000} \text{ mol}$$

$$= 0.015 \text{ mol}$$

$$\text{Therefore, Molarity} = \frac{\text{Moles of solute}}{\text{Volume of solution in litre}} = \frac{\text{Moles of solute} \times 1000}{\text{Volume of solution in Mili litre}}$$

$$\text{Moles of solute} = MV(\text{litre})$$

$$\text{Final Volume of the solution after dilution} = \text{molarity} = \frac{M_i V_i}{V_f} = \frac{0.015}{0.5 \text{ L}} \text{ mol}$$

$$= 0.03 \text{ M}$$

2.4. Calculate the mass of urea (NH_2CONH_2) required in making 2.5 kg of 0.25 molal aqueous solution.

Solution:

Given : molality of aqueous solution of urea = 0.25

we know

$$\text{weight of solution} = 2.5 \text{ kg} = 2500 \text{ g}$$

$$\text{Molar mass of urea } (\text{NH}_2\text{CONH}_2) = 2(1 \times 14 + 2 \times 1) + 1 \times 12 + 1 \times 16 = 60 \text{ g mol}^{-1}$$

0.25 molar aqueous solution of urea means that

$$1000 \text{ g of water contains } 0.25 \text{ mol} = (0.25 \times 60) \text{ g of urea}$$

$$= 15 \text{ g of urea}$$

That is,

Weight of solute = 15 g

Weight of solvent = 1000 g, weight of solution = (15 + 1000)g

(1000 + 15) g of solution contains 15 g of urea = 15 × 2500

Therefore, 2.5 kg (2500 g) of solution contains = $\frac{15 \times 2500}{1000 + 15} = 36.945$ g

= 36.945 g of urea (approximately)

Hence, the mass of urea required = 36.945 g

2.5. Calculate

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

Solution:

Given density of aqueous KI solution = $\frac{1.202 \text{ gm}}{\text{mol}}$

% mass by mass of KI = 20%

- The molar mass of KI = 39 + 127 = 166 g mol⁻¹

That is,

20% (mass/mass) aqueous solution of KI means 20 g of KI is present in 100 g of solution.

20 g Of KI is present in (100 – 20)g of water = 80 g of water

Therefore, molality of the solution (m) = $\frac{\text{Moles of KI}}{\text{Mass of water in kg}}$

Moles of KI = wt/mol = $\frac{20}{166}$ mole

Weight of water (solvent) = 80 g = $\frac{80}{100}$ kg.

$\frac{20}{166}$
= $\frac{20}{166}$ molal

= 1.506 molal

% mass by mass = 20%

Mass by mass means

Mass of solution = 100 g and mass of solvent = 20g

(b) given: the density of the solution = 1.202 g mL^{-1}

Volume of 100 g solution = $\frac{\text{mass}}{\text{density}}$

$$= \frac{100 \text{ g}}{1.202 \text{ g mL}^{-1}}$$

$$= 83.19 \text{ mL}$$

$$= 83.19 \times 10^{-3} \text{ L}$$

Moles of solute = weight/molecular weight = $\frac{20}{166}$

Formulae used molarity (M) = $\frac{\text{moles of solute}}{\text{volume of solution (litre)}}$

Therefore, the molarity of the solution = $\frac{\frac{20}{166} \text{ mol}}{83.19 \times 10^{-3} \text{ L}}$

$$= 1.45 \text{ M}$$

Given 20% w/w KI

Weight of solute = 20 g

Weight of solution = 100 g

Weight of solvent = weight of solution – weight of solute

$$\Rightarrow \text{weight of solvent} = 100 - 20 = 80 \text{ g}$$

(c) Moles of KI = $\frac{\text{weight of KI}}{\text{Molecular weight of KI}}$

$$= \frac{20}{166} = 0.12 \text{ mol}$$

Moles of water = $\frac{\text{weight of H}_2\text{O gm}}{\text{molecular weight of H}_2\text{O}} \frac{80}{18} = 4.44 \text{ mol}$

Exercises

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

Solution:

A solution is a homogeneous mixture of two or more than two components (solute and solvent)

There are three types of solutions, they are:

(i) Gaseous solution:

A gaseous solution is a solution in which the solvent is gas is called a gaseous solution. In these solutions, the solute may be liquid, solid, or gas. For example, a homogenous mixture of oxygen and hydrogen gas is a gaseous solution.

(ii) Liquid solution:

A liquid solution is a solution in which the solvent is a liquid is known as a liquid solution. The solute in these solutions may be gas, liquid, or solid.

The mixture of milk and water is an example of a liquid solution.

(iii) Solid solution:

A solid solution is a solution in which the solvent is a solid dissolved. The solute may be gas, liquid or solid. Metal alloys are examples of solid solutions.

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

Solution:

A solid solution is formed between two substances solute and solvent (where one substance has very large particles and the other is having very small particles), an interstitial solid solution will be formed. Consider an example, a solution of hydrogen in palladium is a solid solution in which the solute is a gas (H_2) and the solvent is Pd.

2.3. Define the following terms:

(i) Mole fraction

(ii) Molality

(iii) Molarity

(iv) Mass percentage.

Solution:

(i) Mole fraction:

The mole fraction of a component in a mixture 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.

i.e., Mole fraction of component (X) =
$$\frac{\text{Number of moles of the component}}{\text{Total number of moles of all components}}$$

If, in a binary solution, the number of moles of the solute is n_A and of the solvent is n_B , then the mole fraction of the solute in the solution is $(x_A) = \frac{n_A}{n_A+n_B}$

Similarly, the mole fraction of the solvent in the solution is $(x_B) = \frac{n_B}{n_A+n_B}$

And, moles fraction of solute (x_A) + mole fractions of the solvent $(x_B) = 1$

Molarity is a unitless quantity.

(ii) Molality

It is unitless quantity

The number of moles of the solute per kilogram of the solvent is called as molality. It is expressed as:

$$\text{Molality (m)} = \frac{\text{Moles of solute (n)}}{\text{Mass of solvent in (kg)}} = \frac{\text{moles of solute} \times 1000}{\text{mass of solvent (g)}}$$

Unit is molal

(iii) Molarity

The number of moles of the solute dissolved in one litre of the solution is called molarity of the solution.

It is expressed as:

$$\text{Molarity (M)} = \frac{\text{Moles of solute}}{\text{Volume of solution in (litre)}} = \frac{\text{moles of solute (1000)}}{\text{volume of solution (ml)}}$$

(iv) Mass percentage:

$$\frac{\text{mass of solute}}{\text{mass of solution}} \times 100 = \frac{\text{mass of solute}}{\text{mass of solute} + \text{mass of solvent}} \times 100$$

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} ?

Solution:

Given density of solution = 1.50 g/mL

We know that concentrated nitric acid used in laboratory work is 68% nitric acid by mass in an aqueous solution. This means that 68 g of nitric acid is dissolved in 100 g of the solution.

Molar mass of nitric acid (HNO_3) = $1 \times 1 + 1 \times 14 + 3 \times 16 = 63 \text{ g mol}^{-1}$

The formula used: $M = \frac{\text{moles of solute}}{\text{volume of solution(litre)}}$

$$\text{Then, the number of moles of HNO}_3 = \frac{\text{weight}}{\text{molecular weight}} = \frac{68}{63} \text{ mol}$$

$$= 1.079 \text{ mol}$$

We know that density of solution = 1.504 g mL^{-1}

i.e., the mass of solution = 100 g by mass percentage

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

$$= 66.49 \text{ mL}$$

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

$$\therefore 1 \text{ ml} = 10^{-3} \text{ litre}$$

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

$$= 16.23 \text{ M}$$

68% Nitric acid by mass means that in a solution of 100g, 68g is solute and rest is solvent i. e. 32g.

- 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 the solution is 1.2 g mL^{-1} , then what shall be the molarity of the solution?

Solution:

Given,

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

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

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

$$= 0.056 \text{ mol}$$

$$\therefore \text{The molality of solution} = \frac{\text{moles of solute}}{\text{weight of solvent (kg)}} = \frac{0.056 \text{ mol}}{0.09 \text{ kg}} = 0.62 \text{ molal}$$

$$\text{Number of moles of water} = \frac{90 \text{ g}}{18 \text{ g mol}^{-1}} = 5 \text{ mol}$$

$$\text{Mole fraction of glucose}(x_g) = \frac{\text{moles of glucose}}{\text{moles of glucose} + \text{moles of H}_2\text{O}} = \frac{0.056}{0.056+5}$$

$$\Rightarrow \text{Mole fraction of glucose} = 0.011$$

$$\text{And, the mole fraction of water } (x_w) = 1 - \text{moles of glucose} = 1 - x_g$$

$$= 1 - 0.011$$

$$= 0.989$$

If the density of the solution is 1.2 g mL^{-1} , then the volume of the 100 g solution

$$= \frac{\text{mass}}{\text{density}}$$

$$\text{Volume of solution } (v) = \frac{100\text{g}}{1.2\text{g mL}^{-1}} = 83.33 \text{ mL} = 83.33 \times 10^{-3}\text{L}$$

$$\therefore \text{Molarity of the solution} = \frac{\text{moles of solute}}{\text{volume of solution (litre)}} = \frac{0.056\text{mol}}{83.33 \times 10^{-3}\text{L}}$$

$$= 0.67 \text{ M}$$

- 2.6. How many mL of 0.1 M HCl are required to react completely with 1 g mixture of Na_2CO_3 and NaHCO_3 containing equimolar amounts of both?

Solution:

Given molarity of HCl = 0.1 M

Volume solution = ?

Assume the amount of Na_2CO_3 in the mixture to be x g.

Then, the amount of NaHCO_3 in the mixture is $(1 - x)$ g.

Weight of mixture of Na_2CO_3 & $\text{NaHCO}_3 = 1 \text{ gm}$

Molar mass of $\text{Na}_2\text{CO}_3 = 2 \times 23 + 1 \times 12 + 3 \times 16 = 106 \text{ g mol}^{-1}$

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

Molar mass of $\text{NaHCO}_3 = 1 \times 23 + 1 \times 12 + 3 \times 16 = 84 \text{ g mol}^{-1}$

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

According to question Na_2CO_3 and NaHCO_3 have equimolar so moles of

$\text{Na}_2\text{CO}_3 = \text{moles of } \text{NaHCO}_3$

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

$$\Rightarrow 84x = 106 - 106x$$

$$\Rightarrow 190x = 106$$

$$x = 0.5579$$

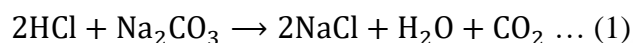
$$\text{Therefore, the number of moles of Na}_2\text{CO}_3 = \frac{x}{106} = \frac{0.5579}{106} \text{ mol}$$

$$= 0.0053 \text{ mol}$$

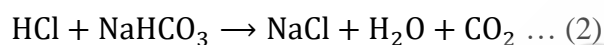
$$\text{And the number of moles of NaHCO}_3 = \frac{1-x}{84} = \frac{1-0.5579}{84}$$

$$= 0.0053 \text{ mol}$$

HCl Reacts with Na_2CO_3 and NaHCO_3 according to the following equation.



2 mol 1 mol



1 mol 1 mol

1 mol Of Na_2CO_3 reacts with 2 mol of HCl.

Therefore, 0.0053 mol of Na_2CO_3 reacts with $2 \times 0.0053 \text{ mol} = 0.0106 \text{ mol}$.

Similarly, 1 mol of NaHCO_3 reacts with 1 mol of HCl. Then for mole of HCl = moles of $\text{NaHCO}_3 = 0.0053$

Total moles of HCl required = $(0.0106 + 0.0053) \text{ mol} = 0.0159 \text{ mol}$

$\Rightarrow 0.1 \text{ M}$ of HCl implies-

0.1 mol of HCl is present in 1000 mL of the solution.

$\therefore 0.0159 \text{ mol}$ of HCl is present in $\frac{1000 \times 0.0159}{0.1} \text{ ml}$ of solution

= 159 mL Of the solution

Hence, 159 mL of 0.1 M of HCl is required to react completely with 1 g mixture of Na_2CO_3 and NaHCO_3 , containing equimolar amounts of both.

- 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.

Solution:

Given the weight of solution = 300 g

% of solution $S_1 = 25\%$

Weight of solution 2 = 400 g

$$\% \text{ of solution } S_2 = 40\%$$

The total amount of solute present in the mixture is given by,

$$= 300 \text{ gm} \times 25\% + 400 \text{ gm} \times 40\%$$

$$300 \times \frac{25}{100} + 400 \times \frac{40}{100}$$

$$= 75 + 160$$

$$= 235 \text{ g}$$

$$\text{Total amount of solution} = 300 + 400 = 700 \text{ g}$$

Therefore, mass percentage (w/w) of the solute in the resulting solution,

$$= \frac{235}{700} \times 100\%$$

$$= 33.57\%$$

And, mass percentage (w/w) 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?

Solution:

Given the weight of ethylene glycol = 222.6 g (solute)

Weight of water = 200 g (solvent)

We know,

The molar mass of ethylene glycol [$\text{C}_2\text{H}_4(\text{OH})_2$]

$$= 2 \times 12 + 6 \times 1 + 2 \times 16$$

$$= 62 \text{ gmol}^{-1}$$

$$\text{Molality} = \frac{\text{moles of solute}}{\text{weight of solvent (kg)}}$$

$$\text{Number of moles of ethylene glycol} = \frac{\text{weight of ethylene glycol}}{\text{molecular weight of ethylene glycol}} = \frac{222.6 \text{ g}}{62 \text{ gmol}^{-1}}$$

$$= 3.59 \text{ mol}$$

Therefore, molality of the solution = $\frac{\text{moles of ethylene glucol}}{\text{weight of solvent (water)kg}} = \frac{3.59 \text{ mol}}{0.200 \text{ kg}} = 17.95 \text{ molal}$

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

Given,

The density of the solution = 1.072 g mL^{-1}

$$\begin{aligned} \therefore \text{The volume of the solution} &= \frac{\text{mass}}{\text{density}} = \frac{422.6 \text{ g}}{1.072 \text{ g mL}^{-1}} \\ &= 394.22 \text{ mL} \\ &= 0.3942 \times 10^{-3} \text{ L} \end{aligned}$$

$$\begin{aligned} \Rightarrow \text{Molarity of the solution} &= \frac{\text{moles of solute}}{\text{volume of solution}} = \frac{3.59}{0.39422 \times 10^{-3} \text{ L}} \\ &= 9.11 \text{ M} \end{aligned}$$

2.9. A sample of drinking water was found to be severely contaminated with chloroform (CHCl_3) supposed to be a carcinogen. The level of contamination was 15 ppm (by mass):

- (i). express this in per cent by mass
- (ii). determine the molality of chloroform in the water sample.

Solution:

Given 15ppm by mass of solution

- (i) 15 ppm (by mass) means 15 parts are present per million(10^6) parts of the solution.

$$\begin{aligned} \text{Therefore, percent by mass} &= \frac{\text{mass of solute} \times 100}{\text{mass of solution}} = \frac{15}{10^6} \times 100\% \\ &= 1.5 \times 10^{-5}\% \end{aligned}$$

- (ii) We know

$$\begin{aligned} \text{Molar mass of chloroform (CHCl}_3) &= 1 \times 12 + 1 \times 1 + 3 \times 35.5 \\ &= 119.5 \text{ g mol}^{-1} \end{aligned}$$

Now, according to the question

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

$$\text{Moles of chloroform solute} = \frac{\text{weight}}{\text{molecular weight}} = \frac{15}{119.5}$$

The formula used molarity = $\frac{\text{moles of solute}}{\text{weight of solvent (kg)}}$

i. e., 15g Of chloroform is present in $(10^6 - 15) \approx 10^6$ g of water.

$$\therefore \text{The molality of the solution} = \frac{\frac{15}{119.5} \text{ mol}}{10^6 \times 10^{-3} \text{ kg}}$$

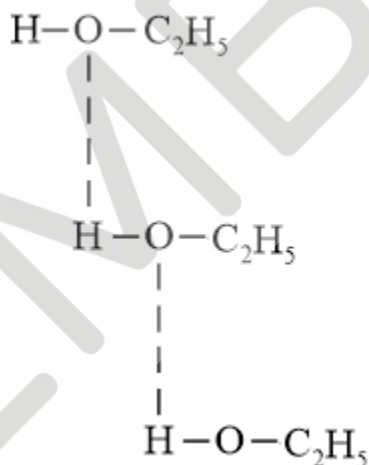
$$= 1.26 \times 10^{-4} \text{ molal}$$

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

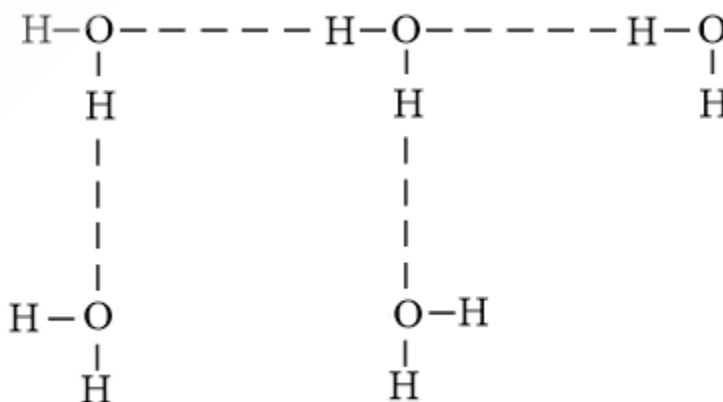
Solution:

In pure alcohol and water, the molecules of alcohol are held tightly by strong hydrogen bonds. The interactions between the molecules of alcohol and water are weaker than alcohol-alcohol and water-water interactions. As a result of this, when alcohol and water are mixed, the intermolecular interactions become weaker and the molecules can easily escape. This lowers the boiling point of the solution.

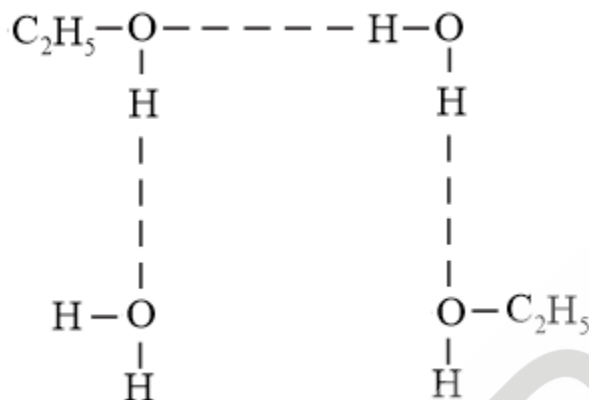
Hydrogen bond in the ethanol molecules



Hydrogen bond in water molecule



Hydrogen bond in between Ethanol and water.



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

Solution:

As temperature increases, the solubility decreases since the dissolution of gases in liquids is an exothermic process.

The solubility of a gas decreases with the increase in temperature.

Gas + Liquid \rightarrow Solution + Heat

The process is exothermic so according to Le Chatelier's Principle the forward reaction is exothermic thus with an increase in temperature or heat supplied the Equilibrium will shift backwards., thereby decreasing the solubility of gases is decreased.

- 2.12. State Henry's law and mention some important applications.

Solution:

According to Henry's law, the partial pressure of a gas in the vapour phase is proportional to the mole fraction of the gas in the solution. If p is the partial pressure of the gas in the vapour phase and x is the mole fraction of the gas, then Henry's law can be expressed as:

$$p \propto x$$

$$p = K_H x$$

Where,

K_H Is Henry's law constant.

Some important applications of Henry's law are mentioned below.

- (i) In soft drink and soda water CO_2 filled bottles are under high-pressure increase the solubility of CO_2 .

Henry's law states that the solubility of gases increases with an increase in pressure. When a scuba diver dives deep into the sea, the high pressure causes nitrogen to get dissolved in his blood. When the diver returns to the surface, the solubility of nitrogen again decreases due to a decrease in pressure upon reaching the surface. The dissolved nitrogen gas is released, which leads to the formation of nitrogen bubbles in the blood as nitrogen is insoluble in blood at low pressures. This process results in the blockage of capillaries and leads to a medical condition known as 'bends' or 'decompression sickness'.

- (ii) The concentration of oxygen is low in the blood vessels and tissues of people living at high altitudes. The partial pressure of oxygen at high altitude is lower than that of the sea level. Therefore people living in high altitudes will have less percentage of oxygen in their blood. This makes them light-headed and weak often preventing 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?

Solution:

Given partial pressure of Ethane over solution = 1 bar

Weight of ethane = 6.56×10^{-3} gm

Weight of Ethane in solution = 5×10^{-2} gm

We know Molar mass of ethane (C_2H_6) = $2 \times 12 + 6 \times 1$

= 30 g mol^{-1}

\therefore Number of moles present in 6.56×10^{-2} g of ethane = $\frac{\text{weight (gm)}}{\text{Molecular weight}} =$

$$\frac{6.56 \times 10^{-2}}{30}$$

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

Assume the number of moles of the solvent to be x.

According to Henry's law,

$$p = K_H x$$

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

$$\Rightarrow 1 \text{ bar} = K_H \frac{2.187 \times 10^{-3}}{x} \text{ (Since } x \gg 2.187 \times 10^{-3} \text{)}$$

$$\Rightarrow K_H = \frac{x}{2.187 \times 10^{-3}} \text{ bar} \dots (1)$$

$$\text{Number of moles present in } 5.00 \times 10^{-2} \text{ of ethane} = \frac{5.00 \times 10^{-2}}{30} \text{ mol}$$

$$= 1.67 \times 10^{-3} \text{ mol}$$

$$\text{Mole fraction of Ethane} = \frac{\text{moles of ethane}}{\text{moles of ethane} + \text{moles of solvent}} = \frac{1.67 \times 10^{-3}}{1.67 \times 10^{-3} + x}$$

According to Henry's law,

$$p = K_H x$$

$$= \frac{x}{2.187 \times 10^{-3}} \times \frac{1.67 \times 10^{-3}}{(1.67 \times 10^{-3}) + x}$$

$$= \frac{x}{2.187 \times 10^{-3}} \times \frac{1.67 \times 10^{-3}}{x} \text{ (Since, } x \gg 1.67 \times 10^{-3} \text{)}$$

$$= 0.764 \text{ bar}$$

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

Solution:

According to Raoult's law, the partial vapour pressure of each volatile component in any solution is directly proportional to its mole fraction.

Present in the solution $p_1 \propto x_1$

And $p_1 = p_1^\circ x_1$

Where p° = vapour pressure of pure component at all the same temperature

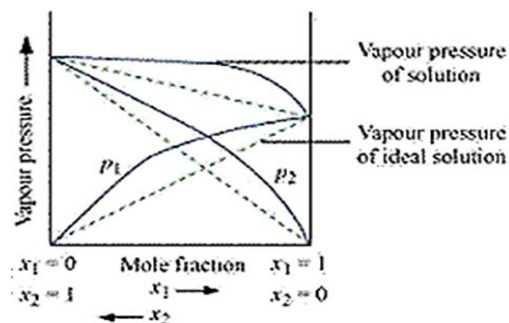
p = partial vapour pressure of component

x = mole fraction of volatile component

The solutions that obey Raoult's law at each and every concentration of the solution are termed as ideal solutions.

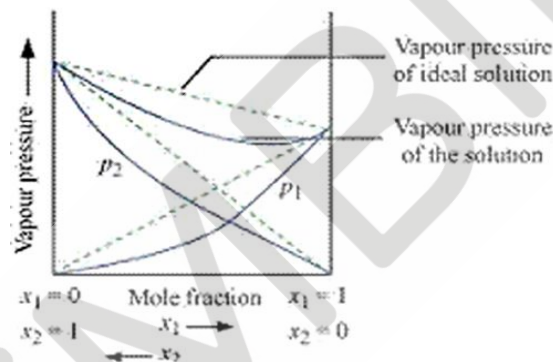
The solutions that do not obey Raoult's law (non-ideal solutions) have vapour pressures either higher or lower than that predicted by Raoult's law. If the vapour pressure is higher, then the solution is said to exhibit positive deviation.

Vapour pressure of a two-component solution showing positive deviation from Raoult's law: $\Delta S > 0$; $\Delta G > 0$; $V_{p(\text{mixture})} > 0$; $\Delta H < 0$



A solution that shows negative deviation: If vapour pressure is lower than the predicted value, then the solution is said to exhibit negative deviation from Raoult's law.

Vapour pressure of a two-component solution showing negative deviation from Raoult's law:



In the case of an ideal solution, the enthalpy of the mixing of the pure components for forming the solution is zero. $\Delta S > 0$; $\Delta G < 0$; $V_{\text{mix}} < 0$; $\Delta H > 0$

In solutions showing negative deviation, heat gets evolved from the solution.

- 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?

Solution:

Aqueous solution of non-volatile solute = 2%

Here,

Vapour pressure of the solution at normal boiling point (p_1) = 1.004 bar

We know Vapour pressure of pure water (solvent) at normal boiling point (p_1^0) = 1.013 bar

Mass of solute, (w_2) = 2 g

Mass of solvent (water), (w_1) mass of solution – mass of solute = 100 – 2

The molar mass of solvent (water), $(M_1) = 2 \times 1 + 16 = 18 \text{ g mol}^{-1}$

According to Raoult's law,

Let n_1 & n_2 are number of moles of solvent and solute

$$\text{Mole of solute} = \frac{(n_1)}{\text{moles of solution } (n_2)}$$

$$\text{Moles of solvent } (n_2) = \frac{\text{weight}}{\text{molecular weight}} = \frac{98}{18}$$

$$\text{Moles of solute } (n_1) = \frac{\text{weight}}{\text{molecule weight}} = \frac{1}{M_1}$$

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

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

$$\Rightarrow \frac{0.009}{1.013} = \frac{2 \times 18}{M_1 \times 98}$$

$$\Rightarrow M_1 = \frac{1.013 \times 2 \times 18}{0.009 \times 98}$$

$$= 41.35 \text{ g mol}^{-1}$$

- 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?

Solution:

Given,

Vapour pressure of heptane (p_1^0) = 105.2 kPa

Weight of heptane = 26.0 g

Weight of octane = 35 g

Vapour pressure of octane (p_2^0) = 46.8 kPa

We know that,

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

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

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

$$= 0.26 \text{ mol}$$

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

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

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

$$= 0.31 \text{ mol}$$

$$\text{Mole fraction of heptane, } (x_1) = \frac{\text{moles of heptane}}{\text{moles of heptane} + \text{moles of octane}} = \frac{0.26}{0.26 + 0.31}$$

$$0.456$$

$$\text{And, mole fraction of octane, } (x_2) = 1 - 0.456$$

$$= 0.544$$

Now, the partial pressure of heptane, = moles fraction of heptane \times vapour pressure of heptane(p_1°);

$$(p_1) = x_1 p_1^\circ$$

$$= 0.456 \times 105.2$$

$$= 47.97 \text{ kPa}$$

Partial pressure of octane, = mole fraction of octane \times vapour pressure of octane(p_2°);

$$(p_2) = x_2 p_2^\circ$$

$$= 0.544 \times 46.8$$

$$= 25.4592 \text{ kPa}$$

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

Solution:

Given vapour pressure of water 12.3 kPa

Molality of solution = 1 (molal)

1 Molal solution means one mol of the solute is present in 1000 g of the solvent (water).

Mole of solute = 1

Weight of solvent = 1000 g

Molar mass of water = 18 g mol^{-1}

$$\begin{aligned} \therefore \text{Number of moles present in } 1000 \text{ g of water} &= \frac{1000}{18} \\ &= 55.56 \text{ mol} \end{aligned}$$

Therefore, mole fraction of the solute present in the solution (x_2) = $\frac{\text{moles of solute}}{\text{moles of solute} + \text{moles of solvent}} = \frac{1}{1 + 55.56} = 0.0177$

Relative vapour pressure = mole fraction of solute (x_2)

p_1° = Vapour pressure of solvent

p_1 = Vapour pressure of solution

Formula, $\frac{p_1^\circ - p_1}{p_1^\circ} = x_2 \Rightarrow \frac{12.3 - p_1}{12.3} = 0.0177 \Rightarrow 12.3 - p_1 = 0.217 \Rightarrow p_1 = 12.0823$

= 12.08 kPa (Approximately)

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^{-1}) which should be dissolved in 114 g octane to reduce its vapour pressure to 80%.

Solution:

Given molar mass of non-volatile solute = 40 g/mol

Solvent = octane

Weight of solvent = 114 g

Let the vapour pressure of pure octane be p_1° .

Then, from the question the vapour pressure of the octane after dissolving the non-volatile solute = $\frac{80}{100} p_1^\circ = 0.8p_1^\circ$.

Assume weight of solute = w_2

Molar mass of octane, (C_8H_{18}), $M_1 = 8 \times 12 + 18 \times 1$
= 114 g mol^{-1}

Formula use

$$\begin{aligned} &\frac{\text{vapour pressure of solvent} - \text{vapour pressure of solutions}}{\text{vapour pressure of solvent}} \\ &= \frac{\text{moles of solute}}{\text{moles of solvent}} \end{aligned}$$

$$\text{Moles of solute} = \frac{w_2}{40}$$

$$\text{Moles of solvent} = \frac{40}{114}$$

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

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

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

$$\Rightarrow 0.2 = \frac{w_2}{40}$$

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

Hence, the mass required 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:

- (i) molar mass of the solute
- (ii) vapour pressure of water at 298 K.

Solution:

Given,

Weight of non-volatile solute = 30 g

Weight of H₂O (solvent) = 90 g

Added amount of water = 18 g in solution

- (i) Let the molar mass of the solute = $M \text{ g mol}^{-1}$

$$\text{Number of moles of solvent (water), } (n_1) = \frac{\text{weight of water}}{\text{molecular weight of water}} = \frac{90 \text{ g}}{18 \text{ g mol}^{-1}} = 5 \text{ mol}$$

$$\text{And, the no. of moles of solute, } (n_2) = \frac{\text{weight of solute}}{\text{molecular weight of solute}} = \frac{30 \text{ g}}{M \text{ mol}^{-1}} = \frac{30}{M} \text{ mol}$$

Formula:

$$\frac{\text{vapour pressure of solvent } (p_1^{\circ}) - \text{vapour pressure } (p_1)}{\text{vapour pressure of solvent } (p_1^{\circ})}$$

= mole fraction of solute(x)

$$\text{mole fraction of solute} = \frac{\text{moles of solute}}{\text{moles of solute} + \text{mole of solvent}}$$

$$= \frac{\frac{30}{M}}{5 + \frac{30}{M}}$$

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

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

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

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

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

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

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

$$\Rightarrow \frac{\frac{2.8}{p_1^{\circ}}}{\frac{2.8}{2.8}} = \frac{5M + 30}{5M} \dots\dots\dots (i)$$

After the addition of 18 g of water:

$$\text{Total weight of solvent (water)} = 90 + 18$$

$$\text{Moles of water } (n_1) = \frac{\text{weight of water}}{\text{molecular weight of water}} = \frac{90+18 \text{ g}}{18} = 6\text{mol}$$

$$\text{Moles of solute} = \frac{\text{moles of solute } (n_2)}{\text{moles of solute}(n_1) + \text{mole of solvent}(n_2)}$$

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

Formula:

p_1° = Vapour pressure of solvent

p_1 = Vapour pressure of solution = $2.9 \times \text{kPa}$

Moles of solute = $\frac{\text{weight}}{\text{molecular weight}} = \frac{30}{M}$

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

$$\Rightarrow \frac{p_1^{\circ} - 2.9}{p_1^{\circ}} = \frac{30/M}{6 + \frac{30}{M}}$$

$$\Rightarrow 1 - \frac{2.9}{p_1^{\circ}} = \frac{30/M}{\frac{6M + 30}{M}}$$

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

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

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

$$\Rightarrow \frac{p_1^{\circ}}{2.9} = \frac{6M + 30}{6M} \dots\dots\dots \text{(ii)}$$

Dividing equation (i) by (ii), we have:

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

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

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

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

$$\Rightarrow 3M = 69$$

$$\Rightarrow M = 23\text{u}$$

Therefore, the molar mass of the solute is 23 g mol^{-1} .

(ii) Substituting the value of M in eqn (i), we get:

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

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

$$\Rightarrow p_1^0 = 3.53$$

Hence, the vapour pressure of water at 298 K is 3.53 kPa.

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

Solution:

Given:

Freezing point can sugar in of water = 271 K

Freezing point of water = 0°C = 273.15K

ΔT_f = Freezing point of solvent – freezing point of solution

Here, ΔT_f = (273.15 – 271)K

= 2.15 K

Molar mass of sugar ($C_{12}H_{22}O_{11}$) = $12 \times 12 + 22 \times 1 + 11 \times 16$

= 342 g mol⁻¹

5% solution by mass of cane sugar in H₂O means 5 g of cane sugar is present in (100 – 5)g = 95g of water.

Now, number of moles of cane sugar = $\frac{\text{weight}}{\text{molecular weight}} = \frac{5}{342}$ mol

= 0.0146 mol

Therefore, molality of the solution, (m) = $\frac{\text{moles of solute}}{\text{weight of solvent (kg)}} = \frac{0.0146 \text{ mol}}{0.095 \text{ kg}}$

= 0.1537 mol kg⁻¹

Formula relative lowering vapour pressure \propto molality

Relative lowering vapour pressure = kg \times molality

$\Delta T_f = K_f \times m$

$\Rightarrow K_f = \frac{\Delta T_f}{m}$

$$= \frac{2.15\text{K}}{0.1537 \text{ mol kg}^{-1}}$$

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

(ii) Given freezing point of glucose in water =?

Freezing point of water (solvent) = 273.15

Weight of solute = 5 g

Weight of solvent = 5 g

$$= 0.95 \text{ g}$$

We know Molar of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) = $6 \times 12 + 12 \times 1 + 6 \times 16$

$$= 180 \text{ g mol}^{-1}$$

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

$$\therefore \text{Number of moles of glucose} = \frac{\text{weight}}{\text{molecular weight}} = \frac{5}{180} \text{ mol}$$

$$= 0.0278 \text{ mol}$$

$$\text{Therefore, molality of the solution, (m)} = \frac{\text{moles of solute}}{\text{weight solvent in kg}} = \frac{0.0278 \text{ mol}}{0.095 \text{ kg}}$$

$$= 0.2926 \text{ mol kg}^{-1}$$

Formula

\Rightarrow Relative lowering in freezing point \propto molality

\Rightarrow Relative lowering in freezing point = $\text{kg} \times \text{molality}$

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

$$= 13.99 \text{ K kg mol}^{-1} \times 0.2926 \text{ mol kg}^{-1}$$

$$\Delta T_f = 4.09 \text{ K}$$

Hence, the freezing point of 5% glucose solution $\Delta T_f =$ freezing point of solvent - freezing point of solution

$$\Rightarrow (273.15 - 4.09)\text{K} = 269.06 \text{ K.}$$

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.

Solution:

Given the formula of compound AB_2 and AB_3

Weight of Benzene = 20 gm

Weight of AB_2 = 1 g ; $(\Delta T_f)_{AB_2} = 3.5$

Weight of AB_4 = 1g ; $(\Delta T_f)_{AB_4} = 1.3$

Formula $\Delta T_f = k_f \times \text{molality} \times i$

$$\therefore \text{Molality} = \frac{\text{moles of solute} \times 100}{\text{weight of solvent (kg)}}$$

i = Van't Hoff factor

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

$$\text{Then, } M_{AB_2} = \frac{1000 \times 1 \times 5.1}{2.3 \times 20}$$

$$= 110.87 \text{ g mol}^{-1}$$

$$M_{AB_4} = \frac{1000 \times 1 \times 5.1}{1.3 \times 20}$$

$$= 196.15 \text{ g mol}^{-1}$$

Now, we have the molar mass of AB_2 and AB_4 as $110.87 \text{ g mol}^{-1}$ and $196.15 \text{ g mol}^{-1}$ respectively.

Assume let the atomic masses of A and B x and y respectively, and they are non-electrolyte.

Now, we can write:

$$x + 2y = 110.87 \dots\dots\dots(i)$$

$$x + 4y = 196.15 \dots\dots\dots(ii)$$

Subtracting equation (i) from (ii), we have

$$2y = 85.28$$

$$y = 42.64$$

Putting the value of ' y ' in equation (1), we have

$$x + 2 \times 42.64 = 110.87$$

$$x = 25.59$$

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?

Solution:

Given,

$i = \text{Van't Hoff factor} \Rightarrow \text{for glucose } (i) = 1$

Concentration of glucose = $\frac{36}{180} = M$

\therefore Molecular weight of glucose = 180

Temperature (T_1) = 300 K

Osmotic pressure (π) = 1.52 bar

Gas constant $R = 0.083 \text{ bar LK}^{-1}\text{mol}^{-1}$

Osmotic pressure(π_2) = 1.52

$T_2 = 300 \text{ K}$

Formula: $\pi = iCRT$

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

$$\Rightarrow \frac{\pi_1}{C_1} = \frac{\pi_2}{C_2} \Rightarrow \frac{4.98}{0.2} = \frac{1.52}{C_2}$$

$$\Rightarrow C_2 = \frac{1.52 \times 0.2}{4.98}$$

$$= 0.61 \text{ M}$$

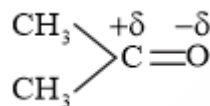
As the volume of the solution is 1L, the concentration of the solution would be 0.061 M.

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

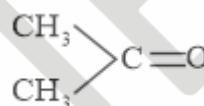
- (i) n-hexane and n-octane
- (ii) I_2 And CCl_4
- (iii) $NaClO_4$ And water
- (iv) Methanol and acetone
- (v) Acetonitrile (CH_3CN) and acetone (C_3H_6O).

Solution:

- (i) n-hexane and n-octane both are gases, so they have Van der Waals forces of attraction.
- (ii) I_2 And CCl_4 both are covalent compound so they have Van der Waals forces of attraction.
- (iii) $Na^+ClO_4^-$ is an ionic compound and H_2O is polar compound, so force between them is Ion-dipole interaction.



- (iv) Methanol CH_3OH and acetone both are polar compounds, so they have Dipole-dipole interaction.



- (v) Acetonitrile ($CH_3 - CN^{-\delta}$) and acetone both are polar compounds, so they have 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 .

Solution:

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

The order of increasing polarity is:

Cyclohexane < CH_3CN < CH_3OH < KCl (ionic)

$\therefore \epsilon$ N of O > N so more polar

Therefore, the order of increasing solubility in nonpolar solvent octane is

KCl < CH_3OH < CH_3CN < Cyclohexane

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

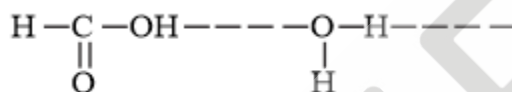
(vi). pentanol.

Solution:

(i) Phenol (C_6H_5OH) has the polar group - OH and non-polar group - C_6H_5 . Thus, phenol is partially soluble in water due to H-bonding, H-bond is weak due to crowding.

(ii) Toluene ($C_6H_5 - CH_3$) has no polar groups. Thus, toluene is insoluble in water.

(iii) Formic acid ($HCOOH$) has the polar group - OH and thus can form H-bond with water. Therefore formic acid is highly soluble in water.



(iv) Ethylene glycol  has polar - OH group and can form H-bond. Thus, it is highly soluble in water.

(v) Chloroform is insoluble in water because CCl_4 is non-polar

(vi) Pentanol ($C_5H_{11}OH$) has polar (-OH) group, but it also contains a bulky non-polar - C_5H_{11} group. Thus, pentanol is partially soluble in water due to poor H-bonding.

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.

Solution:

Given density of lake water = $1.25 \frac{\text{gm}}{\text{ml}}$

Weight of Na^+ ion = 92 gm

Weight of solvent (water) = 1 kg

Moles = $\frac{\text{weight}}{\text{Atomic weight of Na}}$

Number of moles present in 92 g of Na^+ ions = $\frac{92 \text{ g}}{23 \text{ g mol}^{-1}}$

= 4 molal

Formula : molality = $\frac{\text{moles of solute}}{\text{weight of solvent (kg)}}$ molal

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

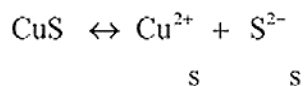
$$= 4 \text{ m}$$

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

Solution:

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

Assume Let s be the solubility of CuS in mol L^{-1} .



Now, $K_{\text{sp}} = [\text{Cu}^{2+}][\text{S}^{2-}]$

$$= s \times s$$

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

Then, we have, $K_{\text{sp}} = s^2 = 6 \times 10^{-16}$

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

$$= 2.45 \times 10^{-8} \text{ mol L}^{-1}$$

Hence, the maximum molarity of CuS in an aqueous solution is $2.45 \times 10^{-8} \text{ mol L}^{-1}$.

- 2.28.** Calculate the mass percentage of aspirin ($\text{C}_9\text{H}_8\text{O}_4$) in acetonitrile (CH_3CN) when 6.5 g of $\text{C}_9\text{H}_8\text{O}_4$ is dissolved in 450 g of CH_3CN .

Solution:

Given the weight of aspirin ($\text{C}_9\text{H}_8\text{O}_4$) = 6.5 gm

6.5 g of $\text{C}_9\text{H}_8\text{O}_4$ is dissolved in 450 g CH_3CN .

Then, the total mass of the solution (weight of $\text{C}_9\text{H}_8\text{O}_4$ + weight CH_3CN) =
(6.5 + 450)g

$$= 456.5 \text{ g}$$

Therefore, mass percentage of $\text{C}_9\text{H}_8\text{O}_4 = \frac{\text{weight of } \text{C}_9\text{H}_8\text{O}_4}{\text{total weight of solution}}$

$$= \frac{\text{weight of } \text{C}_9\text{H}_8\text{O}_4}{\text{weight of } \text{C}_9\text{H}_8\text{O}_4 + \text{weight of } \text{CH}_3\text{CN}} = \frac{6.5}{456.5} \times 100\%$$

$$= 1.424\%$$

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

Solution:

Molality 1.5×10^{-3} m

Dose of nalorphene given = 1.5×10^{-3} gm

The molar mass of nalorphene ($C_{19}H_{21}NO_3$) is given as:

$$19 \times 12 + 21 \times 1 + 1 \times 14 + 3 \times 16 = 311 \text{ g mol}^{-1}$$

In 1.5×10^{-3} m aqueous solution of nalorphine,

$$\begin{aligned} \text{i.e., 1 kg (1000 g) of water contains } & 1.5 \times 10^{-3} \text{ mol} = 1.5 \times 10^{-3} \times 311 \text{ g} \\ & = 0.46665 \text{ g} \end{aligned}$$

$$\begin{aligned} \text{Therefore, the total mass of the solution} & = (1000 + 0.4665) \text{ g} \\ & = 1000.4665 \text{ g} \end{aligned}$$

This implies that the mass of the solution containing 0.4665 g of nalorphine is 1000.4665 g.

$$\begin{aligned} \text{Therefore, the mass of the solution containing 1.5 mg of nalorphine} & = \\ & \frac{1000.4665 \times 1.5 \times 10^{-3}}{0.4665} \text{ g} \\ & = 3.22 \text{ g} \end{aligned}$$

Hence, the mass of aqueous solution required is 3.22 gm.

- 2.30. Calculate the amount of benzoic acid (C_6H_5COOH) required for preparing 250 mL of 0.15 M solution in methanol.

Solution:

Given

$$V_{\text{methanol}} \text{ of solution} = 250 \text{ ml}$$

$$M_{\text{methanol}} \text{ of solution} = 0.15 \text{ M}$$

0.15 M solution of benzoic acid in methanol means,

1000 mL of solution contains 0.15 mole of benzoic acid

$$\begin{aligned} \text{Therefore, 250 mL of solution contains} & = \frac{0.15 \times 250}{1000} \text{ mol of benzoic acid} \\ & = 0.0375 \text{ mol of benzoic acid} \end{aligned}$$

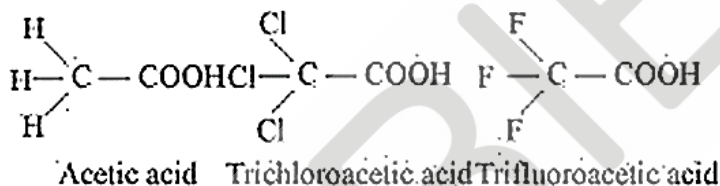
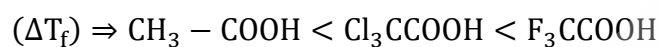
$$\begin{aligned} \text{The molar mass of benzoic acid (C}_6\text{H}_5\text{COOH)} &= 7 \times 12 + 6 \times 1 + 2 \times 16 \\ &= 122 \text{ g mol}^{-1} \end{aligned}$$

Hence, required benzoic acid = mole of benzoic acid \times molecular weight of benzoic acid = $0.0375 \text{ mol} \times 122 \text{ g mol}^{-1}$.

$$= 4.575 \text{ g}$$

- 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.

Solution:



Among H, Cl, and F, H is least electronegative while F is most electronegative in Pauling scale = 4.

$$\text{Cl} = 3.3$$

$$\text{O} = 3.5$$

$$\text{H} = 2.5$$

Then, F can withdraw electrons towards itself more than Cl and H. Thus, trifluoroacetic acid can easily lose H^+ ions i.e., trifluoroacetic acid ionizes to the largest extent. Now, the more ions produced, because as acidic character increases H^+ ion and anion availability increases, so i factor increases. The greater the depression of the freezing point. Hence, the depression in the freezing point increases in the order:

$$\Delta T_f \propto i$$

$$i \propto \text{acidic strength}$$



- 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. $K_a = 1.4 \times 10^{-3}$, $K_f = 1.86 \text{ K kg mol}^{-1}$.

Solution:

Given

Weight of $\text{CH}_3 - \text{CH}_2 - \text{CHClCOOH} = 10\text{gm}$

Weight of water = 250 gm

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

$$k_f = 1.86 \text{ k} \frac{\text{kg}}{\text{ml}}$$

We know Molar mass of $\text{CH}_3\text{CH}_2\text{CHClCOOH} = 122.5 \text{ g mol}^{-1}$

\therefore No. of moles present in 10 g of

$$\text{CH}_3\text{CH}_2\text{CHClCOOH} = \frac{\text{weight}}{\text{molecular weight}} = \frac{10 \text{ g}}{122.5 \text{ g mol}^{-1}}$$

$$= 0.0816 \text{ mol}$$

It is given that 10 g of $\text{CH}_3\text{CH}_2\text{CHClCOOH}$ is added to 250 g of water.

\therefore Molality of the solution(m)

$$= \frac{\text{moles of solute} \times 1000}{\text{weight of solvent}} = \frac{0.0816}{250} \times 1000 \text{ molal or mol kg}^{-1}$$

$$= 0.3264 \text{ mol kg}^{-1}$$

Let α be the degree of dissociation of $\text{CH}_3\text{CH}_2\text{CHClCOOH}$.

$\text{CH}_3\text{CH}_2\text{CHClCOOH}$ undergoes dissociation according to the following equation:

	$\text{CH}_3\text{CH}_2\text{CHClCOOH}$	\leftrightarrow	$\text{CH}_3\text{CH}_2\text{CHClCOO}^-$	$+\text{H}^+$
Initial conc.	$C \text{ mol L}^{-1}$		0	0
At equilibrium	$C(1 - \alpha)$		$C\alpha$	$C\alpha$

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

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

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

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

$$\Rightarrow K_a = C\alpha^2$$

$$\Rightarrow \alpha = \sqrt{\frac{K_{\alpha}}{C}}$$

$$= \sqrt{\frac{1.4 \times 10^{-3}}{0.3264}} \quad (\because K_{\alpha} = 1.4 \times 10^{-3})$$

$$= 0.0655$$

Again,

	$\text{CH}_3\text{CH}_2\text{CHClCOOH}$	\leftrightarrow	$\text{CH}_3\text{CH}_2\text{CHClCOO}^-$	$+\text{H}^+$
Initial moles	1		0	0
At equilibrium	$1 - \alpha$		α	α

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

$$\Rightarrow i = \frac{\text{moles dissociated}}{\text{initial moles}}$$

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

$$= 1 + \alpha$$

$$= 1 + 0.0655$$

$$i = 1.0655$$

Hence, the depression in the freezing point of water is given as:

$$\Delta T_f = iK_f m$$

$$\Delta T_f = 1.0655 \times 1.86 \text{K kg mol}^{-1} \times 0.3264 \text{ mol kg}^{-1}$$

$$= 0.65 \text{ K}$$

- 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.

Solution:

Given that:

$$\text{Weight of water } (w_1) = 500 \text{ g}$$

$$\text{Weight of } \text{CH}_2\text{FCOOH } (w_2) = 19.5 \text{ g}$$

$$K_f = 1.86 \text{ K kg mol}^{-1}$$

$$\Delta T_f = 1 \text{ K}$$

Formula

$$\Rightarrow \Delta T_f = k_f \times m$$

$$\text{Molality} = \frac{\text{moles of solute} \times 1000}{\text{weight of solvent (kg)}}$$

$$M = \frac{\frac{w_2}{M_2}}{\frac{w_1}{1000}} = \frac{w_2}{M_2} \times \frac{1000}{w_1}$$

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

$$= \frac{1.86 \text{ K kg mol}^{-1} \times 19.5 \text{ g} \times 1000 \text{ g kg}^{-1}}{500 \text{ g} \times 1 \text{ K}}$$

$$M_2 = 72.54 \text{ g mol}^{-1}$$

Therefore, the observed molar mass of CH_2FCOOH , $(M_2)_{\text{obs}} = 72.54 \text{ g mol}^{-1}$

Molar mass of CH_2FCOOH is:

$$(M_2)_{\text{cal}} = 14 + 19 + 12 + 16 + 16 + 1$$

$$= 78 \text{ g mol}^{-1}$$

Therefore, van't Hoff factor, $i = \frac{(M_2)_{\text{cal}}}{(M_2)_{\text{obs}}}$

$$= \frac{78 \text{ g mol}^{-1}}{72.54 \text{ g mol}^{-1}}$$

$$= 1.0753$$

Let α be the degree of dissociation of CH_2FCOOH

	CH_2FCOOH	\leftrightarrow	CH_2FCOO^-	$+\text{H}^+$	
Initial conc.	$C \text{ mol L}^{-1}$		0	0	
At equilibrium	$C(1 - \alpha)$		$C\alpha$	$C\alpha$	Total = $C(1 + \alpha)$

α = degree of dissociation

$$\therefore i = \frac{\text{concentration after dissociation } C(1 + \alpha)}{\text{initial concentration } C}$$

$$\Rightarrow i = 1 + \alpha$$

$$\Rightarrow \alpha = i - 1$$

$$= 1.0753 - 1$$

$$= 0.0753$$

Now the value of K_a is given as:

$$K_a = \frac{[\text{CH}_2\text{FCOO}^-][\text{H}^+]}{[\text{CH}_2\text{FCOOH}]}$$

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

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

Taking the volume of the solution as 500 mL, we have the concentration:

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

$$= 0.5\text{M}$$

$$\text{Therefore, } K_a = \frac{C\alpha^2}{1 - \alpha}$$

$$= \frac{0.5 \times (0.0753)^2}{1 - 0.0753}$$

$$= \frac{0.5 \times 0.00567}{0.9247}$$

$$= 0.00307 \text{ (approximately)}$$

$$= 3.07 \times 10^{-3}$$

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.

Solution:

Given,

Vapour pressure of water, $p_1^0 = 17.535$ mm of Hg

Temperature = 293 K

Mass of glucose, $w_2 = 25$ g

Mass of water, $w_1 = 450$ g

We know that,

$$\begin{aligned} \text{The molar mass of glucose } C_6H_{12}O_6, (M_2) &= 6 \times 12 + 12 \times 1 + 6 \times 16 \\ &= 180 \text{ g mol}^{-1} \end{aligned}$$

$$\text{The molar mass of water, } (M_1) = 18 \text{ g mol}^{-1}$$

Then, the number of moles of glucose,

$$(n_2) = \frac{\text{weight}}{\text{molecular weight}} = \frac{25}{180 \text{ g mol}^{-1}} = 0.139 \text{ mol}$$

$$\text{Mole fraction of solute} = \frac{n_1}{n_1+n_2}$$

And, the number of moles of water,

$$n_1 = \frac{\text{weight}}{\text{molecular weight}} = \frac{450 \text{ g}}{18 \text{ g mol}^{-1}} = 25 \text{ mol}$$

We know that,

$$\frac{p_1^0 - p_1}{p_1^0} = \text{mole fraction of solute} = \frac{n_2}{n_2+n_1}$$

$$\Rightarrow \frac{17.535 - p_1}{17.535} = \frac{0.139}{0.139 + 25}$$

$$\Rightarrow 17.535 - p_1 = \frac{0.139 \times 17.535}{25.139}$$

$$\Rightarrow 17.535 - p_1 = 0.097$$

$$\Rightarrow p_1 = 17.44 \text{ mm of Hg}$$

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 4.27×10^5 mm Hg. Calculate the solubility of methane in benzene at 298 K under 760 mm Hg.

Solution:

Given

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

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

According to Henry's law,

$$\text{Partial pressure} = k_4 \times \text{mole fraction methane in benzene}$$

$$\begin{aligned}
 p &= k_H x \\
 \Rightarrow x &= \frac{p}{k_H} \\
 &= \frac{760 \text{ mm Hg}}{4.27 \times 10^5 \text{ mm Hg}} \\
 &= 177.99 \times 10^{-5} \\
 x &= 178 \times 10^{-5} \text{ (approximately)}
 \end{aligned}$$

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.

Solution:

Given the weight of A(liquid) = 100 gm

Molecular weight of A = 140 gm/mole

Weight of liquid B = 1000 gm

Molecular weight of B = 180 gm/mole

Vapour pressure of pure liquid B = 500 torr.

Vapour pressure of pure liquid A = ?

Total vapour pressure = 475 torr.

$$\begin{aligned}
 \text{Number of moles of liquid B, } (n_B) &= \frac{\text{weight}}{\text{molecular weight}} = \frac{1000}{180} \text{ mol} \\
 &= 5.556 \text{ mol}
 \end{aligned}$$

$$\begin{aligned}
 \text{Then, the mole fraction of A, } (x_A) &= \frac{\text{weight}}{\text{molecular weight}} = \frac{n_A}{n_A + n_B} \\
 &= \frac{0.714}{0.714 + 5.556} \\
 &= 0.114
 \end{aligned}$$

And, mole fraction of B, $x_B = 1 - 0.114$

$$= 0.886$$

Vapour pressure of pure liquid B, $(p_B^0) = 500 \text{ torr}$

Therefore, the vapour pressure of liquid B in the solution,

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

Total vapour pressure of the solution, (p_{total}) = 475 torr

$$\therefore p_A + p_B = p_{\text{total}}$$

\therefore Vapour pressure of liquid A in the solution,

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

Now,

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

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

- 2.37.** Vapour pressures of pure acetone and chloroform at 328 K are 741.8 mm Hg and 632.8 mm Hg respectively. Assuming that they form an 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 the mixture is:

Plot this data also on the same graph paper. Indicate whether it has a positive deviation or negative deviation from the ideal solution.

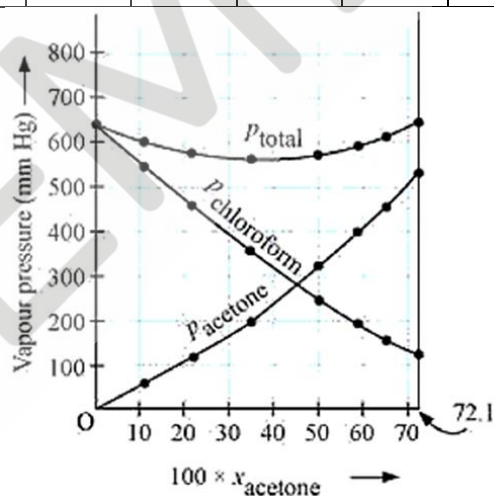
100 $\times x_{\text{acetone}}$	0	11.8	23.4	36.0	50.8	58.2	64.5	72.1
p_{acetone} / mm Hg	0	54.9	110.1	202.4	322.7	405.9	454.1	521.1
$p_{\text{chloroform}}$ / mm Hg								

p_{total} (mm Hg)								

Solution:

From the question, 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_{acetone} / mm Hg	0	54.9	110.1	202.4	322.7	405.9	454.1	521.1
$p_{\text{chloroform}}$ / mm Hg	632.8	548.1	469.4	359.7	257.7	193.6	161.2	120.7
p_{total} (mm Hg)	632.8	603.0	579.5	562.1	580.4	599.5	615.3	641.8



The plot of p_{total} slopes downwards. Therefore, the solution is said to show a negative deviation from Raoult's law.

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

Solution:

Given,

Vapour pressure of pure benzene = 50.71 mm of Hg

Vapour pressure of toluene = 32.06 mm of Hg

Weight of benzene in vapour phase = 80 g

Weight of toluene in vapour phase = 100 g

We know, Molar mass of benzene (C_6H_6) = $6 \times 12 + 6 \times 1$
= 78 g mol^{-1}

Molar mass of toluene ($C_6H_5CH_3$) = $7 \times 12 + 8 \times 1$
= 92 g mol^{-1}

Now, no. of moles present in 80 g of benzene

$$= \frac{\text{weight}}{\text{molecular weight}} = \frac{80}{78} \text{ mol} = 1.026 \text{ mol}$$

And, no. of moles present in 100 g of toluene

$$= \frac{\text{weight}}{\text{molecular weight}} = \frac{100}{92} \text{ mol} = 1.087 \text{ mol}$$

$$\therefore \text{Mole fraction of benzene, } (x_b) = \frac{1.026}{1.026+1.087} = 0.486$$

$$\therefore x_B + x_t = 1 \text{ so } x_t = 1 - x_B$$

And, mole fraction of toluene, (x_t) = $1 - 0.486 = 0.514$

Vapour pressure of pure benzene, (p_b^0) = 50.71 mm Hg

And, vapour pressure of pure toluene, (p_t^0) = 32.06 mm Hg

Therefore, partial vapour pressure of benzene, (p_b) = $x_b \times p_b^0$

$$= 0.486 \times 50.71$$

$$= 24.645 \text{ mm Hg}$$

And, partial vapour pressure of toluene, (p_t) = $x_t \times p_t^0$

$$= 0.514 \times 32.06$$

$$= 16.479 \text{ mm Hg}$$

\therefore the mole fraction of C_6H_6 in vapour phase is given to be

$$\begin{aligned} &= \frac{p_b}{p_b + p_t} = \frac{24.645}{24.645 + 16.479} \\ &= \frac{24.645}{41.124} \\ &= 0.599 \\ &= 0.6 \end{aligned}$$

- 2.39.** The air is a mixture of a number of gases. The major components are oxygen and nitrogen with an 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 at 298 K are 3.30×10^7 mm and 6.51×10^7 mm respectively, calculate the composition of these gases in water.

Solution:

Given

$$v \propto n;$$

$$n_{O_2} = 20 \text{ and } n_{N_2} = 79$$

$$\text{Total moles} = 100$$

$$\text{Percentage of oxygen (O}_2\text{) in air} = 20\%$$

$$\text{Percentage of nitrogen (N}_2\text{) in atmosphere} = 79\%$$

It is also given that water is in equilibrium with air at a total pressure of 10 atm,
 $(P_T) = (10 \times 760) \text{ mm Hg} = 7600 \text{ mm Hg}$

Therefore,

Partial pressure of oxygen,

$$\begin{aligned} (p_{O_2}) &= \text{mole fraction of } O_2 \times P_t = \frac{\text{moles of } O_2}{\text{total mass}} \times P_t = \frac{20}{100} \times 7600 \text{ mm Hg} \\ &= 1520 \text{ mm Hg} \end{aligned}$$

Partial pressure of nitrogen,

$$\begin{aligned} (p_{N_2}) &= \text{mole fraction of } N_2 \times P_t = \frac{\text{moles of } N_2}{\text{total mass}} = \frac{79}{100} \times 7600 \text{ mm Hg} \\ &= 6004 \text{ mm Hg} \end{aligned}$$

Now, according to Henry's law:

$$p = K_H \cdot x$$

For oxygen:

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

$$\Rightarrow x_{O_2} = \frac{p_{O_2}}{K_H}$$

$$= \frac{1520 \text{ mm Hg}}{3.30 \times 10^7 \text{ mm Hg}} \text{ (Given } K_H = 3.30 \times 10^7 \text{ mm Hg)}$$

$$= 4.61 \times 10^{-5}$$

For nitrogen:

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

$$\Rightarrow x_{N_2} = \frac{p_{N_2}}{K_H}$$

$$= \frac{6004 \text{ mm Hg}}{6.51 \times 10^7 \text{ mm Hg}}$$

Therefore, the mole fraction of nitrogen is 4.61×10^{-5} and of oxygen is 9.22×10^{-5} .

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

Solution:

Given

$$\text{Formula } \pi = iCRT \therefore C = \frac{\text{moles}}{\text{volume}}$$

We know that,

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

$$\text{(Volume) } V = 2.5 \text{ L}$$

$$\text{(Van't Hoff factor) } i = 2.47$$

$$\text{Temperature (T) } = (27 + 273)\text{K} = 300\text{K}$$

We know

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

$$\text{The molecular weight of } \text{CaCl}_2 = 1 \times 40 + 2 \times 35.5$$

$$= 111 \text{ g mol}^{-1}$$

$$\pi = i \frac{n}{V} RT$$

$$\Rightarrow \pi = i \frac{w}{MV} RT$$

$$\Rightarrow w = \frac{\pi MV}{iRT}$$

$$\text{Therefore, } w = \frac{0.75 \times 111 \times 2.5}{2.47 \times 0.0821 \times 300}$$

$$= 3.42 \text{ g}$$

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 litres of water at 25°C , assuming that it is completely dissociated.

Solution:

Given,

$$\text{weight of } \text{K}_2\text{SO}_4 = 25 \text{ mg} = 0.025 \text{ g}$$

$$\text{Volume of solution} = 2 \text{ L}$$

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

We know that:

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

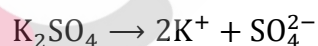
$$M = (2 \times 39) + (1 \times 32) + (4 \times 16) = 174 \text{ g mol}^{-1}$$

Applying the following relation,

$$\pi = iCRT \quad \therefore C = \frac{\text{mole}}{\text{volume}}$$

$$\pi = i \frac{n}{V} RT$$

When K_2SO_4 is dissolved in water, K^+ and SO_4^{2-} ions are produced.



Total number of ions produced = 3 (100% dissociation)

$$\therefore i = 3$$

$$\pi = i \frac{n}{V} RT$$

$$\begin{aligned} &\Rightarrow i \frac{w}{M} \frac{1}{V} RT \\ &= 3 \times \frac{0.025}{174} \times \frac{1}{2} \times 0.0821 \times 298 \\ &= 5.27 \times 10^{-3} \text{ atm} \end{aligned}$$

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