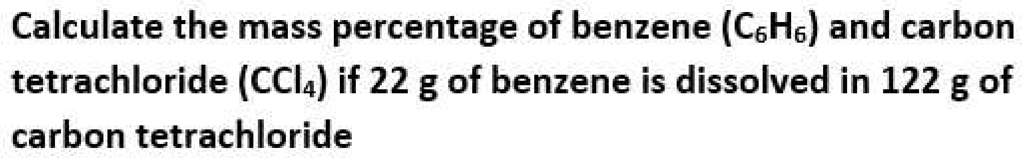
NCERT Solutions for Class 12 Chemistry
Chapter 1 – Solutions

Intext Questions with Solutions of Class 12 Chemistry Chapter 1 – Solutions

1.1



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

📌 ***Key Points to Note from Question***

Given Data:

- Mass of benzene (C_6H_6): **22 g**
- Mass of carbon tetrachloride (CCl_4): **122 g**

Key Concept:

- Mass Percentage Formula

What to Calculate:

- Mass percentage of **benzene**.
- Mass percentage of **carbon tetrachloride**.

Ans – The mass percentage is the ratio between the solute's mass and the total mass of the solution, multiplied by 100.

$$\text{Mass\%} = \frac{\text{mass}_{solute}}{\text{mass}_{solution}} \times 100$$

The mass proportion of each component is provided as follows;

$$\text{Mass\% of } C_6H_6 = \frac{\text{mass}_{C_6H_6}}{\text{mass}_{C_6H_6} + \text{mass}_{CCl_4}} \times 100$$

$$= \frac{22}{22 + 122} \times 100 = 15.28\%$$

The solution consists just in two elements: C_6H_6 and CCl_4.
Consequently, the mass percentage of CCl_4 is computed as 100 – 15.28, thereby obtaining 84.72%.
C_6H_6 has mass percentages of 15.28% and CCl_4 of 84.72%. respectively.

1.2

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

📌 ***Key Points to Note from Question***

Given Data:

- Mass percentage of benzene (C_6H_6): **30%**
- Solvent: Carbon tetrachloride (CCl_4)
- Assume **100 g of solution** for simplicity:
  - Mass of benzene = **30 g**
  - Mass of carbon tetrachloride = **70 g**

Key Concept:

- Mole Fraction Formula:- Moles of benzene/Total moles of solution
- Moles of a substance:- Mass/Molar Mass

What to Calculate:

- Moles of benzene.
- Moles of carbon tetrachloride.
- Mole fraction of benzene.

Ans – The mole fraction is the ratio of the moles of solute to the total moles of the solution.

$$\text{Mass fraction} = \frac{n_{\text{solute}}}{n_{\text{solution}}}$$

This solution has 30% by mass of benzene in carbon tetrachloride, indicating that 30 g of benzene is contained in 100 g of the solution. Therefore, carbon tetrachloride constitutes 70 g inside 100 g of the solution.
It is known that,

The molar mass of benzene is 78 g/mol.
The molar mass of carbon tetrachloride is 154 g/mol.

⇒ No. of moles of benzene = 30/78 = 0.385mol
⇒ No. of moles of carbon tetrachloride = 70/154 = 0.455mol

$$\text{Mass fraction of } C_6H_6 = \frac{n_{C_6H_6}}{n_{C_6H_6} + n_{CCl_4}}$$

$$= \frac{0.385}{0.385 + 0.455} = \frac{0.385}{0.84} = 0.458$$

∴ The mole fraction of benzene (C_6H_6) in the solution = 0.458.

1.3

Calculate the molarity of each of the following solutions: (a) 30 g of $Co(NO_3)_2$. 6H2O in 4.3 L of solution (b) 30 mL of 0.5 M H2SO4 diluted to 500 mL

📌 ***Key Points to Note from Question***

Part (a):

Given Data:

- Mass of $Co(NO_3)_2{\cdot}6H_2O$: **30 g**
- Volume of solution: **4.3 L**

Key Concept:

- **Molarity Formula:** Moles of solute/Volume of solution (L)
- **Moles Formula:** Mass/Molar Mass
- Molar Mass of $Co(NO_3)_2{\cdot}6H_2O$ = **291 g/mol**

What to Calculate:

- Moles of $Co(NO_3)_2{\cdot}6H_2O$.
- Molarity of the solution.

Part (b):

Given Data:

- Initial volume of H_2SO_4: **30 mL**
- Initial molarity: **0.5 M**
- Final volume after dilution: **500 mL = 0.5 L**

Key Concept:

- **Dilution Formula:** $M_1V_1 = M_2V_2$

What to Calculate:

- Final molarity (M_2) of the diluted H_2SO_4 solution.

Ans – a) Molarity is defined as the quantity of moles of solute per litre of solution.

$$\text{Molarity} = \frac{n_{\text{solute}}}{V_{\text{solution}}}\ \text{M}$$

Here V_{solution} is given as 4.3L

The molar mass of $Co(NO_3)_2.6H_2O$ is 310.7 g/mol.

Number of moles = 30/310.7 = 0.0966

Therefore, molarity = 0.0966/4.3 = 0.022 M

b) The molarity of 30 mL of 0.5 M H2SO4 diluted to 500 mL.

There are 0.5 moles of H2SO4 in 1000 ml of 0.5 M H2SO4.

Consequently, 30 ml of 0.5 M H_2SO_4 has 0.5/1000 × 30 = 0.015 moles of H_2SO_4.

The volume of the solution is 500 millilitres, equivalent to 0.5 litres.

This means that the molarity is 0.015/0.5, which is 0.03 M.

1.4

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

📌 ***Key Points to Note from Question***

Given Data:

- Mass of solvent (water): **2.5 kg**
- Molality (mmm): **0.25 mol/kg**
- Molar mass of urea (NH_2CONH_2): **60 g/mol**

Key Concept:

- Molality Formula
- Moles of Solute Formula
- Mass of Solute Formula

What to Calculate:

- Moles of urea required.
- Mass of urea required.

Ans – Molality is defined as the ratio of moles of solute to grammes of solvent mass.

$$\text{Molality} = \frac{n_{\text{solute}}}{m_{\text{solvent}}}\ \text{m}$$

A 0.25 molal solution of urea signifies that 0.25 moles of urea are dissolved in 1000 grammes of solvent.
The molar mass of the solute, urea, is 60 g/mol.
Hence, the mass of urea in the solution is 15 grammes, calculated as 0.25 multiplied by 60.
Total mass of solution = 1000 g solvent + 15 g solute = 1015 g equals 1.015 kg.
This indicates that, a 1.015 kg solution contains 15 g of urea.
As a result, 2.5 kg solution will contain 37 g of urea, calculated as follows:
15/1.015 x 2.5 = 37.
∴ the requisite mass of urea will be 37 g.

1.5

Calculate (a) molality (b) molarity and (c) mole fraction of KI if the density of 20% (mass/mass) aqueous KI is 1.202 g mL^{-1}.

📌 ***Key Points to Note from Question***

Given Data:

- Mass percentage of KI: **20% (mass/mass)**
- Density of solution: **1.202 g/mL**
- Solvent: Water (assume remaining mass is water).
- Molar mass of KI: **166 g/mol**

Key Concept:

- Molality Formula
- Molarity Formula

What to Calculate:

- Molality
- Molarity
- Mole fraction of KI

Ans – a) In a 20% aqueous KI solution, there are 20 grammes of KI in 100 grammes of solution.
Mass of water (solvent) = 100 – 20 = 80 g.
The solution's volume is given as follows:

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

$$= \frac{100}{1.202} = 83.194 \text{ ml}$$

Potassium iodide's (KI) molar mass is 166.0028 g/mol.

$$\text{No. of moles of KI} = \frac{20}{166.0028} = 0.120\ \text{mol}$$

Water has a molar mass of 18 g/mol.

$$\Rightarrow \text{No. of moles of Water} = \frac{80}{18} = 4.44\ \text{mol}$$

$$\text{Molality} = \frac{0.120}{0.08} = 1.50\text{m}$$

b)

$$\text{Molarity} = \frac{0.120}{0.083194} = 1.44\text{M}$$

c)

$$\text{Mole fraction} = \frac{0.120}{0.120 + 4.44} = 0.0263$$

1.6

H_2S, a toxic gas with rotten egg like smell, is used for the qualitative analysis. If the solubility of H_2S in water at STP is 0.195 m, calculate Henry's law constant.

📌 ***Key Points to Note from Question***

Given Data:

- Solubility of H_2S: **0.195 mol/kg (molality)**
- Temperature: **STP (0°C or 273 K, 1 atm)**

Key Concept:

- **Pressure (PPP)** at STP: **1 atm**
- **Concentration (CCC)** is given in molality, assumed equivalent to molarity in dilute solutions.
- **Henry's Law Constant Formula,** $k_H = P/C$

What to Calculate:

- Henry's law constant

Ans – The molality of H_2S in water is 0.195 m, indicating that there are 0.195 moles of H_2S in 1 kg of water (1000 g).
Molar mass of water = 18 g/mol.

The quantity of moles of water is calculated as 1000 divided by 18, resulting in 55.55 moles.
The mole fraction of H_2S is:

$$X_{H_2S} = \frac{0.195}{0.195 + 55.55} = 0.00349$$

Pressure at standard temperature and pressure (STP) = 0.987 bar
∴ Using Henry's law, we obtain:

$$P_{H_2S} = K_H \times X_{H_2S}$$

$$K_H = \frac{P_{H_2S}}{X_{H_2S}} = \frac{0.987}{0.00349} = 282.80\ bar$$

Thus, Henry's constant will equal 282.80 bar.

1.7

Henry's law constant for CO_2 in water is 1.67×10^8 Pa at 298 K. Calculate the quantity of CO_2 in 500 mL of soda water when packed under 2.5 atm CO_2 pressure at 298 K.

📌 ***Key Points to Note from Question***

Given Data:

- Henry's Law Constant (k_H): **1.67×10^8 Pa**
- Pressure of CO_2 (P): **2.5 atm**
- Soda water volume: **500 mL = 0.5 L**
- Temperature: **298 K**
- 1 atm = **1.013×10^5 Pa**

Key Concept:

- **Henry's Law Constant Formula,** $C = P/k_H$
- Moles of CO_2, $n = C \times V$
- Mass of CO_2 = $n \times M$

What to Calculate:

- Concentration of CO_2 in water.
- Moles of CO_2.
- Mass of CO_2 in soda water.

Ans –

$\text{Henry's Law constant}, K_H = 1.67 \times 10^8 \text{ Pa}$

$P_{CO_2} = 2.5\,atm = 2.5 \times 101325$

$= 253312.5 \text{ Pa}$

By Henry's Law,

$P_{CO_2} = K_H \times X_{CO_2}$

$$X_{CO_2} = \frac{253312.5}{1.67 \times 10^8} = 1.5168 \times 10^{-3}$$

500 ml of soda water is equivalent to 500 ml of water, which indirectly indicates 500 grammes of the same.
Molar mass of water = 18 g/mol.

$$\Rightarrow \text{No. of moles of Water} = \frac{500}{18} = 27.78 \text{ mol}$$

No. of moles of CO_2 will be,

$$X_{CO_2} = \frac{n_{CO_2}}{n_{CO_2} + n_{water}}$$

$n_{CO_2} = X_{CO_2}(n_{CO_2} + n_{water})$

$n_{CO_2} = 42.2 \times 10^{-3}\text{mol}$

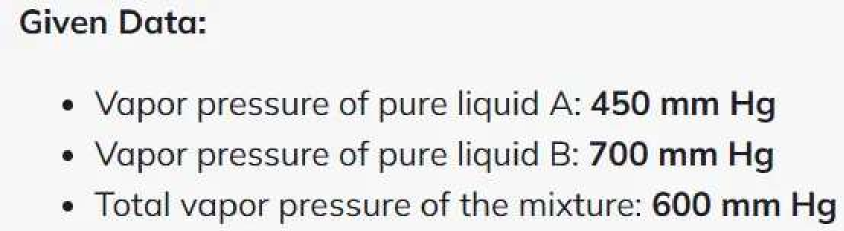
Molar mass of CO_2 = 44 g/mol.

$\Rightarrow \text{Mass} = n_{CO_2} \times 44 = 1.857\,gm$

1.8

The vapour pressure of pure liquids A and B are 450 and 700 mm Hg respectively, at 350 K. Find out the composition of the liquid mixture if total vapour pressure is 600 mm Hg. Also find the composition of the vapour phase.

📌 ***Key Points to Note from Question***



Given Data:

- Vapor pressure of pure liquid A: **450 mm Hg**
- Vapor pressure of pure liquid B: **700 mm Hg**
- Total vapor pressure of the mixture: **600 mm Hg**

Key Concept:

- Raoult's Law

What to Calculate:

- Mole fraction of A (x_A) and B (x_B) in the liquid phase.
- Mole fraction of A (y_A) and B (y_B) in the vapor phase.

Ans – The vapour pressures of the specified pure liquids are as follows:

P_A^0 = 450mm Hg

P_B^0 = 700mm Hg

P_{Total} = 600mm Hg

According to Raoult's law we know,

$P_{Total} = P_A + P_B$

$= X_A P_A^0 + X_B P_B^0$

$= X_A P_A^0 + (1 - X_A) P_B^0$

$= P_B^0 + (P_A^0 - P_B^0) X_A$

$\therefore 600 = 700 + (450 - 700) X_A$

$\Rightarrow X_A = 0.4$

$\Rightarrow X_B = 1 - 0.4 = 0.6$

$P_A = X_A P_A^0$

$= 0.4 \times 450 = 180$mm Hg

$P_B = X_B P_B^0$

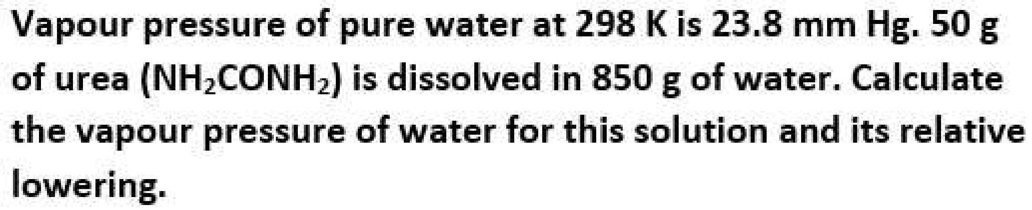
$= 0.6 \times 700 = 420$mm Hg

In the vapor phase,

$$\text{Mole fraction of A} = \frac{180}{180 + 420} = 0.3$$

$$\text{Mole fraction of B} = 1 - 0.3 = 0.7$$

1.9



Vapour pressure of pure water at 298 K is 23.8 mm Hg. 50 g of urea (NH_2CONH_2) is dissolved in 850 g of water. Calculate the vapour pressure of water for this solution and its relative lowering.

📌 *Key Points to Note from Question*

Given Data:

- Pure water Vapor pressure (P_0): **23.8 mm Hg**
- Mass of urea (NH_2CONH_2): **50 g**
- Mass of water: **850 g**
- Molar mass of urea: **60 g/mol**
- Molar mass of water: **18 g/mol**

Key Concept:

- Raoult's Law

What to Calculate:

- Moles of urea (n_{solute})
- Moles of water ($n_{solvent}$)
- Relative lowering of vapor pressure
- Vapour pressure of the solution (P)

Ans – The vapour pressure of pure water, P^0, is 23.8 mm Hg.

The mass of urea, W_2, is 50 g

The molar mass of urea, M_2, is 60 g/mol.

Mass of water, W_1 = 850 g

Molar mass of water = 18 g/mol.

Applying Raoult's law,

$$\frac{P^0 - P_s}{P^0} = \frac{n_2}{n_1 + n_2} = \frac{\frac{W_2}{M_2}}{\frac{W_1}{M_1} + \frac{W_2}{M_2}}$$

$$= \frac{\frac{50}{60}}{\frac{850}{18} + \frac{50}{60}} = \frac{0.83}{47.22 + 0.83}$$

$$= 0.017$$

Putting P^0 = 23.8 mm, we will get,

$$\frac{23.8 - P_s}{P_s} = 0.017$$

$$23.8 - P_s = 0.017\ P_s$$

$$1.017\ P_s = 23.8$$

$$P_s = 23.4\ \text{mm}$$

1.10

Boiling point of water at 750 mm Hg is 99.63°C. How much sucrose is to be added to 500 g of water such that it boils at 100°C.

📌 ***Key Points to Note from Question***

Given Data:

- Boiling point of water at 750 mm Hg: **99.63°C**
- Mass of water: **500 g = 0.5 kg**
- Molal boiling point elevation constant (K_b) for water: **0.52 °C kg/mol**

Key Concept:

- Elevation in Boiling Point Formula
- Molality Formula
- Moles Formula

What to Calculate:

- Molality of the solution (m)
- Moles of sucrose
- Mass of sucrose required

Ans – Water's boiling point = 100°C.
Water's boiling point at 750 mm Hg = 99.63°C.
Boiling point increase, ΔT_b = 100 – 99.63 = 0.37
The ebullioscopic constant, K_b = 0.52.
Mass of water, W_1 = 500 g
Molar mass of water = 18 g/mol.
The molar mass of sucrose, M_2 = 342 g/mol.
The rise in boiling point is specified as;
$\Delta T_b = K_b \times m$

$$\Delta T_b = K_b \times \frac{n_2}{W_1}$$

$$= K_b \times \frac{W_2}{M_2} \times \frac{1000}{W_1}$$

$$W_2 = \frac{\Delta T_b \times M_2 \times W_1}{K_b \times 1000} = 121.67g$$

∴ 121.67 g of sucrose is to be dissolved in water to achieve a boiling point of 100°C.

1.11

Calculate the mass of ascorbic acid (Vitamin C, $C_6H_8O_6$) to be dissolved in 75 g of acetic acid to lower its melting point by 1.5°C. K_f = 3.9 K kg mol^{-1}

📌 ***Key Points to Note from Question***

Given Data:

- Depression in freezing point (ΔT_f): **1.5°C**
- Mass of acetic acid (solvent): **75 g = 0.075 kg**
- Cryoscopic constant (K_f) for acetic acid: **3.9 K kg mol^{-1}**

Key Concept:

- Freezing Point Depression Formula

What to Calculate:

- Molality of the solution (m)
- Moles of ascorbic acid
- Mass of ascorbic acid required

Ans – Lowering in freezing point, ΔT_f = 1.5° C

Cryoscopic constant, K_f = 3.9Kg/mol

Mass of acetic acid, W_1 = 75 g

Molar mass of acetic acid, M_1 = 60 g/mol

Molar mass of ascorbic acid, M_2 = 176 g/mol

Mass of ascorbic acid, W_2 = ?

The lowering in freezing point is given as;

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

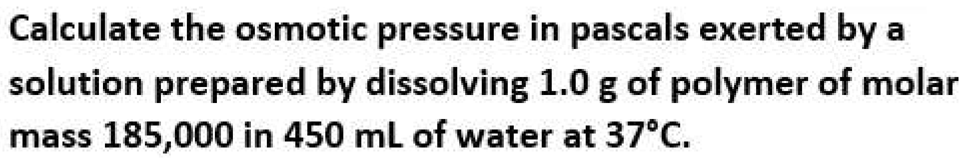
$$\Delta T_f = K_f \times \frac{n_2}{W_1}$$

$$= K_f \times \frac{W_2}{M_2} \times \frac{1000}{W_1}$$

$$W_2 = \frac{\Delta T_f \times M_2 \times W_1}{K_f \times 1000} = 5.077g$$

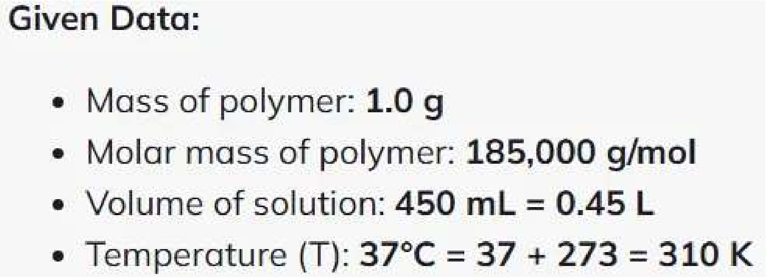
Ascorbic acid has a mass of 5.077 g.

1.12



Calculate the osmotic pressure in pascals exerted by a solution prepared by dissolving 1.0 g of polymer of molar mass 185,000 in 450 mL of water at 37°C.

📌 ***Key Points to Note from Question***



Given Data:

- Mass of polymer: **1.0 g**
- Molar mass of polymer: **185,000 g/mol**
- Volume of solution: **450 mL = 0.45 L**
- Temperature (T): **37°C = 37 + 273 = 310 K**

Key Concept:

- Osmotic Pressure Formula
- Volume from mL to m^3: 1 mL=10^{-6} m^3

What to Calculate:

- Moles of polymer
- Molar concentration of the solution (C)
- Osmotic pressure (Π)

Ans –

Mass of polymer, W = 1 g

Molar mass of polymer, M = 185000 g/mol

Volume of water, V = 450 ml = 0.45 L

Temperature, T = 37°C = 310 K

Gas constant, R = 8.314 $kPa.L.K^{-1}mol^{-1}$

$= 8.314 \times 10^3\ Pa.L.K{-}1mol^{-1}$

Now, applying the osmotic pressure formula,

$$\pi = C\,R\,T = \frac{n}{V}R\,T$$

$$= \frac{W}{M} \times \frac{1}{V} \times R\,T$$

$$\pi = 30.96\ Pa$$

Exercise Questions with Solutions of Class 12 Chemistry Chapter 1 – Solutions

1.1

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

Ans – A homogeneous combination of over 2 chemically non-reacting elements is defined as a solution. 9 different variants of solutions are deemed under the 3 major states: Gaseous, Liquid, and Solid.

Here are the various types of solutions with desired examples –

- **Solid in Solid:** Ornaments like Cu/Ag with Au.
- **Gas in Gas:** Air, The combination of O_2 & N_2 elements.
- **Liquid in Gas:** Water vapor is a relevant example of this variant.
- **Solid in Gas:** Smoke, Camphor vapors in N_2 gas.
- **Gas in Liquid:** Aerated water, O_2 dissolved in H_2O.
- **The liquid in Liquid:** Vinegar solution is a prime example of this variant.
- **Solid in Liquid:** Glucose is dissolved in water and saline water.
- **Gas in Solid:** Solution of hydrogen in platinum.
- **Liquid in Solid:** Amalgams such as Mg and Hg.

1.2

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

Ans – We all know that a gaseous state is defined as a solute in a solid solution, which is also referred to as a gas-solid solution. Some of the well-known examples are solid CO_2 which is commonly seen in fire extinguishers, combined solution of H_2 in palladium, multiple dissolved gases in earth-based minerals, and the list goes on.

1.3

Define the following terms:

(i) Mole Fraction (ii) Molality

(iii) Molarity (iv) Mass Percentage

Ans – (i) Mole fraction can be defined as the overall ratio of moles present in a solute when compared to the total moles present in a solution.

$$\text{Mole fraction} = \frac{n_{\text{solute}}}{n_{\text{solution}}}$$

(ii) Molality is defined as the ratio of solute moles per gram of solvent mass.

$$\text{Molality} = \frac{n_{\text{solute}}}{m_{\text{solvent}}}\ \text{m}$$

It is a much better and more reliable option to showcase the overall concentration of the total solute since this does not affect the modifications in temperature compared to molarity (based on the total volume of the solution).

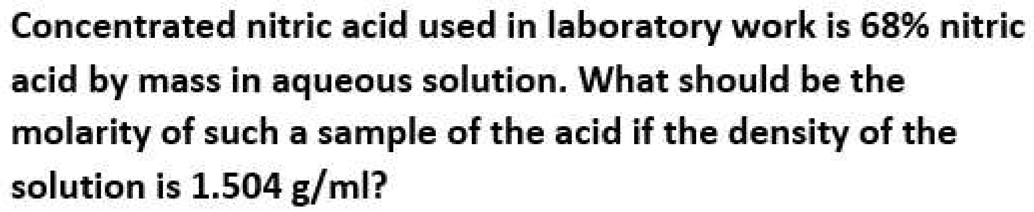
(iii) Molarity is the overall no. of moles of solute available per litre of a solution.

$$\text{Molarity} = \frac{n_{\text{solute}}}{V_{\text{solution}}}\ \text{M}$$

(iv) Mass percentage is the total ratio of the mass of the solute that gets separated by the total mass of the solution which is then multiplied by 100.

$$\text{Mass\%} = \frac{\text{mass}_{solute}}{\text{mass}_{solution}} \times 100$$

1.4



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?

📌 ***Key Points to Note from Question***

Given Data:

- Mass percentage of nitric acid (HNO_3): **68%**
- Density of the solution: **1.504 g/mL**
- Molar mass of nitric acid (HNO_3): **63 g/mol**

Key Concept:

- Molarity Formula
- Total mass of solution is derived from density and volume.

What to Calculate:

- Molarity of the nitric acid solution.

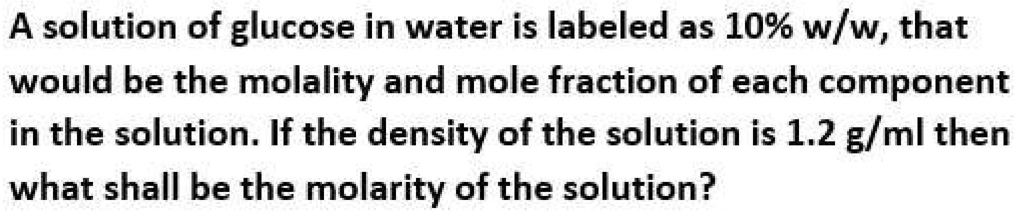
Ans – As we all know, we already contain a 68% nitric acid regarding mass in an aqueous solution where 68 grams of nitric acid is readily available in 100 grams of solution.
Then, the overall molar mass of nitric acid remains at 63 gms/mol
Hence, the no. of moles of nitric acid can be given as, 68/63 = 1.079mol
So, the total density of the solution stands at 1.504 gms/ml.
Finally,

$$\text{Volume} = \frac{\text{mass}}{\text{density}} = \frac{100}{1.504} = 66.489\, ml$$

Similarly, the molarity can be defined as,

$$\text{Molarity} = \frac{1.079 \times 1000}{66.489} = 16.22M$$

1.5



A solution of glucose in water is labeled as 10% w/w, that would be the molality and mole fraction of each component in the solution. If the density of the solution is 1.2 g/ml then what shall be the molarity of the solution?

📌 ***Key Points to Note from Question***

Given Data:

- Glucose solution concentration: **10% w/w**
- Density of the solution: **1.2 g/mL**
- Molar mass of glucose: **180 g/mol**

Key Concept:

- Molality Formula
- Mole Fraction Formula
- Molarity Formula
- Volume of solution is derived from density and mass.

What to Calculate:

- Molality of glucose solution
- Mole fraction of glucose and water
- Molarity of the glucose solution

Ans – As we already know, a 10% w/w mixture of sugar in water, which is equivalent to 10 g of glucose in 90 grams of liquid, is what we have provided.
As per our recent understanding;
Glucose molar mass in total remains 180 gms/mol
The water's molar mass stands at 18 g/mol.
Therefore, 10/180 = 0.055 mol is the number of moles of glucose in the water solution.
90/18 = 5mol is the number of moles of water in the solution.
Considering the values mentioned previously;
The Molality can be provided as,

$$\text{Molality} = \frac{0.055 \times 1000}{90} = 0.617\, m$$

Any element's molecular percentage shall be expressed as follows:

$$\text{Mole fraction of glucose} = \frac{0.055}{0.055 + 5} = 0.0108$$

$$\text{Mole fraction of water} = 1 - 0.0108 = 0.9892$$

We are already been informed that the overall density of the solution remains at 1.2 gms/ml. So,

$$\text{Volume} = \frac{\text{mass}}{\text{density}} = \frac{100}{1.2} = 83.33\, ml$$

Thus, molarity can be given as,

$$\text{Molarity} = \frac{0.055 \times 1000}{83.33} = 0.66\, M$$

1.6

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?

📌 ***Key Points to Note from Question***

Given Data:

- Molarity of HCl\text{HCl}HCl: **0.1 M**
- Mass of the mixture (Na_2CO_3 and $NaHCO_3$): **1 g**
- Both components are in **equimolar amounts**.

What to Calculate:

- Moles of Na_2CO_3 and $NaHCO_3$ in the mixture.
- Total moles of HCl required.
- Volume of HCl solution needed.

Ans – As we already know,

Assume for the moment that we are dealing with x g of Na_2CO_3 in an amalgam of 1 g.

Consequently, (1 – x) g of $NaHCO_3$ will be in the same.

We are aware that,

The total molar mass of Na_2CO_3 = 106 gms/mol

The total molar mass of $NaHCO_3$ = 84 gms/mol

$$\text{Number of moles of } Na_2CO_3 \text{in mixture} = \frac{x}{106}$$

$$\text{Number of moles of } NaHCO_3 \text{in mixture} = \frac{(1-x)}{84}$$

Since we now know that they are equimolar,

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

$$84x = 106 - 106x$$

$$\therefore x = 0.557$$

Thus,

$$\text{Number of moles of } Na_2CO_3 \text{in mixture} = \frac{0.557}{106} = 0.00526$$

$$\text{Number of moles of } NaHCO_3 \text{in mixture} = \frac{(1 - 0.557)}{84} = 0.00503$$

We must examine the effects for each to determine the amount of millilitres of 0.1 M HCl needed to fully interact with a 1 g combination of Na_2CO_3 & $NaHCO_3$.

$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + HO_2 + CO_2$$

From this we get that each mole of Na_2CO_3 requires 2 moles of HCl.

⇒ 0.00526 moles of Na_2CO_3 requires 0.00526 x 2 = 0.01052 moles of HCl

$$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$$

From this we get that each mole of $NaHCO_3$ requires a mole of HCl.

⇒ 0.00503 moles of $NaHCO_3$ requires 0.00503 moles of HCl

Thus, total moles of HCl required = 0.01052 + 0.00503 = 0.01555 moles

There are 1.1 moles of 0.1 M HCl in a 1000 ml solution.

$$0.01555 \text{ moles will be present in } = \frac{0.01555 \times 1000}{0.1}$$
$$= 155.5\, ml \text{ of solution}$$

Thus, the volume necessary for complete reaction with the mixture is 155.5 ml.

1.7

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.

📌 ***Key Points to Note from Question***

Given Data:

- Mass of first solution: 300 g, concentration: 25%.
- Mass of second solution: 400 g, concentration: 40%.

Key Concept:

- Mass Percentage Formula for the Resulting Solution
- Mass of solute in a solution = (Percentage Concentration × Total Mass)/100

What to Calculate:

- Total mass of solute in the resulting solution.

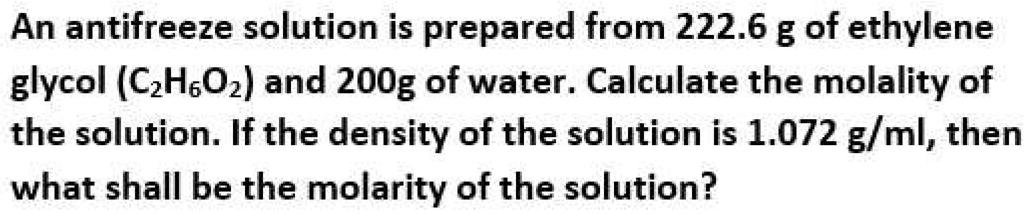
- Total mass of the resulting solution.
- Mass percentage of the resulting solution.

Ans – As we already know,
The total weight of the solute is divided by the mass of the solution itself, multiplied by 100.
Therefore, the total percentage value of mass is defined as,

$$\text{Mass\%} = \frac{\text{mass}_{solute}}{\text{mass}_{solution}} \times 100$$

A combination of 400 g of 40% solutions by mass & 300 g of 25% solution is provided here. As an outcome, the final solution will contain the following:
Solute total mass = 75 + 160 = 235 g
The solution's total weight is 300 + 400 = 700 g.
Consequently, the mass percentage is provided as
235/700 * 100 = 33.57% is the mass proportion of solute present in the desired solution.
Likewise, 100 – 33.57 = 66.42% is the mass proportion of water in the solution.

1.8



An antifreeze solution is prepared from 222.6 g of ethylene glycol ($C_2H_6O_2$) and 200g of water. Calculate the molality of the solution. If the density of the solution is 1.072 g/ml, then what shall be the molarity of the solution?

📌 ***Key Points to Note from Question***

Given Data:

- Mass of ethylene glycol: 222.6 g
- Mass of water (solvent): 200 g = 0.2 kg
- Density of the solution: 1.072 g/mL

Key Concept:

- Volume of solution = Total mass of solution / Density.

What to Calculate:

- Molality of the solution.
- Volume of the solution.
- Molarity of the solution.

Ans –

$$\text{Mass of solute}(C_2H_6O_2) = 222.6\ g$$

$$\text{Molar Mass }(C_2H_6O_2) = 62\ g/mol$$

$$\text{Number of moles of } C_2H_6O_2 = \frac{222.6}{62} = 3.59\ mol$$

$$\text{Mass of solvent (water)} = 200\text{ g}$$

$$\text{The total mass of solution} = 422.6\text{ g}$$

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

$$\text{Volume} = \frac{\text{mass}}{\text{density}} = \frac{422.6}{1.072} = 394.21\ ml$$

Thus,

$$\text{Molality} = \frac{3.59}{200} = 17.95\ m$$

$$\text{Molarity} = \frac{3.59 \times 1000}{394.21} = 9.106\ M$$

1.9

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 water samples.

📌 ***Key Points to Note from Question***

Given Data:

- Contamination level of chloroform ($CHCl_3$): 15 ppm (by mass).
- Molar mass of chloroform: 119.5 g/mol.

What to Calculate:

- Percent by mass of chloroform.
- Moles of chloroform.
- Molality of chloroform in the water sample.

Ans – (i) We are aware that the value of measurement might indicate the quantity in parts per million (ppm) or parts in 106106 of the mixture.

Meanwhile, 15 ppm in this context refers to 15 parts per 10^6 of solution.
As a result, the mass proportion is provided as,

$$\text{Mass }\% = \frac{15}{10^6} \times 100 = 15 \times 10^{-4}$$

(ii) 15 ppm indicates that there is 15 g of chloroform in 10^6 grams of solution, meaning that the solvent density is equal to 10^6 grams.
Chloroform's mole fraction is 119.5 grams/mol.

$$\text{Number of moles of chloroform} = \frac{15}{119.5} = 0.125\ mol$$

$$\Rightarrow \text{Molality} \ = \frac{0.125 \times 1000}{10^6} = 125\ \times 10^{-6} m$$

1.10

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

📌 ***Key Points to Note from Question***

Key Concept:

- **Molecular interactions in the mixture:**
    - **Hydrogen bonding**: Both alcohol and water can form hydrogen bonds with each other.
    - **Dipole interactions**: Polar nature of both molecules leads to dipole-dipole interactions.

Ans – Both of these substances have a high propensity to create intermolecular hydrogen bonds. Alcohol & H_2O molecules establish H-bonds if these two liquids are mixed, resulting in the creation of a solution. Compared to pure H_2O, these connections are weaker & less widespread. As a result, they exhibit a favourable departure from optimum conduct. Consequently, compared to pure water & pure alcohol, the normal versions of the alcohol and water combination will possess a reduced boiling point & a larger vapor pressure.

1.11

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

📌 ***Key Points to Note from Question***

Key Concept:

- **Gas solubility and temperature:**
    - Solubility of gases decreases with temperature because dissolving gas in liquid is an exothermic process.
    - Higher temperature provides energy for gas molecules to escape from the liquid phase into the vapor phase.

Ans – Exothermic means that surplus thermal energy is released during the dissolution of gasses in water. The state of equilibrium moves backward following Le Chatelier's principle if the operation's temperature rises more. As a result, gasses lose their solubility in liquids.

1.12

State Henry's law and mention some important applications.

Ans – As Henry's law states that:
The partial pressure of a gas is entirely linked to its solubility in a liquid at a certain temperature level. This law governs how pressure affects the solubility of a gas element that is already present in a liquid state.
In mathematics, $P = K_H x$
wherein P is defined as the partial pressure of the gas element.
The molecular proportion of gas present in the mixture is denoted by x.
Henry's Law constant is K_H.
Some common applications of Henry's law include:

a. Plunging into the sea.
b. In the manufacturing of fizzy drinks.

1.13

The partial pressure of ethane over a solution containing 6.56×10^{-3}g of ethane is 1 bar. If the solution contains 5×10^{-2}g of ethane, then what shall be the partial pressure of the gas?

📌 ***Key Points to Note from Question***

Given Data:

- Initial mass of ethane: 6.56×10^{-3} g.
- Partial pressure corresponding to the initial mass: 1 bar.
- New mass of ethane: 5×10^{-2} g.

Key Concept:

- Henry's Law Proportionality

What to Calculate:

- New partial pressure (P_2) corresponding to the new mass of ethane.

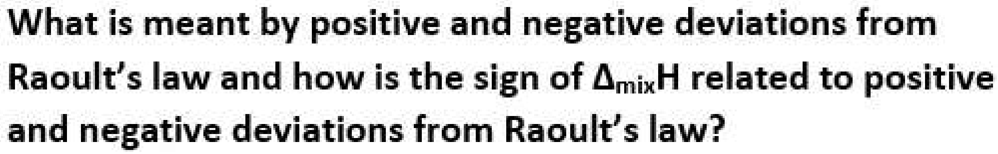
Ans – Using Henry's law, we all can understand that the overall solubility of a gas in a liquid medium can be directly proportional to the overall pressure of the gas. This nominal proportionality sign can be swiftly replaced using Henry's constant. Hence, it can be defined as;
$m = K_H \times P$

Therefore, the mole fraction can be directly proportional to the overall ethane mass.
Hence,
In case 1 we can write,
$6.56 \times 10^{-3} = K_H \times 1$
In case 2 we can write,
$5 \times 10^{-2} = K_H \times x$, where x is the partial pressure of the gas when the solution comprises 5×10^{-2}g of ethane.
Now, equating both the equations together we get,

$$6.56 \times 10^{-3} = \frac{5 \times 10^{-2}}{x}$$

Therefore, partial pressure of the gas will be, x = 7.62 bar

1.14

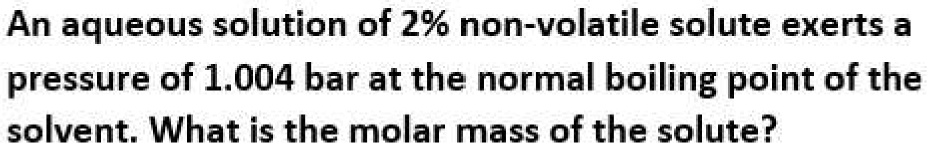


What is meant by positive and negative deviations from Raoult's law and how is the sign of $\Delta_{mix}H$ related to positive and negative deviations from Raoult's law?

Ans – Going by Raoult's law, the positive deviation can be depicted as,
Solutions that showcase positive deviation from Raoult's law if they contain higher vapor pressure well above the prescribed limit as mentioned in the law.
Meanwhile, $\Delta_{mix}H$ remains positive since the energy can be gathered for dismantling the robust connection and building a feeble connection. Similarly, $\Delta_{mix}V$ stands in the positive zone based on the expansion of volume that occurs during this event.
Going by Raoult's law, the negative deviation can be depicted as,

Solutions that portray negative deviation using Raoult’s law if they contain lower vapor pressure well above the prescribed limit as mentioned in the law.
Meanwhile, $\Delta_{mix}H$ stays at the negative zone since the energy is dispersed because of the external replacement of feeble interactions shadowed by the much more robust ones.

1.15



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?

📌 ***Key Points to Note from Question***

Given Data:

- Mass percentage of solute: **2%** (2 g solute in 100 g solution).
- Vapor pressure of pure solvent (P_0) at boiling point: **1.013 bar**.
- Vapor pressure of the solution (P): **1.004 bar**.
- Mass of solvent: **98 g = 0.098 kg**.

Key Concept:

- Relative Lowering of Vapor Pressure Formula
- Rearrange to solve for the molar mass of the solute.

What to Calculate:

- Relative lowering of vapor pressure.
- Molar mass of the solute.

Ans – As we all know,

The vapor pressure of pure water during its boiling point is defined as P^0 = 1.103 bar.
Meanwhile, the vapor pressure of the solution can be calculated as P_s = 1.004 bar.
Based on the above statement, an aqueous solution has a 2% non-volatile solute. Hence, this gives us an idea that the solution remains at 100 gms from which 2 gms (W_2) is regarded as the non-volatile solute & the desired solvent stands at 98 gms (W_1).
Then, the overall molar mass of water will be, (M_1) = 18 g/mol. Considering Raoult’s law for dilute solutions; we get,

$$\frac{P^0 - P_s}{P^0} = \frac{n_2}{n_1 + n_2} \approx \frac{n_2}{n_1}$$

$$\Rightarrow \frac{P^0 - P_s}{P^0} = \frac{\frac{W_2}{M_2}}{\frac{W_1}{M_1}}$$

$$\frac{1.103 - 1.004}{1.103} = \frac{\frac{2}{M_2}}{\frac{98}{18}}$$

$$\frac{0.099}{1.103} = \frac{2}{M_2} \times \frac{18}{98}$$

$$M_2 = \frac{2 \times 18 \times 1.103}{0.099 \times 98} = 41.34\ \mathrm{g/mol}$$

Finally, the molar mass of the solute is defined by (M_2) = 41.34 g/mol.

1.16

Heptane and octane form an ideal solution. At 373 K, the vapor pressure of the two liquid components is 105.2 kPa and 46.8 kPa respectively. What will be the vapor pressure of a mixture of 26.0 g of heptane and 35.0 g of octane?

📌 ***Key Points to Note from Question***

Given Data:

- Vapor pressure of heptane: 105.2 kPa.
- Vapor pressure of octane: 46.8 kPa.
- Mass of heptane: 26.0 g.
- Mass of octane: 35.0 g.

Key Concept:

- Raoult's Law for Vapor Pressure of a Solution

What to Calculate:

- Moles of heptane and octane.
- Mole fractions of heptane and octane.
- Total vapor pressure of the mixture.

Ans – Given,

Vapour Pressure of heptane = 105.2 kPa

Vapour Pressure of octane = 46.8 kPa

Mass of heptane = 26 g

Mass of octane = 35 g

Molar mass of heptane = 100 g/mol

Molar mass of octane = 114 g/mol

$$\Rightarrow \text{No. of moles of heptane present} = \frac{26}{100} = 0.26 \text{ mol}$$

$$\Rightarrow \text{No. of moles of octane present} = \frac{35}{114} = 0.307 \text{ mol}$$

$$\text{Mole fraction of heptane} = \frac{0.26}{0.26+0.307} = 0.458$$

$$\text{Mole fraction of octane} = 1 - 0.458 = 0.541$$

Vapor pressure of heptane = $x_H \times P^0$ = 0.458 × 105.2

= 48.1816 kPa

Vapor pressure of octane = $x_O \times P^0$ = 0.541 × 46.8

= 25.3188 kPa

∴ Vapor pressure of mixture = 48.1816 + 25.3188 = 73.5004kPa

1.17

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

📌 ***Key Points to Note from Question***

Given Data:

- Vapor pressure of pure water (P_0): **12.3 kPa**.
- Solution concentration: **1 molal** (1 mole of solute in 1 kg of water).
- Molar mass of water: **18 g/mol**.

Key Concept:

- Relative Lowering of Vapor Pressure Formula
- Rearrange to find the vapor pressure of the solution (P)

What to Calculate:

- Moles of solvent (water).
- Vapor pressure of the solution (P).

Ans – One molal solution is one mole of solute in 1000g of solvent.

Molar mass of water (solvent) $= 18\ \text{g mol}^{-1}$

$\therefore$ Moles of water $= \frac{1000}{18} = 55.5$ moles

$\therefore$ Mole fraction of salute $= \frac{1}{1+55.5} = 0.0177$

Now, $\frac{p^{o}-p_{s}}{p^{o}} = x_2$

$$\frac{1.23 - P_s}{12.3} = 0.0177$$

$P_s = 12.08$ kPa

1.18

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

📌 ***Key Points to Note from Question***

Given Data:

- Molar mass of non-volatile solute: **40 g/mol**.
- Mass of octane (solvent): **114 g**.
- Vapor pressure of pure octane (P_0): Assume it to be 100% for this calculation.
- Desired vapor pressure of the solution: **80%** of pure octane's vapor pressure.

Key Concept:

- Raoult's Law for Vapor Pressure
- Moles of Solute Formula

What to Calculate:

- Mole fraction of solute.
- Moles of solute required.
- Mass of the non-volatile solute.

Ans – According to Raoult's Law,

$$\frac{P_A^o - P_S}{P_S} = \frac{n_B}{n_A} = \frac{W_B}{M_B} \times \frac{M_A}{W_A}$$

$P_A^o = 1\ atm$

$P_S = 0.8\ atm$

$P_A^o - P_S = 0.2\ atm$

$M_B = 40\ g\ \mathrm{mol}^{-1}$

$W_A = 114\mathrm{g};\ M_A(C_8H_{18}) = 114\ \mathrm{g\ mol}^{-1}$

$$W_B = \frac{(P_A^o - P_S)}{P_S} \times \frac{M_B \times W_A}{M_A}$$

$$= \frac{(0.2\ \mathrm{atm})}{(0.8\ \mathrm{atm})} \times \frac{(40\ \mathrm{g\ mol}^{-1})\times(114\mathrm{g})}{(114\ \mathrm{g\ mol}^{-1})} = 10.0\ \mathrm{g}$$

1.19

A solution containing 30g of non-volatile solute exactly in 90 g of water has a vapor pressure of 2.8 kPa at 298 K. Further, 18g of water is then added to the solution and the new of vapor pressure becomes 2.9 kPa at 298 K. Calculate
(i) the molar mass of the solute
(ii) vapor pressure of water at 298 K

📌 ***Key Points to Note from Question***

Given Data:

- Initial solution:
    - Mass of solute: **30 g**.
    - Mass of water: **90 g = 0.09 kg**.
    - Vapor pressure: **2.8 kPa**.
- After adding water:
    - Additional mass of water: **18 g = 0.018 kg**.
    - New vapor pressure: **2.9 kPa**.

Key Concept:

- Relative Lowering of Vapor Pressure Formula

What to Calculate:

- Initial vapor pressure data to find molar mass of the solute.
- Second vapor pressure data to calculate pure water's vapor pressure (P_0).

Ans – Let molar mass of solute be M g mol^{-1}

Moles of solute present

$$= \frac{30\text{g}}{\text{M g mol}^{-1}} = \frac{30}{\text{M}}\text{mol}$$

Moles of solvent present, $(n_1) = \frac{90}{18} = 5$ moles

$$\therefore \frac{P^o - P_S}{P^o} = \frac{n_2}{n_1 + n_2}$$

$$\frac{P^o - 2.8}{P^o} = \frac{30/M}{5 + 30/M}$$

$$1 - \frac{2.8}{P^o} = \frac{30}{(5M + 30)}$$

$$1 - \frac{30}{(5M + 30)} - \frac{2.8}{P^o}$$

$$1 - \frac{6}{(M + 6)} = \frac{2.8}{P^o}$$

$$\frac{M + 6 - 6}{M + 6} = \frac{2.8}{P^o}$$

$$\frac{M}{M + 6} = \frac{2.8}{P^o}$$

$$\frac{2.8}{P^o} = 1 + \frac{6}{M} \qquad \ldots \quad \text{(i)}$$

After adding 18 g of water,

Moles of water becomes

$$= \frac{90 + 18}{18} = \frac{108}{18} = 6 \text{ moles}$$

$$\therefore \frac{P^o - P_S}{P^o} = \frac{30/M}{6 + 30/M}$$

P_S New Vapour pressure $= 2.9$ kPa

$$\frac{P^o - 2.9}{P^o} = \frac{\frac{30}{M}}{M(6M + 30)} = \frac{5}{M + 5}$$

$$1 - \frac{2.9}{P^o} = \frac{5}{M + 5}$$

$$1 - \frac{5}{M + 5} = \frac{2.9}{P^o}$$

$$\frac{M + 5 - 5}{M + 5} = \frac{2.9}{P^o}$$

$$\frac{2.9}{P^o} = \frac{M + 5}{M} \Rightarrow = 1 + \frac{5}{M}$$

$$\frac{P^o}{2.9} = 1 + \frac{5}{M} \qquad \text{... (ii)}$$

Dividing equation (i) by (ii), we get,

$$\frac{2.9}{2.8} = \frac{1 + 6/\mathrm{M}}{1 + 5/\mathrm{M}}$$

$$2.9\left(1 + \frac{5}{\mathrm{M}}\right) = 2.8\left(1 + \frac{6}{\mathrm{M}}\right)$$

$$2.9 + \frac{14.5}{\mathrm{M}} = 2.8 + \frac{2.8 \times 6}{\mathrm{M}}$$

$$2.9 + \frac{14.5}{\mathrm{M}} = 2.8 + \frac{16.8}{\mathrm{M}}$$

$$0.1 = \frac{16.8}{\mathrm{M}} - \frac{14.5}{\mathrm{M}} = \frac{2.3}{\mathrm{M}}$$

$$\mathrm{M} = \frac{2.3}{0.1}$$

$$\mathrm{M} = 23\ \mathrm{g\ mol^{-1}}$$

Putting $\mathrm{M} = 23$, in equation (i) we get,

$$\frac{\mathrm{P^o}}{2.8} = 1 + \frac{6}{23} = \frac{29}{23}$$

$$\mathrm{P^o} = \frac{29}{23} \times 2.8 = 3.53\ kPa.$$

1.20

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

📌 ***Key Points to Note from Question***

Given Data:

- Freezing point of 5% solution of cane sugar: **271 K**.
- Freezing point of pure water: **273.15 K**.

Key Concept:

- Freezing Point Depression Formula
- For identical percentages, the freezing point depression is proportional to the molality, which depends on the molar mass of the solute.

What to Calculate:

- Freezing point depression for glucose solution (ΔTf).

- Freezing point of the glucose solution.

Ans – In 5% (by mass) solution, mass of sugar denotes 5g in 100g of solvent (water).

$$\text{Molar mass of sugar} = 342 \text{ g mol}^{-1}$$

$$\text{Molality of sugar solution} = \frac{5\times1000}{342\times100} = 0.146$$

$$\therefore \Delta T_f \text{ for sugar solution} = 273.15 - 271$$

$$= 2.15^\circ$$

$$\Delta T_f - K \times m$$

$$\Delta T_f = K_f \times 0.146$$

$$\Rightarrow K_f = \frac{2.15}{0.146}$$

Molality of glucose solution

$$= \frac{5}{18} \times \frac{1000}{100} = 0.278$$

$$\text{Molar mass of glucose} = 180 \text{ g mol}^{-1}$$

$$\Delta T_f = K_f \times m = \frac{2.15}{0.146} \times 0.278 = 4.09^\circ$$

$\therefore$ Freezing point of glucose solution

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

1.21

Two elements A and B form compounds having formula AB_2 and AB_4. When dissolved in 20g 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. Calculate atomic masses of A and B.

📌 ***Key Points to Note from Question***

Given Data:

- **Compound AB_2:**
    - Mass of AB_2: **1 g**.
    - Depression in freezing point (ΔT_f): **2.3 K**.
- **Compound AB_4:**

  - Mass of AB_4: **1 g**.
  - Depression in freezing point (ΔT_f): **1.3 K**.
- Solvent: **Benzene (C_6H_6)**.
- Mass of benzene: **20 g = 0.02 kg**.
- Freezing point depression constant (K_f): **5.1 K·kg/mol**.

Key Concept:

- Freezing Point Depression Formula
- Molar Mass of Solute

What to Calculate:

- Molar masses of AB_2 and AB_4.
- Atomic masses of A (M_A) and B (M_B) by solving the system of equations.

Ans – By using the relation we know,

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

$$\therefore M_{AB_2} = \frac{1000 \times 5.1 \times 1}{20 \times 2.3} = 110.87 \text{ g mol}^{-1}$$

$$M_{AB_4} = \frac{1000 \times 5.1 \times 1}{20 \times 1.3} = 196.15 \text{ g mol}^{-1}$$

The atomic masses of A and B are denoted as 'p' and 'q', respectively.

Then molar mass of

$AB_2 = p + 2q = 110.87 \text{g mol}^{-1}$ ..... (i)

And molar mass of

$AB_4 = p + 4q = 196.15 \text{ g mol}^{-1}$ .... (ii)

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

$2q = 85.28$

$\Rightarrow q = 42.64$

Putting $q = 42.64$ in equation, (i), we get

$p = 110.87 - 85.28$

$p = 25.59$

$\therefore$ atomic mass of A $= 25.59 \text{ g mol}^{-1}$

atomic mass of B $= 42.64 \text{ g mol}^{-1}$

1.22

At 300 K, 36g of glucose present in a litre of its solution has an osmotic pressure of 4.08 bar. If the osmotic pressure of the solution is 1.52 bars at the same temperature, what would be its concentration?

📌 ***Key Points to Note from Question***

Given Data:

- **Initial Data for Glucose Solution:**
    - Mass of glucose: **36 g**.
    - Volume of solution: **1 L**.
    - Osmotic pressure (Π): **4.08 bar**.
    - Temperature (T): **300 K**.
- **New Data:**
    - Osmotic pressure (Π): **1.52 bar**.

Key Concept:

- Osmotic Pressure Formula

What to Calculate:

- Molar concentration of the solution when $\Pi = 1.52$ bar.

Ans –

$$\pi = CRT = \frac{W_B \times R \times T}{M_B \times V}$$

R, T and V are constants for both the solution.

$$\text{1st case: } (4.98 \text{ bar}) = \frac{(36 \text{ g}) \times R \times T}{(180 \text{ g mol}^{-1}) \times V}$$

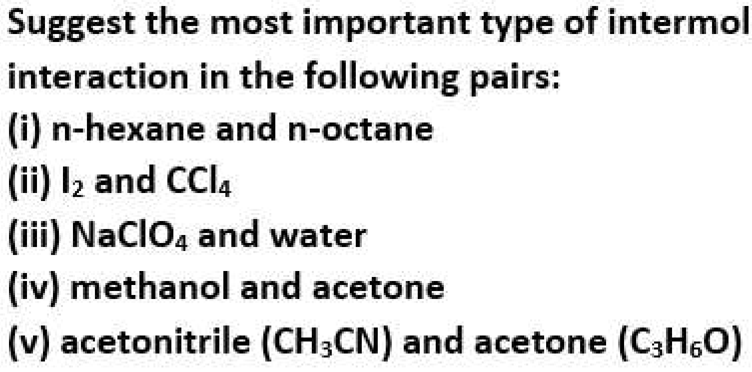
$$\text{2nd case: } (1.52 \text{ bar}) = \frac{W_B \times R \times T}{M_B \times V}$$

Divide eqn. (ii) by eqn. (i),

$$\frac{(1.52 \text{ bar})}{(4.98 \text{ bar})} = \frac{W_B}{M_B} \times (5 \text{ mol})$$

$$\text{Or } \frac{W_B}{M_B} = \frac{1.52}{4.98} \times \frac{1}{(5 \text{ mol})} = 0.0610 \text{ mol}^{-1}$$

1.23



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)

Ans – (i) n-hexane & n-octane – These 2 are nonpolar molecules. So, the intermolecular interaction will happen in London-oriented dispersion forces.

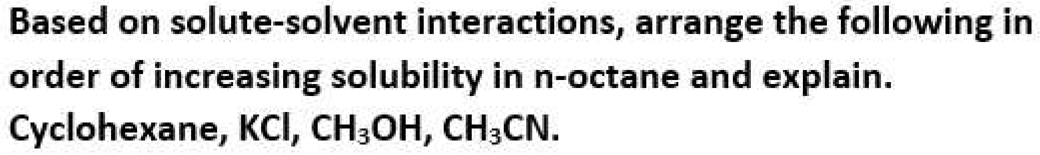
(ii) I_2 & CCl_4 – These 2 molecules are nonpolar & the intermolecular interactions will happen as London dispersion forces.

(iii) $NaClO_4$ & H_2O – The intermolecular interactions are completely based on ion-dipole interactions where $NaClO_4$ is an ionic compound and the H_2O element is a polar molecule.

(iv) Methanol & Acetone – These 2 elements are polar molecules and their intermolecular conversion might occur as dipole-dipole interactions.

(v) Acetonitrile (CH_3CN) & Acetone (C_3H_6O) – The aforementioned elements are polar molecules and the intermolecular interactions in them occur as dipole-dipole-based interactions.

1.24



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

Cyclohexane, KCl, CH_3OH, CH_3CN.

Ans – The segregation of elements based on the ascending order of solubility in n-octane can be defined as,
KCl < CH_3OH < CH_3CN < Cyclohexane.

It happens due to the below factors,

- KCl is an ionic element and won't be dissolved in n-octane.
- CH_3OH is a polar molecule & may get dissolved in n-octane.
- CH_3CN is another polar molecule but has a lesser value when compared to CH_3OH. Hence, it might get dissolved in n-octane to a considerable range.
- Cyclohexane is also a polar & basic molecule that impacts its dissolving features across different proportions in the group.

1.25

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

📌 ***Key Points to Note from Question***

Key Concept:

- **Solubility in Water:**
    - **Highly soluble** compounds: Polar molecules, especially those capable of hydrogen bonding with water.
    - **Partially soluble** compounds: Compounds that have a balance of polar and nonpolar characteristics, leading to moderate solubility in water.
    - **Insoluble** compounds: Nonpolar compounds or compounds that cannot interact well with water molecules.

Ans – (i) Phenol – Remains insoluble in H_2O

(ii) Toluene – It isn't soluble with one another because of its non-polar specification.

(iii) Formic Acid – It remains partially soluble in water.

(iv) Ethylene Glycol – It is partially soluble in water.

(v) Chloroform – It remains highly insoluble in water.

(vi) Pentanol – Pentanol is partially soluble in water because of its presence of the alcohol functional group.

1.26

If the density of some lake water is 1.25 g/ml and contains 92g of Na^+ ions per kg of water, calculate the molality of Na^+ ions in the lake.

📌 ***Key Points to Note from Question***

Given Data:

- Density of lake water: **1.25 g/mL**.
- Mass of Na^+: **92 g per kg of water**.

Key Concept:

- Molality Formula
- Since 92 g of Na^+ is dissolved in 1 kg of water, the solvent mass is already given as **1 kg**.

What to Calculate:

- Molality

Ans –

$$\text{Molality of Na}^+ \text{ ions (m)} = \frac{\text{No.of moles of Na}^+ \text{ ions}}{\text{Mass of water in kg}}$$

$$= \frac{(92g)/(23g\text{mol}^{-1})}{1\text{ kg}}$$

$$= 4\text{ mol kg}^{-1} = 4\text{ m}$$

1.27

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

Ans – Maximum molarity of CuS = Solubility of CuS

Solubility of CuS will be S mol L^{-1}

$$\text{CuS} \rightleftharpoons \text{Cu}^{2+} + \text{S}^{2-}$$

$$\text{K}_{\text{sp}} = 6 \times 10^{-16}$$

$$\text{K}_{\text{sp}} = [\text{Cu}^{2+}][\text{S}]^{2-}$$

$$\Rightarrow 6 \times 10^{-16} = \text{S} \times \text{S} = \text{S}^2$$

$$\therefore \text{S} = \sqrt{6 \times 10^{-16}}$$

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

1.28

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 450g of CH_3CN.

📌 *Key Points to Note from Question*

Given Data:

- Mass of aspirin: **6.5 g**
- Mass of acetonitrile: **450 g**

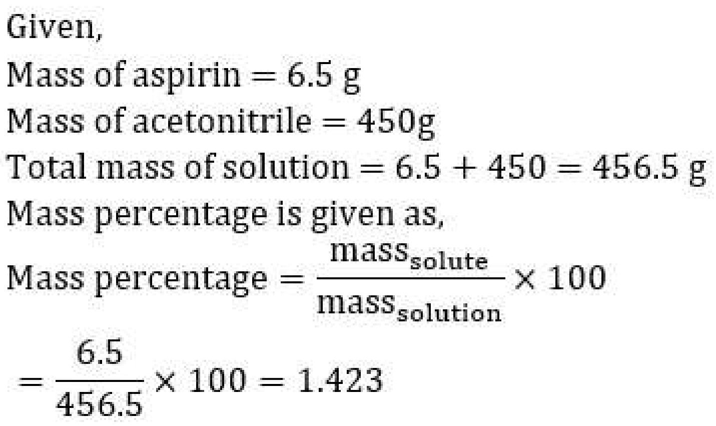
Key Concept:

- Mass Percentage Formula

What to Calculate:

- Total mass of the solution.
- Mass percentage of aspirin in the solution.

Ans –



Given,

Mass of aspirin = 6.5 g

Mass of acetonitrile = 450g

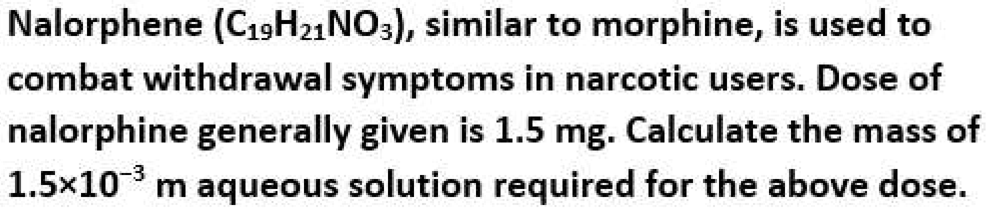
Total mass of solution = 6.5 + 450 = 456.5 g

Mass percentage is given as,

$$\text{Mass percentage} = \frac{\text{mass}_{\text{solute}}}{\text{mass}_{\text{solution}}} \times 100$$

$$= \frac{6.5}{456.5} \times 100 = 1.423$$

1.29



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

📌 *Key Points to Note from Question*

Given Data:

- Molarity of solution(m): **1.5×10^{-3} m** (molality).
- Dose of nalorphene: **1.5 mg = 1.5×10^{-3} g**.

Key Concept:

- Molality Formula
- Moles of Solute Formula

What to Calculate:

- Moles of nalorphene.
- Mass of aqueous solution required to deliver the given dose.

Ans – As we have mentioned,

The molality of solution = 1.5×10^{-3} m shows us that the 1.5×10^{-3} moles are still available in 1 kg of solvent.
Molar mass of nalorphine = 311 g/mol
Hence, The mass of solute = $1.5 \times 10^{-3} \times 311 = 0.4665$g
Total mass of the solution = 0.4665 + 1000 = 1000.4665 g
Based on the above equations, we can conclude that 0.4665 g Narlophene demands a total solution of 1000.4665 gms.
Therefore, 1.5 mg of Narlophene might consist of,

$$\text{Mass of Solution} = \frac{1.5 \times 10^{-3} \times 1000.4665}{0.4665} = 3.216\text{g}$$

1.30

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

📌 ***Key Points to Note from Question***

Given Data:

- Volume of solution: **250 mL = 0.25 L**
- Molarity (M): **0.15 M**

Key Concept:

- Molarity Formula
- Mass of Solute Formula

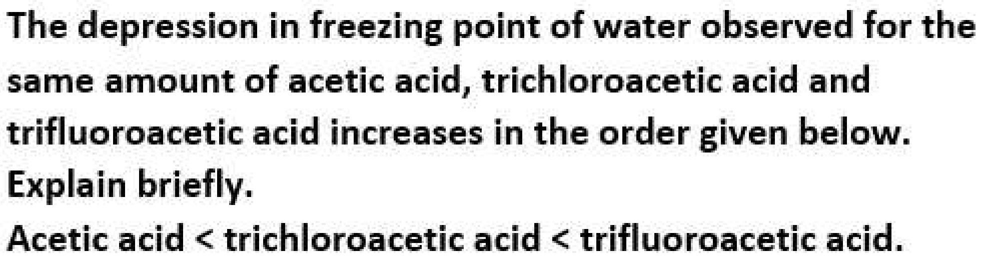
What to Calculate:

- Moles of benzoic acid required.
- Mass of benzoic acid needed to prepare the solution.

Ans – As we all know,

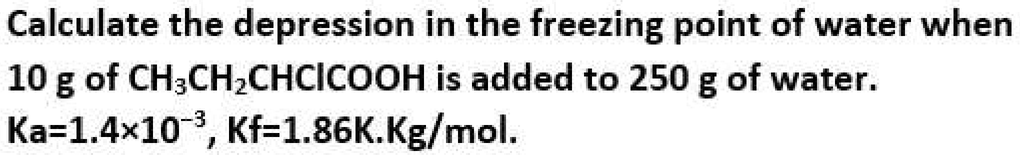
We also contain a 0.15M solution depicting the solution of 1000 ml with 0.15 moles of solute.
Hence, to build 250 ml of solution we might need = (250 × 0.15)/1000 = 0.0375mol
The total molar mass of benzoic acid = 122 g/mol
So, the total mass of benzoic acid can be defined as,
Mass = 0.0375 × 122 = 4.575g

1.31


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 below. Explain briefly.
Acetic acid < trichloroacetic acid < trifluoroacetic acid.

Ans – As we know, fluorine has been regarded as highly electronegative when compared to chlorine, which has the maximum number of electron-withdrawing inductive effects. Hence, trifluoroacetic acid will be considered as the strongest trichloroacetic acid. The second most common one, acetic acid is deemed as the weakest acid of all due its the absence of an electron-withdrawing pair of ions. So, F_3CCOOH ionizes to the maximum possible threshold whereas CH_3COOH ionizes to a minimum considerable threshold in water. The higher the value of ionization, the more massive the depression in the freezing point will get. That is the reason why the depression order in the freezing point can be framed as $CH_3COOH < Cl_3CCOOH < F_3CCOOH$.

1.32


Calculate the depression in the freezing point of water when 10 g of $CH_3CH_2CHClCOOH$ is added to 250 g of water. $Ka=1.4×10^{-3}$, Kf=1.86K.Kg/mol.

📌 ***Key Points to Note from Question***

Given Data:

- Mass of solute ($CH_3CH_2CHClCOOH$): **10 g**
- Mass of solvent (water): **250 g = 0.25 kg**
- Acid dissociation constant (Ka): **$1.4 × 10^{-3}$**
- Cryoscopic constant (Kf): **1.86 K·kg/mol**

Key Concept:

- Freezing Point Depression Formula

What to Calculate:

- Moles of solute.
- Degree of dissociation (α).
- Van't Hoff factor (i).
- Depression in freezing point (ΔT_f)

Ans – As you all know,
The total mass of $CH_3CH_2CHClCOOH$ is 10 g
Molar mass of $CH_3CH_2CHClCOOH$ = 122.5 g/mol

No. of moles of the same will be n = 10/122.5 = 0.0816moles
The total mass of water (solvent) can be defined as W_1 = 250 g

Molality is given by; $\frac{n}{W_1} = \frac{0.0816 \times 1000}{250} = 3.265 \times 10^{-1}$ m

Given,

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

$K_f = 1.86$ K. Kg/mol.

Dissociation reaction is given as,

$$\text{at initial conc. } \underset{C}{CH_3CH_2CHClCOOH} \leftrightarrow \underset{0}{CH_3CH_2CHClCOO^-} + H^+_0$$

$$\text{at equilibrium } \underset{C(1-\alpha)}{CH_3CH_2CHClCOOH} \leftrightarrow \underset{C\alpha}{CH_3CH_2CHClCOO^-} + H^+_{C\alpha}$$

Where, α is the degree of dissociation.

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

As, $\alpha \ll 1, (1-\alpha) \approx 1$

$K_a = C\alpha^2$

$$\alpha = \sqrt{\frac{K_a}{C}} = 0.06548$$

Again,

$$\text{at initial conc. } \underset{C}{CH_3CH_2CHClCOOH} \leftrightarrow \underset{0}{CH_3CH_2CHClCOO^-} + H^+_0$$

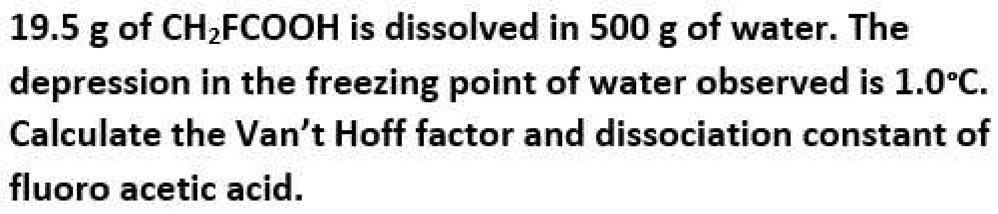
$$\text{at equilibrium } \underset{C(1-\alpha)}{CH_3CH_2CHClCOOH} \leftrightarrow \underset{C\alpha}{CH_3CH_2CHClCOO^-} + H^+_{C\alpha}$$

$$\text{initial mole } \underset{1}{CH_3CH_2CHClCOOH} \leftrightarrow \underset{0}{CH_3CH_2CHClCOO^-} + H^+_0$$

$$\text{at equilibrium } \underset{1-\alpha}{CH_3CH_2CHClCOOH} \leftrightarrow \underset{\alpha}{CH_3CH_2CHClCOO^-} + H^+_{\alpha}$$

Hence,
The total moles that remain at equilibrium = 1 – α + 2α = 1 + α
Meanwhile, the Van’t Hoff factor can be defined as i = 1 + α
Finally, placing the total values we obtain will be i = 1 + 0.06548 = 1.06548
Ultimately, the reduction in freezing point can be mentioned as,

1.33



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 fluoro acetic acid.

📌 *Key Points to Note from Question*

Given Data:

- Mass of solute (CH_2FCOOH): **19.5 g**
- Mass of water (solvent): **500 g = 0.5 kg**
- Depression in freezing point (ΔT_f): **1.0°C**

Key Concept:

- Freezing Point Depression Formula
- Relation Between Van't Hoff Factor and Dissociation
- Degree of Dissociation
- Dissociation Constant

What to Calculate:

- Molality of the solution.
- Van't Hoff factor (i).
- Degree of dissociation (α).
- Dissociation constant (K_a)

Ans – Given the values,

Mass of water, $w_1 = 500$ g = 0.5 kg

Mass of fluoroacetic acid (CH_2FCOOH), $w_2 = 19.5$ g

Molal freezing point depression constant for water, $K_f = 1.86 \text{ K kg mol}^{-1}$

$\Delta T_f = 1°C$

Molar mass of CH_2FCOOH (M_2)

$= 2 \times 12 + 3 \times 1 + 1 \times 19 + 2 \times 16$

$= 24 + 3 + 19 + 32 = 78 \text{ g mol}^{-1}$

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

$i = \frac{\Delta T_f}{K_f m}$ ...(i)

$m = \frac{w_2}{M_2 \times w_1}$

$= \frac{19.5\ \text{g}}{(78\ \text{g mol}^{-1}) \times (0.5\ \text{kg})}$

$= 0.5\ \text{mol kg}^{-1}$ .... (ii)

From eq. (i), we get

$i = \frac{1}{(1.86\ \text{K kg mol}^{-1})(05\ \text{mol kg}^{-1})}$

$= 1.0753$

Calculation of dissociation constant, K_a

Let α be the degree of dissociation of CH_2FCOOH then

$$CH_2FCOOH \rightleftharpoons CH_2FCOO^- + H^+$$

Initial conc.	$C\ \text{mol L}^{-1}$	0	0
At equilibrium	$C(1-\alpha)$	$C\alpha$	$C\alpha$

Total $= C(1+\alpha)$

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

$\Rightarrow i = 1 + \alpha$

$\Rightarrow \alpha = i - 1$

$= 1.0753 - 1 = 0.0753$

Now, $K_a = \frac{[CH_2FCOO^-][H^+]}{[CH_2FCOOH]}$

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

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

Considering 500 mL as the solution's volume, we have:

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

$$= 0.5\ \mathrm{M}$$

$$\therefore K_a = \frac{C\alpha^2}{(1-\alpha)}$$

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

$$= \frac{0.5 \times 0.00567}{0.9247} = 0.00307$$

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

1.34

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

📌 ***Key Points to Note from Question***

Given Data:

- Vapor pressure of pure water: **17.535 mm Hg**
- Mass of glucose ($C_6H_{12}O_6$): **25 g**
- Mass of water (solvent): **450 g**

Key Concept:

- Relative Lowering of Vapor Pressure Formula

What to Calculate:

- Moles of glucose (n_{solute}).
- Moles of water ($n_{solvent}$).
- Vapor pressure of water in the solution (P)

Ans – Apply Raoult's Law

$$\frac{P_A^o - P_S}{P_S} = \frac{n_B}{n_A}$$

$$\Rightarrow \frac{P_A^o}{P_S} - 1 = \frac{n_B}{n_A}$$

$$\Rightarrow \frac{P_A^o}{P_S} = 1 + \frac{n_B}{n_A} = 1 + \frac{W_B}{M_B} \times \frac{M_A}{W_A}$$

$W_B = 25\text{g}\,; W_A = 450\text{ g}$

$M_B = 180\text{ g mol}^{-1};\ M_A = 18\text{ g mol}^{-1}$

$P_A^o = 17.535\text{ mm}$

$$\frac{P_A^o}{P_S} = 1 + \frac{(25\text{g}) \times (18\text{gmol}^{-1})}{(180\text{gmol}^{-1}) \times (450\text{g})}$$

$$= 1 + 0.0055 = 1.0055$$

$$P_S\text{ (V.P of water in solution)} = \frac{(173535\text{ mm})}{(1.0055)} = 17.44\text{ mm}$$

1.35

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.

📌 Key Points to Note from Question

Given Data:

- Henry's law constant (k_H): **4.27 × 10^5 mm Hg**
- Pressure of methane: **760 mm Hg**
- Temperature: **298 K**

Key Concept:

- Henry's Law Formula

What to Calculate:

- Solubility of methane in benzene.

Ans – Given values are Henry's law constant, $K_H = 4.27 \times 10^5$ mm Hg and Pressure = 760 mm Hg. Using Henry's law we can calculate the molality as,

By henry's law,
$P=K_Hm$

$$m = \frac{P}{K_H} = \frac{760}{4.27 \times 10^5}$$

Molality / solubility = 1.779×10^{-3} m

1.36

100 g of liquid A (molar mass 140 g/mol) was dissolved in 1000 g of liquid B (molar mass 180 g/mol). The vapor pressure of pure liquid B was found to be 500 torr. Calculate the vapor pressure of pure liquid A and its vapor pressure in the solution if the total vapor pressure of the solution is 475 Torr.

📌 ***Key Points to Note from Question***

Given Data:

- **Liquid A (solute):**
    - Mass of A: **100 g**
    - Molar mass of A: **140 g/mol**
- **Liquid B (solvent):**
    - Mass of B: **1000 g**
    - Molar mass of B: **180 g/mol**
    - Vapor pressure of pure B: **500 Torr**
- **Solution:**
    - Total vapor pressure (P_{total}): **475 Torr**

Key Concept:

- Raoult's Law for Total Vapor Pressure
- Mole Fraction Calculation

What to Calculate:

- Moles of A and B.
- Mole fractions of A and B.
- Vapor pressure of pure liquid A.
- Vapor pressures of A and B in the solution.

Ans – For liquid A number of moles, $n_A = w_1/M_1$ = 100/140 mol = 0.714 mol

For liquid A number of moles, $n_B = w_2/M_2$ = 1000/180 mol = 5.556 mol

Mole fraction of A will be,

$$x_A = \frac{n_A}{n_A + n_B} = \frac{0.714}{0.714 + 5.556} = 0.114$$

Mole fraction of B will be, $x_B = 1 - x_A = 1 - 0.114 = 0.886$

For pure liquid B vapour pressure, P^0_B = 500 torr

⇒ In the solution vapour pressure of liquid B will be,

$P_B = P^0_B\, x_B = 500 \times 0.886 = 443$ torr

Given P_{total} = 475 torr

⇒ $P_A = P_{total} - P_B$

= 475 – 443 = 32 torr

In the solution vapour pressure of liquid A will be,

$P_A = P^0_A\, x_A$

⇒ Vapour pressure of pure liquid A, $P^0_A = P_A/x_A$

= 32/0.114 = 280.7 torr

1.37

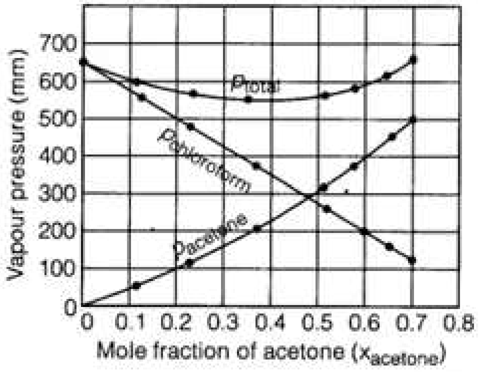
Vapor pressure of pure acetone and chloroform at 328 K are 741.8 mm Hg and 632.8 mm Hg respectively. Assuming that they form ideal solution over the entire range of composition, plot P_{Total}, $P_{Chloroform}$ and $P_{Acetone}$ as a function of $x_{Acetone}$. The experimental data observed for different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

$x_{Acetone}$×100	0	11.8	23.4	36.0	50.8	58.2	64.5	72.1
$P_{Acetone}$/mmHg	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

Ans – The following data can be derived from the provided table;

$x_{Acetone} \times 100$	0	11.8	23.4	36.0	50.8	58.2	64.5	72.1
$P_{Acetone}$/mmHg	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}/mmHg	632.8	603.0	579.5	562.1	580.4	599.5	615.3	641.8

The graph can be represented as follows;



The table and graph indicate that the solution exhibits a negative deviation from ideal behaviour, as evidenced by the downward trend of P_{total}.

1.38

Benzene and toluene form an ideal solution over the entire range of composition. The vapor 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 vapor phase if 80 g of benzene is mixed with 100 g of toluene.

📌 ***Key Points to Note from Question***

Given Data:

- **Benzene:**
  - Mass: **80 g**
  - Molar mass: **78 g/mol**
  - Vapor pressure: **50.71 mm Hg**
- **Toluene:**

    - Mass: **100 g**
    - Molar mass: **92 g/mol**
    - Vapor pressure: **32.06 mm Hg**

Key Concept:

- Raoult's Law for Partial Pressures
- Total Vapor Pressure
- Mole Fraction in Vapor Phase

What to Calculate:

- Moles of benzene and toluene.
- Mole fractions of benzene and toluene in the liquid phase.
- Partial pressures.
- Mole fraction of benzene in the vapor phase.

Ans – From the given data we get,

Benzene mass = 80 g and Molar mass = 78 g/mol

⇒ Number of moles of benzene, n_B = 80/78

= 1.025 moles

Similarly toluene mass = 100 g and Molar mass = 92 g/mol

⇒ Number of moles of benzene, n_T = 100/92

= 1.086 moles

Mole fraction of benzene,

$$x_B = \frac{n_B}{n_B + n_T} = \frac{1.025}{1.025 + 1.086} = 0.4855$$

Mole fraction of toluene, $x_T = 1 – x_B = 1 – 0.4855 = 0.5144$

Given vapor pressures $P^0{}_B$ = 50.71 mm Hg and $P^0{}_T$ = 32.06 mm Hg

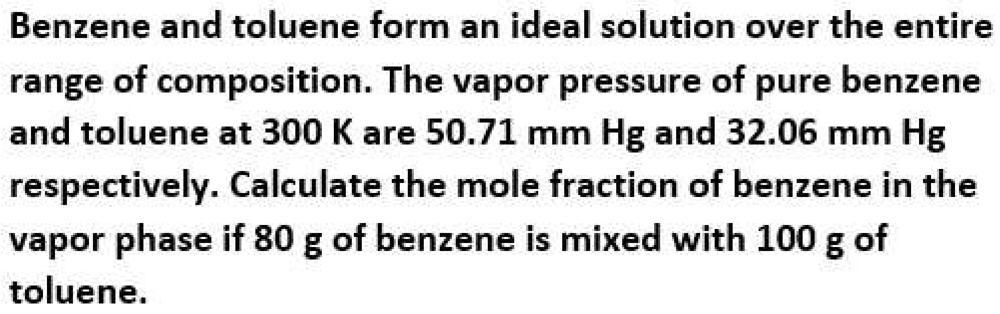
Using raoult's law,

$P_B = P^0{}_B\, x_B = 50.71 \times 0.4855 = 24.61$ mm Hg

$P_T = P^0{}_T\, x_T = 32.06 \times 0.5144 = 16.491$ mm Hg

$$\text{Mole fraction} = \frac{P_B}{P_B + P_T} = \frac{24.61}{24.61 + 16.491} = 0.598$$

1.39



Benzene and toluene form an ideal solution over the entire range of composition. The vapor 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 vapor phase if 80 g of benzene is mixed with 100 g of toluene.

📌 *Key Points to Note from Question*

Given Data:

- **Composition of Air:**
  - Oxygen (O_2): **20% by volume**
  - Nitrogen (N_2): **79% by volume**
- **Henry's Law Constants:**
  - For O_2: $k_H(O_2)=3.3\times10^7$ mm Hg
  - For N_2: $k_H(N_2)=6.51\times10^7$ mm Hg
- **Pressure of Air:**
  - Total pressure (P_{Total}): **10 atm = 7600 mm Hg**

Key Concept:

- Partial Pressure of Each Gas in Air
- Henry's Law for Solubility
- Composition in Water

What to Calculate:

- Partial pressures of O_2 and N_2.
- Solubility of O_2 and N_2 in water.
- Mole fractions of O_2 and N_2 in water

Ans – % of oxygen (O_2) in air = 20%

% of nitrogen (N_2) in air = 79%

Also, it is given that water is in equilibrium with air at a total pressure of 10 atm = (10 × 760) mm = 7600 mm

∴ partial pressure of oxygen,

PO_2 = 20/100 x 7600 mm

= 1520 mm Hg

Partial pressure of nitrogen,

pN_2 = 79/100 x 7600 mm

= 6004 mm Hg

According to Henry's law,
$p = K_{H}.\chi$

For oxygen:

$$P_{O_2} = K_H . \chi_{O_2}$$

$$\Rightarrow \chi_{O_2} = \frac{P_{O_2}}{K_H}$$

$$= \frac{1520 \text{ mm}}{3.30 \times 10^7 \text{ mm}}$$

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

(Given $K_H = 3.30 \times 10^7$ mm)

For nitrogen:

$$P_{N_2} = K_H . \chi_{N_2}$$

$$\Rightarrow \chi_{N_2} = \frac{P_{N_2}}{K_H}$$

1.40

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.

📌 *Key Points to Note from Question*

Given Data:

- Van't Hoff factor (i): **2.47**
- Osmotic pressure (Π): **0.75 atm**
- Volume of solution (V): **2.5 L**
- Temperature (T): **27°C = 300 K**
- Molar mass of $CaCl_2$ = 111 g/mol

Key Concept:

- Osmotic Pressure Formula
- Mass of Solute Formula
- Number of Moles Formula

What to Calculate:

- Molar concentration.
- Moles of $CaCl_2$.

- Mass of $CaCl_2$ required.

Ans – Given values are
Volume, V= 2.5 L
Van't Hoff factor, i = 2.47
Osmotic pressure, π = 0.75 atm
Gas constant, R = 0.0821 L atm K^{-1} mol^{-1}
Temperature, T = 273 + 27 = 300 K
Molar mass of $CaCl_2$, M = 111g/mol
The formula for osmotic pressure is

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

$$W = \frac{\pi MV}{iRT} = \frac{0.75 \times 111 \times 2.5}{2.47 \times 0.0821 \times 300} = 3.42\ g$$

1.41

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

📌 ***Key Points to Note from Question***

Given Data:

- Mass of K_2SO_4: **25 mg = 0.025 g**
- Volume of solution (V): **2 L**
- Temperature (T): **25°C = 298 K**

Key Concept:

- Osmotic Pressure Formula
- Molar Concentration Formula

What to Calculate:

- Moles of K_2SO_4.
- Molar concentration.
- Osmotic pressure (Π) of the solution

Ans –

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

Number if ions produced = Van’t Hoff factor = i = 3

∴ Osmotic pressure is given as,

$$\pi = iCRT$$

$$\pi = i\frac{n}{V}RT = \frac{W}{M \times V}RT$$

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

$$= 5.272 \times 10^{-3}\ \text{atm}$$