CHAPTER 1

Solution

CLASS 12TH

NCERT EXERCISE AND SOLUTIONS - CHEMISTRY

https://t.me/veda11and12

Q. 1. **Define the term solution. How many types of solutions are formed? Write briefly about each type with an example.**

ANSWER:-

Homogenous mixtures of two or more than two components are known as solution.

Type of Solution	Example
Gaseous Solutions	
(a) Gas in Gas	Air (mixture of O_2 and N_2), etc.
(b) Liquid in Gas	Water vapour (H_2O in air)
(c) Solid in Gas	Camphor vapours in N_2 gas, smoke, etc.
Liquid Solutions	
(a) Gas in Liquid	CO_2 dissolved in water (aerated water), O_2 dissolved in water, etc.
(b) Liquid in Liquid	Ethanol dissolved in water, etc.
(c) Solid in Liquid	Sugar dissolved in water, saline solution, etc.
Solid Solutions	
(a) Gas in Solid	Solution of hydrogen in palladium
(b) Liquid in Solid	Amalgams, e.g., Na-Hg (sodium in mercury)
(c) Solid in Solid	Gold ornaments (Cu/Ag with Au)

Q. 2. **Give an example of a solid solution in which the solute is a gas.**

ANSWER:-

An example of a solid solution in which the solute is a gas is hydrogen dissolved in palladium.

In this case, hydrogen (H_2) gas acts as the solute and palladium (Pd) is the solvent. Palladium can absorb large amounts of hydrogen gas at room temperature, forming a solid solution. This property of palladium is utilized in hydrogen storage and fuel cell technologies.

Q. 3. **Define the following terms:**

1. **Mole fraction**
2. **Molality**
3. **Molarity**
4. **Mass percentage**

1. **Mole fraction:** It is defined as the ratio of the number of moles of the solute to the total number of moles in the solution. If A is the number of moles of solute dissolved in B moles of solvent, then Mole fraction of solute

   $$(\chi_A) = \frac{n_A}{n_A + n_B} \quad \ldots(1)$$

   $$\text{Mole fraction of solvent } (\chi_B) = \frac{n_B}{n_A + n_B} \ldots (2)$$

Adding the above two equations, we get.

$$\chi_A + \chi_B = \frac{n_A}{n_A + n_B} + \frac{n_B}{n_A + n_B} = \frac{n_A + n_B}{n_A + n_B} = 1$$

i.e., $\chi_A + \chi_B = 1$

$\therefore \chi_A = 1 - \chi_B$ or $\chi_B = 1 - \chi_A$

2. **Molality:** Molality is defined as the number of moles of solute dissolved in one kilogram of solvent.

   $$\text{Molality}(m) = \frac{\text{Number of moles of solute}}{\text{Weight of solvent in kg}} = \frac{n}{W}$$

3. **Molarity:** It is defined as the number of moles of solute present in one litre of solution.

   $$\text{Molarity(M)} = \frac{\text{Number of moles of solute}}{\text{Volume of Solution in litre}} = \frac{n}{V}$$

4. **Mass percentage:** It is the amount of solute in grams present in 100g of solution.

   $$= \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$

Q. 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:-

Mass of HNO_3 in solution = 68 g

Molar mass of HNO_3 = 63 g mol^{-1}

Mass of solution = 100 g

Density of solution = 1·504 g mL^{-1}

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

$$= \frac{(100\text{g})}{(1.504\text{gmL}^{-1})} = 66.5\text{mL} = 0.0665\text{L}$$

$$\text{Molarity of solution(M)} = \frac{\text{Mass of HNO}_3 \,/\, \text{Molar mass of HNO}_3}{\text{Volume of solution in Litres}}$$

$$= \frac{\left(68\text{g} / 63\text{ g mol}^{-1}\right)}{(0.0665\text{ L})} = 16\cdot 23\text{mol L}^{-1} = 16\cdot 23\text{ M.}\cdot$$

Q. 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:-

10 percent w/w solution of glucose in water means 10g glucose and 90g of water.

Molar mass of glucose = 180g mol^{-1} and molar mass of water = 18g mol^{-1}

$$\therefore \quad 10\text{g of glucose} = \frac{10}{180} = 0.0555\text{ moles}$$

$$\text{and} 90\text{g of } H_2O = \frac{90}{18} = 5\text{ moles}$$

: Molality of solution

$$= \frac{\text{Moles of solute} \times 1000}{\text{Mass of solution in grams}}$$

$$= \frac{0.0555}{90} \times 1000 = 0.617\text{m}$$

Mole fraction of glucose

$$= \chi_g = \frac{\text{No. of moles of glucose}}{\text{No. of moles of glucose} + \text{No. of moles of water}}$$

$$= \frac{0.0555}{5 + 0.0555} = 0.01$$

Mole fraction of water

$$= \chi_w = \frac{\text{No. of moles of water}}{\text{No. of moles of glucose} + \text{No. of moles of water}}$$

$$= \frac{5}{5 + 0.0555} = 0.99.$$

Volume of 100g of solution

$$= \frac{\text{Mass of solution}}{\text{Density}} = \frac{100}{1.2} = 83.33\text{mL}$$

$$\therefore \text{ Molarity of solution} = \frac{0.0555}{83.33} \times 1000$$

= 0.67M.

Q. 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?**

ANSWER:-

Let the mixture contains x g of sodium carbonate and 1−x g of sodium bicarbonate.

The molar masses of sodium carbonate and sodium bicarbonate are 106 g/mol and 84 g/mol respectively

The number of moles of sodium carbonate and sodium bicarbonates are x/106 and 1−x/84 respectively.

Since, it is an equilmolar mixture,

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

84x = 106 − 106

190x = 106

x = 0.5579

Number of moles of sodium carbonate $=\frac{0.5579}{106}=0.005263$

Number of moles of sodium hydrogen carbonate $=\frac{1-0.5579}{84}=0.005263$

One mole of sodium carbonate will react with 2 moles of HCl and 1 mole of sodium bicarbonate will react with 1 mole of HCl.

Total number of moles of HCl that will completely neutralize the mixture = 2 × 0.005263 + 0.005263 = 0.01578 moles

Volume of 0.1 M HCl required $=\frac{0.01578}{0.1}=0.158\text{L}=158\text{mL}$

Q. 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:-

Given that a solution is obtained by mixing 300 g of solution and 400 g of 40% solution by mass.

Total mass of solution = 300 + 400 = 700g

Mass of solute = 0.25 × 300 + 0.40 × 400 = 75+160 = 235 g

Mass percentage of solution $=\frac{235}{700}\times 100=33.6\%$

Hence, the mass percentage of the resulting solution is 33.6%

Q. 8. **An antifreeze solution is prepared from 222.6 g of ethylene glycol ($C_2H_6O_2$) 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:-

Step 1: Calculating the molality of the solution

Since ethylene glycol is present in excess, it acts as a solvent and water is a solute.
The molar mass of water is 18 g/mol

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

$$\text{Number of moles of solute} = \frac{200}{18} = 11.11\text{moles}$$

Molality is defined as the number of moles of solute present in the 1 kg of solvent.

$$\text{Molality} = \frac{\text{Number of moles of solute}}{\text{mass of solvent(g)}} \times 1000$$

$$\text{Molality} = \frac{11.11}{222.6} \times 1000 = 49.9\text{m}$$

Step 2: Calculating the molarity of the solution

The molarity of a solution is defined as the number of moles of solute dissolved in 1 Litre of solution.

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{Volume of solution(mL)}} \times 1000$$

Total mass of solution =222.6 + 200 = 422.6 g

Density of solution = 1.072 g/ mL

$$\text{Volume of solution} = \frac{\text{Mass}}{\text{Density}} = \frac{422.6}{1.072} = 394\text{mL}$$

$$\text{Molarity} = \frac{11.11}{0.394} \times 1000 = 28.2\text{M}$$

Q. 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 percent by mass

(ii) determine the molality of chloroform in the water sample.

ANSWER:-

(i) 15 parts of chlorofom in 1,000,000 parts of water corresponds to 15 ppm. Percentage by mass of chloroform $= \frac{15 \times 100}{1,000,000} = 1.5 \times 10^{-3}\%$

(ii) The molar mass of chloroform is 12 + 1 + 3 + × 35.5 = 119.5 g/mol.

The number of moles of chloroform are $\frac{1.5 \times 10^{-3}\text{g}}{119.5\text{g/mol}} = 1.255 \times 10^{-5}\text{mol}$

The mass of water is 100 g .

Hence, the molality of the solution is $\frac{1.255 \times 10^{-5}}{100} \times 1000 = 1.255 \times 10^{-4}\text{m}$

Q. 10. **What role does the molecular interaction play in a solution of alcohol and water?**

ANSWER:-

There is strong hydrogen bonding in alcohol molecules as well as water molecules. On mixing, the molecular interactions are weakened. Hence, their solution will show positive deviation from ideal behaviour. As a result, the solution will have higher vapour pressure and lower boiling point than that of water and alcohol.

Q. 11. **Why do gases always tend to be less soluble in liquids as the temperature is raised?**

ANSWER:-

Dissolution of gas in a liquid is an exothermic process and heat is evolved during dissolution.

Gas + Liquid $\rightleftharpoons$ Solution + Heat

With an increase in temperature, heat is supplied which shifts the equilibrium in the backward direction.

Hence, gases always tend to be less soluble in liquids as the temperature is raised.

Q. 12. **State Henry's law and mention some important applications.**

ANSWER:-

According to Henry's law, the solubility of a gas in a liquid is directly proportional to the pressure of the gas.

$\chi = K_h \times P$

χ is the mole fraction of gas, K_h is Henry's law constant and P is the partial pressure of the gas.

Important applications of Henry's law:

(1) In packing of soda cans: Soda water bottles are always packed under higher pressure to increase the solubility of CO_2 gas.

(2) In deep-sea diving: Nitrogen is more soluble than Helium in our blood. In the deep sea, the pressure is higher than at the surface of the water. When diver tries to come rapidly towards the surface of the water, pressure decreases and dissolved nitrogen comes back from blood and makes bubbles in veins. Hence, divers use oxygen diluted with helium.

Q. 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:-

According to Henry's law, the solubility of gas in a liquid is directly proportional to the pressure of the gas.

$\chi = K_h \times P$

χ is the mole fraction of gas, K_h is the Henry's law constant and P is the partial pressure of the gas.

Also mole fraction of ethane will be directly proportional to its mass.

WP' = W' P

W is the mass of ethane in first solution and P is the partial pressure of ethane in first solution.

W is the mass of ethane in second solution and P' is the partial pressure of ethane in second solution.

6.56×10^{-3} P' = $5.00 \times 10^{-2} \times 1$

P' = 7.6

Hence, the partial pressure of gas will be 7.6 bar.

Q. 14. What is meant by positive and negative deviations from Raoult's law and how is the sign of ΔmixH related to positive and negative deviations from Raoult's law?

ANSWER:-

Positive deviation from Raoult's law occurs when the total vapour pressure of the solution is more than corresponding vapour pressure in case of ideal solution.

$$P = P_A + P_B > P_A^o \chi_A + P_B^o \chi_B$$

Negative deviation from Raoult's law occurs when the total vapour pressure of the solution is less than corresponding vapour pressure in case of the ideal solution.

$$P = P_A + P_B < P_A^\circ \chi_A + P_B^\circ \chi_B$$

For positive deviation from Raoult's law, Δ_{mix} H has a positive sign.

For negative deviation from Raoult's law, Δ_{mix} H has a negative sign.

Q. 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:-

Vapour pressure of pure water at boiling point (P^o) = 1 atm = 1.013bar

Vapour pressure of solution (P^s) = 1.004bar

Mass of solute (w) = 2g

Mass of solution = 100g

Mass o f solvent = 98g

Applying Roault's law for dilute solution (being 2%)

$$\frac{p^0 - p_s}{p^0} = \frac{n_2}{n_1 + n_2} \approx \frac{n_2}{n_1} = \frac{w_2 / M_2}{w_1 M_1} = \frac{w_2}{M_2} \times \frac{M_1}{w_1}$$

$$\therefore \frac{(1.013 - 1.004)}{1.013\text{bar}} = \frac{2g}{M_2} \times \frac{18\text{gmol}^{-1}}{98g}$$

$$M_2 = \frac{2 \times 18}{98} \times \frac{1.013}{0.009} \text{gmol}^{-1}$$

M_2 = 41.35 mol^{-1}.

Q. 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:-

26 g of heptane corresponds to $\frac{26}{100} = 0.26$ moles

35 g of octane corresponds to $\frac{35}{114} = 0.31$moles

Mole fraction of heptane $\chi = \frac{0.20}{0.26+0.31} = 0.456$

Mole fraction of octane $\chi' = 1 - 0.456 = 0.544$

Partial pressure of heptane $p = 0.456 \times 105.2 = 47.97$kPa

Partial pressure of octane $p' = 0.544 \times 46.8 = 25.46$kPa

Vapour pressure of solution $P = p + p' = 47.97 + 25.46 = 73.43$kPa

Q. 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:-

The vapour pressure of water is 12.3 kPa at 300 K. We need to calculate the vapour pressure of 1 molal solution of a non-volatile solute in it.

1000 g of water contains 1 mole of solute.

The molar mass of water is 18 g/mol.

Number of moles of water $= \frac{1000}{18} = 55.56$mol

Mole fraction of the solute in the solution is $\chi = \frac{1}{1+55.56} = 0.0177$

The relative lowering in the vapour pressure is equal to the mole fraction of solute.

$$\frac{p^\circ - p}{p^\circ} = \chi$$

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

$\Rightarrow p = 12.08$

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

Q. 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%.**

ANSWER

Let p be the vapour pressure of pure octane. The vapour pressure of solution will be $\frac{80}{100}p = 0.8p$

Molar mass of solute (M) and octane (m) are and respectively. Mass of octane, wis 114 g .

$$\frac{p-p'}{p}=\frac{Wm}{Mw}$$

$$\frac{p-o.8p}{p}=\frac{W\times 114}{40\times 114}$$

W = 8g

Hence, 8 g of solute are required.

Q. 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

ANSWER:-

Weight of solute, $W_B = 30$ g,

Weight of water, $W_A = 90$ g

Vapour pressure of solution $P_A = 2.8$ kPa

According to Raoult's law, $\frac{P_{\mathring{A}}^o - P_A}{P_{\mathring{A}}^o} \doteqdot \quad {}_B \simeq \frac{W_BM_A}{M_BW_A}$

$$\frac{P_A^o-2.8}{P_A^\circ}=\frac{30\times 18}{M_B\times 90}$$

$$\frac{2.8}{P_A^\circ}=\frac{M_B-6}{M_B}\ldots\ldots$$

After adding water-

Weight of solute $W_B = 30\text{g}$, Weight of water $W_A = 90+18=108\text{g}$

Vapour pressure of solution $P_A = 2.9$kPa

According to Raoult's law, $\frac{P_A^o - P_A}{P_A^o} = \chi_B \simeq \frac{W_BM_A}{M_BW_A}$

$$\frac{P_A^\circ-2.9}{P_A^\circ}=\frac{30\times 18}{M_B\times 108}$$

$$\frac{2.9}{P_A^\circ}=\frac{M_B-5}{M_B}\ldots\ldots \quad (2)$$

Divide equation (1) by equation (2), we get

$$\frac{2.8}{2.9}=\frac{M_B-6}{M_B-5}$$

M_B= 34 g/mol

Substituting the values of M_B in equation (1), we get

$$\frac{2.8}{P_A^\circ} = \frac{34-6}{34}$$

$$P_A^\circ = 3.4\text{kPa}$$

Q. 20. **A 5% solution (by mass) of cane sugar in water has freezing point of 271K. Calculate the freezing point of 5% glucose in water if freezing point of pure water is 273.15 K.**

ANSWER:-

The depression in the freezing point of the solution is given by

ΔT_f = freezing point of water- freezing point of solution"

ΔT_f = 273.15-271 = 2.15K

Molar masses of glucose and sucrose are 180 g/mol and 342g/mol respectively.

100 g of solution will contain 5 g of glucose or 5 g of sucrose.

Number of moles of glucose $= \frac{5}{180} = 0.028$ moles

Number of moles of sucrose $= \frac{5}{342} = 0.0146$ moles

Mass of solvent total mass of solution - mass of solute = 100 − 5 = 95g or 0.095 kg

Molality of sucrose solution $= \frac{0.0146}{0.095} = 0.154\text{m}$

$$K_f = \frac{\Delta T_f}{\text{molality}} = \frac{2.15}{0.154} = 13.97$$

For glucose solution, $\Delta T_f = K_f \times m = 13.97 \times 0.29 = 4.08$

Freezing point of glucose solution in water freezing point of water -ΔT_f = 273.15 − 4.08 = 269.07K

Q. 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 K kg mol^{-1}. Calculate atomic masses of A and B.**

ANSWER:-

In case of compound AB_2 :

$$M_B = \frac{K_f W_B \times 1000}{W_A \Delta T_f}$$

$$\Delta T_f = 2.3\text{K}, W_B = 1.0\text{g}, W_A = 20.0\text{g}, K_f = 5.1\text{KKg/mol}$$

$$M_B = \frac{5.1 \times 1.0 \times 1000}{20.0 \times 2.3} = 110.87\text{g/mol}$$

In case of compound AB_4 :

$$\Delta T_f = 1.3\text{K}, W_B = 1.0\text{g}, W_A = 20.0\text{g}$$

$$M_B = \frac{5.1 \times 1.0 \times 1000}{20.0 \times 1.3} = 196.15\text{g/mol}$$

Let a g/mol and bg/mol be the atomic masses of A and B respectively.

$$M_{AB_2} = a + 2b = 110.87\ldots$$

$$M_{AB_4} = a + 4b = 196.15\ldots\ldots$$

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

-2b = -85.28

Atomic mass of B is b = 42.64.

Substituting the values of b in equation (i), we get

a+2×42.64 = 110.87

Atomic mass of A is a = 25.59 g/mol.

Hence, the atomic mass of a is 25.59 g/mol and b is 42.64 g/mol

Q. 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:-

As per the van’t Hoff equation, the relationship between the osmotic pressure and the molar concentration is Π=CRT. Here, R is the ideal gas constant and T is absolute temperature.

$$C_1 = \frac{36}{180}\text{M}$$

$$\Pi_1 = 4.98\text{bar}$$

$$C_2 = ?$$

$$\Pi_2 = 1.52\text{bar}$$

$$4.98 = \frac{36}{180}\text{RT} \quad \ldots.(i)$$

$$1.52 = C_2\,\text{RT} \quad \ldots.(ii)$$

Divide equation (ii) with equation (i),

$$\frac{C_2}{36} \times 180 = \frac{1.52}{4.98}$$

C_2 = 0.061 M

Hence, second solution has concentration of 0.061 M .

Q. 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:-

The most important type of inter-molecular attractive interaction in the given pairs are given below:

(i) Van der waals interactions
(ii) Van der waals interactions
(iii) Ion dipole interactions
(iv) Hydrogen bonding
(v) Dipole dipole interactions

Q. 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:-

n-Octane is non polar and can dissolve non-polar solutes. It cannot dissolve polar (and ionic) solutes.
Cyclohexane is non-polar. Hence, easily soluble in n-octane.
Methanol and acetonitrile are polar and have very low solubility in n-octane.
KCl is ionic compound and hence, insoluble in n-octane.
The increasing order for solubility in n-octane is as follows:
$KCl < CH_3OH < CH_3CN <$ Cyclohexane.

Q. 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.**

ANSWE:-

(i) Phenol is partially soluble in water.
(ii) Toluene is water insoluble.
(iii) Formic acid is water soluble.
(iv) Ethylene glycol is water soluble.

(v) Chloroform is water insoluble.

(vi) Pentanol is partially soluble in water.

Q. 26. **If the density of some lake water is 1.25g mL^{-1} and contains 92 g of Na^+ ions per kg of water, calculate the molarity of Na^+ ions in the lake.**

ANSWER:-

Step 1: Calculate the volume of lake water

Given the density of lake water is and the mass of lake water is 1 kg (which is 1000 g), we can use the formula for density:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

Rearranging this gives us:

$$\text{Volume} = \frac{\text{Mass}}{\text{Density}}$$

Substituting the values:

$$\text{Volume} = \frac{1000\text{g}}{1.25\text{g}/\text{mL}} = 800\text{mL}$$

Step 2: Convert the volume from mL to L

Since molarity is expressed in moles per liter, we need to convert the volume from mL to L :

800 mL = 0.8 L

Step 3: Calculate the number of moles of ions

We know that the mass of ions in the lake water is 92 g and the molar mass of sodium (Na) is 23g/mo. We can calculate the number of moles using the formula:

$$\text{Moles} = \frac{\text{Mass}}{\text{Molar Mass}}$$

Substituting the values:

$$\text{Moles of Na}^+ = \frac{92\text{g}}{23\text{g}/\text{mol}} = 4\text{mol}$$

Step 4: Calculate the molarity of ions

Now we can calculate the molarity using the formula:

$$\text{Molarity}(M) = \frac{\text{Moles of solute}}{\text{Volume of solution in L}}$$

Substituting the values:

$$\text{Molarity} = \frac{4\text{mol}}{0.8\text{L}} = 5\text{mol}/\text{L}$$

Q. 27. **If the solubility product of CuS is 6×10^{-16}, calculate the maximum molarity of CuS in aqueous**

solution.

ANSWER:-

Dissociation of CuS in aqueous solution is:

$$CuS(s)\overset{aq}{\Leftrightarrow}Cu^{2+}(aq)+S^{2-}(aq)$$

$$\begin{array}{l}CuS(s)\overset{aq}{\Leftrightarrow}Cu^{2+}(aq)+S^{2-}(aq)\\ \qquad\qquad\;[s]\qquad\quad\;\;[s]\end{array}$$

$$\text{Solubility product }\left(K_{\text{sp}}\right)=\left[\text{Cu}^{2+}(aq)\right]\left[S^{2-}(aq)\right]=[S]\times[S]$$

$$[S]=\left(K_{sp}\right)^{1/2}=\left(6\times10^{-16}\right)^{1/2}=2.45\times10^{-8}M$$

Q. 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_3CN.**

ANSWER:-

Given, Mass of aspirin = 6.5 g, Mass of acetonitrile =450 g

$$\text{Mass percent of aspirin}=\frac{\text{Mass of aspirin}}{\text{Mass of aspirin}+\text{ Mass of acetonitrile}}\times100$$

$$=\frac{6.5}{6.5+450}\times100$$

=1.424%

Q. 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 * 10^{-3} m aqueous solution required for the above dose.**

ANSWER:-

To solve the problem, we need to calculate the mass of a solution that has a molality of 1.5×10^{-3} mol/kg and contains a dose of 1.5 mg of nalorphene

$C_{19}H_{21}NO_3$

1. Convert the Dose from mg to g :

Dose=1.5mg=1.5×10^{-3} g

2. Calculate the Molar Mass of Nalorphene:

The molecular formula is $C_{19}H_{22}NO_3$.

Calculate the molar mass:

Molar mass = (19 × 12) + (21 × 1) + (14) + (3 × 16)

= 228 + 21 + 14 + 48 = 311 g/mol)

3. Calculate the Number of Moles of Nalorphene:

$$\text{Moles of Nalorphene} = \frac{\text{mass}}{\text{molar mass}}$$

$$= \frac{1.5\times10^{-3}\,\text{g}}{311\,\text{g}/\text{mol}} \approx 4.8\times10^{-6}\,\text{mol}$$

4. Use the Definition of Molality:

- Molality (m) is defined as:

$$m = \frac{\text{moles of solute}}{\text{mass of solvent (kg)}}$$

- Rearranging gives us:

$$\text{mass of solvent(kg)} = \frac{\text{moles of solute}}{m}$$

- Substitute the values:

$$\text{mass of solvent(kg)} = \frac{4.8\times10^{-6}\,\text{mol}}{1.5\times10^{-3}\,\text{mol}/\text{kg}}$$

≈ 0.0032 kg.

5. Convert Mass of Solvent to grams:

mass of solvent (g)=0.0032 kg×1000 g/kg=3.2 g

6. Calculate the Mass of the Solution:

- The mass of the solution is the sum of the mass of the solute and the mass of the solvent: mass of solution mass of solute + mass of solvent
- Substitute the values:

mass of solution $=1.5\times10^{-3}$ g+3.2 g≈3.2015 g

Final Answer:

The mass of the solution required for the dose of nalorphene is approximately 3.2015 g.

Q. 30. **Calculate the amount of benzoic acid (C_6H_5COOH) required for preparing 250 mL of 0.15 M solution in methanol.**

ANSWER:-

Number of moles of benzoic acid required $= 0.15\times\frac{250}{1000} = 0.0375$ moles

Molar mass of benzoic acid = 7(12) + 6(1) + 2(16) = 84 + 6 + 32 = 122 g/mol

Mass of benzoic acid required = 122 × 0.0375 = 4.575 g

Q. 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:-

When strongly electron withdrawing groups are present on alpha C atom of acetic acid, the acid strength and the degree of dissociation increases. This increases the vant Hoff factor i and the depression in the freezing point. Trifluoroacetic is most acidic because fluorine is most electron withdrawing in nature. Hence, trifluoroacetic acid has maximum depression in the freezing point.

Q. 32. **Calculate the depression in the freezing point of water when 10 g of $CH_3CH_2CHClCOOH$ is added to 250 g of water. $K_a = 1.4 \times 10^{-3}$, $K_f = 1.86$ K kg mol^{-1}**

ANSWER:-

Molar mass of $CH_3\,CH_2\,CHClCOOH = 122.5\,g\,mol^{-1}$

No. of moles present in 10 g of $CH_3CH_2CHClCOOH = \frac{10}{122.5}\,mol = 8.16 \times 10^{-2}\,mol$

Molality of the solution(M) $= \frac{8.16 \times 10^{-2}}{250} \times 1000 = 0.3264$

Let α be the degree of dissociation of $CH_3\,CH_2\,CHClCOOH$

$CH_3\,CH_2\,CHClCOOH$ undergoes dissociation according to the following equation:

$CH_3\,CH_2\,CHClCOOH \longrightarrow CH_3\,CH_2\,CHClCOO^- + H^+$			
Initial cone.	($C\,mol\,L^{-1}$)	0	0
At equilibrium	$C(1-\alpha)$	$C\alpha$	$C\alpha$

$$\therefore K_\alpha = \frac{C\alpha \cdot C\alpha}{C(1-\alpha)} \simeq C^2$$

$$\text{or, } \alpha = \sqrt{K_\alpha / C} = \sqrt{\frac{1.4 \times 10^{-3}}{0.3264}} = 0.065$$

$CH_3\,CH_2\,CHClCOOH \longrightarrow CH_3\,CH_2\,CHClCOO^- + H^+$			
Initial moles	1	0	0
Moles at equilibrium	$1-\alpha$	α	α

Total moles $= 1 + \alpha$

$i = 1 + 0.065$

$\therefore i = 1.065$

$\Delta T_f = iK_f\,m = 1.065 \times 1.86 \times 0.3264$

$= 0.649 \approx 0.65°C$

Q. 33. **19.5 g of CH_2FCOOH is dissolved in 500 g of water. The depression in the freezing point of water observed is 1.00 C. Calculate the van't Hoff factor and dissociation constant of fluoroacetic acid.**

ANSWER:-

It is given that:

w_1=500 g

w_2=19.5g

K_f=1.86 K kg mol^{-1}

ΔT_f=1.0°C

We know that:

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

$$= \frac{1.86 K kg mol^{-1} \times 19.5 g \times 1000 g kg^{-1}}{500 g \times 1.0}$$

$$= 72.54 g mol^{-1}$$

$$= \frac{1.86 K kg mol^{-1} \times 19.5 g \times 1000 g kg^{-1}}{500 g \times 1.0}$$

= 72.54 g mol^{-1}

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

The calculated molar mass of is $(M_2)_{cal}$ = 14+19+12+16+16+1

= 78 g mol^{-1}

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

$$= \frac{78}{72.54}$$

= 1.0753

Let α be the degree of dissociation of CH_2FCOOH

	$CH_2FCOOH \rightleftharpoons$	$CH_2FCOO^- +$	H^+
Initial Conc.	C mol L^{-1}	-1	0
At equilibrium	C(1-α)	Cα	Cα

$$\therefore i = \frac{C(1+\alpha)}{C} = 1 + \alpha \text{ Or } \alpha = i - 1 = 1.0753 - 1 = 0.0753$$

Now, the value of K_a is given as:

$$K_\alpha = \frac{\left[CH_2FCOO^-\right]\left[H^+\right]}{\left[CH_2FCOOH\right]}$$

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

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

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

$$C = \frac{19.58}{78} \times \frac{1}{500} \times 1000 = 0.5M$$

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

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

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

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

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

Q. 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:-

To calculate the vapor pressure of water at 293 K when 25 g of glucose is dissolved in 450 g of water, we will use Raoult's Law. Here's the step-by-step solution:

Step 1: Identify the given data

- Vapor pressure of pure water at 293 K, P_0 = 17.535 mmHg
- Mass of glucose, W_2 = 25 g
- Mass of water, W_1 = 450g

Step 2: Calculate the number of moles of glucose and water

- Molecular weight of glucose $C_6H_{12}O_6$): 180 g/mol
- Molecular weight of water H_2O:18 g/mol

Number of moles of glucose n_2:

$$n_2 = \frac{W_2}{\text{Molecular weight of glucose}} = \frac{25g}{180g/mol}$$

$$\approx 0.1389mol$$

Number of moles of water (n_1):

$$n_1 = \frac{W_1}{\text{Molecular weight of water}} = \frac{450\text{g}}{18\text{g / mol}} = 25$$

mol

Step 3: Calculate the mole fraction of water

The mole fraction of water (χ_1) is given by:

$$\chi_1 = \frac{n_1}{n_1 + n_2} = \frac{25}{25 + 0.1389} \approx \frac{25}{25.1389} \approx 0.993$$

Step 4: Apply Raoult's Law

Raoult's Law states that the vapor pressure of the solution (P_s) is given by:

$$P_s = P_0 \cdot \chi_1$$

Substituting the values:

$$P_s = 17.535 \text{ mmHg} \cdot 0.993 \approx 17.42 \text{ mmHg}$$

Step 5: Final Result

The vapor pressure of water at 293 K when 25 g of glucose is dissolved in 450 g of water is approximately: 17.42 mmHg

Q. 35. **Henry's law constant for the molality of methane in benzene at 298 K is 4.27 × 105 mm Hg. Calculate the solubility of methane in benzene at 298 K under 760 mm Hg.**

ANSWER:-

It is given that the molality of methane in benzene at 298 K is 4.27×10^5 mm of Hg .

$$C = \frac{p}{k}$$

$$\text{C} = \frac{760\text{mmHg}}{4.27 \times 10^5 \text{ mmHg / m}}$$

$\text{C} = 1.78 \times 10^{-3}$ m

Here, C is the molality of methane in benzene, p is the pressure of methane and k is Henry's law constant.

Q. 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:-

To solve the problem step by step, we will follow these calculations:

Step 1: Calculate the number of moles of liquid B

Given:

- Mass of liquid B = 1000 g
- Molar mass of liquid B=180 g/mol

Using the formula for moles:

$$\text{Number of moles of B} = \frac{\text{Mass}}{\text{Molar mass}} = \frac{1000\text{g}}{180\text{g / mol}}$$

= 5.56 mol

Step 2: Calculate the number of moles of liquid A

Given:

Mass of liquid A=100g

Molar mass of liquid A=140 g/mol

Using the formula for moles:

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

$$= \frac{100\text{g}}{140\text{g / mol}} = 0.714\text{mol}$$

Step 3: Calculate the mole fraction of liquid

The total number of moles in the solution is:

Total moles B = 0.714

+5.56 = 6.274 mol

Now, calculate the mole fraction of B(χ_B) :

$$\chi_B = \frac{\text{Moles of B}}{\text{Total moles}} = \frac{5.56}{6.274} = 0.886$$

Step 4: Calculate the mole fraction of liquid

Using the relationship $\chi_A + \chi_B = 1$:

$$\chi_A = 1 - \chi_B = 1 - 0.886 = 0.114$$

Step 5: Calculate the vapor pressure of liquid B in the solution

Using Raoult's Law:

$$P_B = P_B^0 \times \chi_B$$

Given P_B^0 =500 torr:

$$P_B = 500 \text{ torr} \times 0.886 = 443 \text{ torr}$$

Step 6: Use the total vapor pressure to find the vapor pressure of liquid A in the solution

Given the total vapor pressure P_{total} = 475 torr:

$$P_{\text{total}} = P_A + P_B$$

Rearranging gives:

$$P_A = P_{\text{total}} - P_B = 475 \text{ torr} - 443 \text{ torr} = 32 \text{ torr}$$

Step 7: Calculate the vapor pressure of pure liquid Using Raoult's Law again:

$$P_A = P_A^0 \times \chi_A$$

Rearranging gives:

$$P_A^0 = \frac{P_A}{\chi_A} = \frac{32 \text{ torr}}{0.114} \approx 280.7 \text{ torr}$$

Summary of Results

- Vapor pressure of pure liquid : 2807 torr
- Vapor pressure of liquid A in the solution (P_A):32 torr

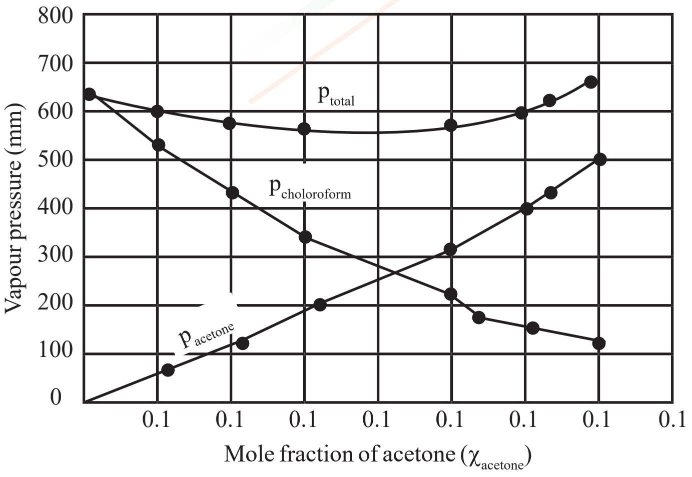
Q. 37. **Vapour pressures of pure acetone and chloroform at 328 K are 741.8 mm Hg and 632.8 mm Hg respectively. Assuming that they form ideal solution over the entire range of composition, plot P_{total}, $P_{chlroform}$ and $P_{acetone}$ as a function of $\chi_{acetone}$. The experimental data observed for different compositions of mixtures is:**

$100\times x_{acetone}$	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
$p_{chloroform}$ /mmHg	632.8	548.1	469.4	359.7	257.7	193.6	161.2	120.7

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

ANSWER:-

$100\times x_{acetone}$	0.0	0.118	023.4	0.360	0.508	0.582	0.645	0.721
$P_{acetone}$ /mm Hg	0	54.9	110.1	202.4	322.7	405.9	454.1	521.1
$P_{chloroform}$ /mmHg	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



It can be observed from the graph that the plot for the P_{total} of the solution curves downward. Therefore, the solution shows a negative deviation from the ideal behaviour.

Q. 38. **Benzene and toluene form 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 vapour phase if 80 g of benzene is mixed with 100 g of toluene.**

ANSWER:-

Molar mass of benzene (C_6H_6) = 6 × 12 + 6 × 1
= 78 g mol^{-1}
Molar mass of toluene ($C_6H_5CH_3$) = 7 × 12 + 8 × 1
= 92 g mol^{-1}

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

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

$\therefore$ Mole fraction of benzene, $\chi_b = \frac{1.026}{1.026 + 1.087} = 0.486$

And, mole fraction of toluene, $\chi_t = 1 - 0.486 = 0.514$

It is given that the vapour pressure of pure benzene, $p_b^0 = 50.71 \text{mmHg}$

And, vapour pressure of pure toluene, $p_t^0 = 32.06 \text{mmHg}$

Therefore, the partial vapour pressure of benzene, $p_b = \chi_b \times p_b^0$

$= 0.486 \times 50.71$

$= 24.65 \text{mmHg}$

And, partial vapour pressure of toluene, $p_t = \chi_t \times p_t^0$

$= 0.514 \times 32.06$

$= 16.48 \text{mmHg}$

As a result, the mole fraction of benzene in the vapour phase is as follows:

$$\frac{p_b}{p_b + p_t}$$

$$= \frac{24.65}{24.65 + 16.48}$$

$$= \frac{24.65}{41.13}$$

$= 0.599$

$= 0.6$

Q. 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 at 298 K are 3.30 × 107 mm and 6.51 × 107 mm respectively, calculate the composition of these gases in water.**

ANSWER:-

The vapour pressure of air over water is 10 atm.

The partial pressure of nitrogen

$$= P_{N_2} = \frac{79\times 10}{100} = 7.9\text{atm} = 7.9\times 760\text{mmHg} = 6004\text{mmHg}$$

The partial pressure of oxygen

$$P_{O_2} = \frac{20\times 10}{100} = 2.0\text{atm} = 2.0\times 760\text{mmHg} = 1520\text{mmHg}$$

According to Henry's law,

$$P_{N_2} = K_H\left(N_2\right)\times \chi_{N_2}$$

$$\chi_{N_2} = \frac{P_{N_2}}{K_H\left(N_2\right)} = \frac{6004}{6.51\times 10^7} = 9.22\times 10^{-5}$$

$$\chi_{O_2} = \frac{P_{O_2}}{K_H O_2} = \frac{1520}{3.30\times 10^7} = 4.6\times 10^{-5}$$

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

Q. 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:-

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

$$\text{or, } n = \frac{\pi\times V}{i\times R\times T}$$

$$= \frac{0.75\text{atm}\times 2.5\text{L}}{2.47\times 0.0821\text{LatmK}^{-1}\text{mol}^{-1}\times 300\text{K}}$$

$$= \frac{1.875}{60.836}$$

= 0.0308 mol

Molar mass of $CaCl_2$ = 40+2×35.5 = 111 g mol^{-1}

Amount of dissolved = 0.0308×111g = 3.42g

Q. 41. **Determine the osmotic pressure of a solution prepared by dissolving 25 mg of K_2SO_4 in 2 litre of water at 25° C, assuming that it is completely dissociated.**

ANSWER:-

When $K_2\ SO_4$ is dissolved in water, $K^+$and SO_4^{2-}. ions are produced.

$$K_2SO_4 \rightarrow 2K^+ + SO_4^{2-}$$

Total number of ions produced = 3

$\therefore i = 3$

Given,

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

$V = 2\text{ L}$

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

Also, we know that:

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

$M = (2 \times 39) + (1 \times 32) + (4 \times 6) = 174\text{g mol}^{-1}$

Applying the following relation,

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

$$= i \times \frac{w}{M} \times \frac{1}{V}RT$$

$$= 3 \times \frac{0.025}{174} \times \frac{1}{2} \times 0.0821 \times 298$$

$= 5.27 \times 10^{-3}$ atm