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

Answer :

We know that solute and solvent forms solution.

So mass percentage of benzene (solute) :-

$$= \frac{22}{22+122} \times 100 = \frac{22}{144} \times 100 = 15.28\%$$

Similarly mass percentage of CCl_4 :-

$$= \frac{122}{22+122} \times 100 = \frac{122}{144} \times 100 = 84.72\%$$

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

Answer :

For calculating mole fraction, we need moles of both the compounds.

It is given that benzene is 30% in the solution by mass.

So if we consider 100g of solution then 30g is benzene and 70g is CCl_4.

$$Moles of CCl_4 = \frac{Given\ mass}{Molar\ mass} = \frac{70}{154} = 0.4545\ mol$$

$$(Molar\ mass\ of\ CCl_4 = 12 + 4(35.5))$$

Similarly moles of benzene :

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$= \frac{30}{78} = 0.3846 \; (Molar\ mass\ of\ benzene = 6(12) + 6(1))$

So **mole fraction of benzene** is given :

$= \frac{0.3846}{0.3846 + 0.4545}$

$= 0.458$

Question 2.3(a) Calculate the molarity of each of the following solutions:

$30\ g$ of $Co(NO_3)_2.\ 6H_2O$ in $4.3\ L$ of solution

Answer :

For finding molarity we need the moles of solute and volume of solution.

So moles of solute :

$= \frac{Given\ mass}{Molar\ mass} = \frac{30}{291} = 0.103$

$Since\ molar\ mass\ of\ Co(NO_3)_2.6H_2O = 59 + 2(14 + 316) + 6(16 + 12)$

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

Now, $Molarity = \frac{No.\ of\ moles\ of\ solute}{Volume\ of\ solution\ (l)}$

$= \frac{0.103}{4.3} = 0.023\ M$

Question 2.3(b) Calculate the molarity of each of the following solutions:

$30\ mL$ of $0.5\ M\ H_2SO_4$ diluted to $500\ mL$.

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Answer :

By conservation of moles we can write :

$M_1 V_1 = M_2 V_2$

Given that M_1 = 0.5 M and V_1 = 30 ml ; V_2 = 500 ml

$$M_2 = \frac{M_1 V_1}{V_2} = 0.03\ M$$

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

Answer :

Let us assume that the mass of urea required be x g.

So moles of urea will be :

$$Moles = \frac{Given\ mass}{Molar\ mass} = \frac{x}{60}\ moles$$

$$Molality = \frac{Moles}{Mass\ of\ solvent\ in\ Kg} = \frac{\frac{x}{60}}{2.5 - 0.001x} = 0.25$$

we get x = 37

Thus mass of urea required = 37 g.

Question 2.5 Calculate

(a) molality

(b) molarity and

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(c) mole fraction

of KI if the density of 20% (mass/mass) aqueous KI is $1.202\ g\ mL^{-1}$.

Answer :

If we assume our solution is 100 g. Then according to question, 20 g KI is present and 80 g is water.

So moles of KI :

$$= \frac{20}{166} \left(Molar\ mass = 39 + 127 = 166\ g\ mol^{-1} \right)$$

(a) Molality :-

$$Molality = \frac{Moles}{Mass\ of\ solvent\ in\ Kg} = \frac{\frac{20}{166}}{0.08} = 1.506\ m.$$

(b) Molarity :-

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

$$Volume = \frac{Mass}{Density} = \frac{100}{1.202} = 83.19mL$$

$$Molarity = \frac{Moles}{Volume(l)} = \frac{\frac{20}{166}}{83.19 \times 10^{-3}} = 1.45\ M$$

(c) Mol fraction :- Moles of water :-

$$= \frac{80}{18} = 4.44$$

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So, mol fraction of KI :-

$$= \frac{0.12}{0.12 + 4.44} = 0.0263$$

Question 2.6 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.

Answer :

For finding Henry's constant we need to know about the mole fraction of H_2S.

Solubility of H_2S in water is given to be 0.195 m .

i.e., 0.195 moles in 1 Kg of water.

$$Moles\ of\ water := \frac{1000}{18} = 55.55\ moles$$

So $x_{H_2S} = Mole\ fraction\ of\ H_2S$

$$= \frac{0.195}{0.195 + 55.55} = 0.0035$$

At STP conditions, pressure = 1 atm or 0.987 bar

Equation is : $p_{H_2S} = K_h \times x_{H_2S}$

So we get :

$$K_h = \frac{0.987}{0.0035} = 282\ bar$$

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Question 2.7 Henry's law constant for CO_2 in water is $1.67 \times 10^8\ Pa$ 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 $298\ K$.

Answer :

We know that ,

$$p = k_h \times x$$

Pressure of CO_2 = 2.5 atm

We know that : $1\ atm = 1.01 \times 10^5\ Pa$

So, Pressure of CO_2 = 2.53×10^5 Pa

By Henry Law we get,

$$x = \frac{p}{k_h} = \frac{2.53 \times 10^5}{1.67 \times 10^8} = 1.52 \times 10^{-3}$$

Taking density of soda water = 1 g/ml

We get mass of water = 500 g.

So, Moles of water :

$$= \frac{500}{18} = 27.78$$

Also, $$x_{H_2O} = \frac{n_{CO_2}}{n_{H_2O} + n_{CO_2}} \approx \frac{n_{CO_2}}{n_{H_2O}}$$

So, moles of CO_2 = 0.042 mol

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Using relation of mole and given mass, we get

Mass of CO_2 = 1.848 g.

Question 2.8 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.

Answer :

Let the composition of liquid A (mole fraction) be x_A .

So mole fraction of B will be $x_B = 1 - x_A$.

Given that, $P_A^\circ = 450\ mm\ of\ Hg\ ;\ P_B^\circ = 700\ mm\ of\ Hg$

Using Raoult's law ,

$$p_{total} = p_A^\circ\ x_A\ +\ p_B^\circ\ (1 - x_A)$$

Putting values of p_{total} and vapour pressure of pure liquids in the above equation, we get :

$600 = 450.x_A + 700.(1 - x_A)$

or $600 - 700 = 450x_A - 700x_A$

or $x_A = 0.4$

and $x_B = 0.6$

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Now pressure in vapour phase :

$$P_A = p_A^\circ \; x_A$$

= 450(0.4) = 180 mm of Hg

$$P_B = p_B^\circ \; x_B$$

= 700(0.6) = 420 mm of Hg

$$Mole\ fraction\ of\ liquid\ A = \frac{P_A}{P_A\ + P_B}$$

$$= \frac{180}{180\ + 420} = 0.30$$

And mole fraction of liquid B = 0.70

Question 2.9 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.

Answer :

Given that vapour pressure of pure water, $p_w^\circ = 23.8\ mm\ of\ Hg$

Moles of water :

$$= \frac{850}{18} = 47.22$$

Moles of urea :

$$= \frac{50}{60} = 0.83$$

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Let the vapour pressure of water be p_w.

By Raoult's law, we get :

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

or $$\frac{23.8 - p_w}{23.8} = \frac{0.83}{47.22 + 0.83} = 0.0173$$

or p_w = 23.4 mm of Hg.

Relative lowering :- Hence, the vapour pressure(v.P) of water in the solution = 23.4 mm of Hg

and its relative lowering = 0.0173.

Question 2.10 Boiling point of water at $750\ mm\ Hg$ is $99.63^\circ C$. How much sucrose is to be added to $500\ g$ of water such that it boils at $100^\circ C$.

Answer :

Here we will use the formula :

$$\Delta T_b = \frac{K_b \times 1000 \times w_2}{M_2 \times w_1}$$

Elevation in temperature = 100 - 99.63 = 0.37

K_b = 0.52 ; $Molar\ mass\ of\ sucrose = 11(12) + 22(1) + 11(16) = 342\ g\ mol^{-1}$

Putting all values in above formula, we get :

$$w_2 = \frac{0.37 \times 342 \times 500}{0.52 \times 1000} = 121.67\ g$$

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Thus 121.67 g of sucrose needs to be added.

Question 2.11 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 $1.5°C$. $K_f = 3.9\ K\ kg\ mol^{-1}$

Answer :

Elevation in melting point = 1.5 degree celsius.

Here we will use the following equation :

$$\Delta T_b = \frac{K_b \times 1000 \times w_2}{M_2 \times w_1}$$

Putting given values in the above equation :

$$w_2 = \frac{1.5 \times 176 \times 75}{3.9 \times 1000} = 5.08\ g$$

Thus 5.08 ascorbic acid is needed for required condition.

Question 2.12 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 $37°C$.

Answer :

We know that :

$$Osmotic\ Pressure = \Pi = \frac{n}{v}RT$$

We are given with :-

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$$Moles\ of\ polymer = \frac{1}{185000}$$

Volume, V = 0.45 L

Thus osmotic pressure :

$$= \frac{\frac{1}{185000} \times 8.314 \times 10^3 \times 310}{0.45} = 30.98\ Pa$$

NCERT solutions for class 12 chemistry chapter 2 Solutions : Exercises

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

Solution :- A solution is a homogeneous mixture of two or more non-reacting substances. It has two components :- solute and solvent.

Types of solutions are given below :-

Type of Solution	Solute	Solvent	Common Examples
Gaseous Solutions	Gas	Gas	Mixture of oxygen and nitrogen gases
	Liquid	Gas	Chloroform mixed with nitrogen gas
	Solid	Gas	Camphor in nitrogen gas
Liquid Solutions	Gas	Liquid	Oxygen dissolved in water
	Liquid	Liquid	Ethanol dissolved in water
	Solid	Liquid	Glucose dissolved in water
Solid Solutions	Gas	Solid	Solution of hydrogen in palladium
	Liquid	Solid	Amalgam of mercury with sodium
	Solid	Solid	Copper dissolved in gold

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

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Answer:

Solution of hydrogen in palladium is such an example in which solute is a gas and solvent is solid.

Q.2.3(i) Define the following terms:

Mole fraction

Answer :

Mole fraction is defined as the ratio of number of moles of a component and total number of moles in all components.

i.e., $$Mole\ fraction = \frac{Number\ of\ moles\ in\ a\ component}{Total\ number\ of\ moles\ in\ all\ components}$$

Q.2.3(ii) Define the following terms:

Molality

Answer :

It is defined as the number of moles of solute dissolved per kg (1000g) of solvent

i.e., $$Molality = \frac{Number\ of\ moles\ of\ solute}{Mass\ of\ solvent\ in\ Kg}$$

It is independent of temperature.

Q2.3(iii) Define the following terms:

Molarity

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Answer :

Molarity is defined as number of moles of solute dissolved per litre(or 1000ml) of solution.

i.e., $$Molarity = \frac{No.\ of\ moles\ of\ solute}{Volume\ of\ solution\ in\ litre}$$

It depends on temperature because volume is dependent on temperature.

Q2.3(iv) Define the following terms:

Mass percentage.

Answer :

Mass percentage is defined as the percentage ratio of mass of one component to the total mass of all the components.

i.e., $$Mass\ percentage = \frac{Mass\ of\ a\ component}{Total\ mass\ of\ solution} \times 100$$

Q2.4 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 :

According to given question, in 100 g of solution 68 g is nitric acid and rest is water.

So moles of 68 g HNO_3 :-

$$= \frac{68}{63} = 1.08$$

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Density of solution is given to be 1.504.

So volume of 100 g solution becomes :-

$$= \frac{100}{1.504} = 66.49\ mL$$

Thus, molarity of nitric acid is :

$$Molarity = \frac{1.08}{\frac{66.49}{1000}} = 16.24\ M$$

Q2.5 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 :

According to question, 10% mass percentage means in 100 g of solution 10 g glucose is dissolved in 90 g water.

Molar mass of glucose ($C_6H_{12}O_6$) = $180\ g\ mol^{-1}$

So moles of glucose are :

$$\frac{10}{180} = 0.056mol$$

$$Moles\ of\ water = \frac{90}{18} = 5mol$$

$$Molality = \frac{0.056}{0.09} = 0.62\ m$$

Mole fraction :-

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

Molarity :- Volume of 100 g solution :

$$= \frac{100}{1.2} = 83.3\ mL$$

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

Q2.6 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 :

Total amount of mixture of Na_2CO_3 and $NaHCO_3$ = 1 g.

Let the amount of Na_2CO_3 be x g.

So the amount of $NaHCO_3$ will be equal to (1 - x) g.

$$Molar\ mass\ of\ Na_2CO_3 = 106\ ;\ molar\ mass\ of\ NaHCO_3 = 84$$

Now it is given that it is an equimolar mixture.

So, Moles of Na_2CO_3 = Moles of $NaHCO_3$.

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

or x = 0.558 g

$$\text{So } Moles\ of\ Na_2CO_3 = \frac{0.558}{106} = 0.00526$$

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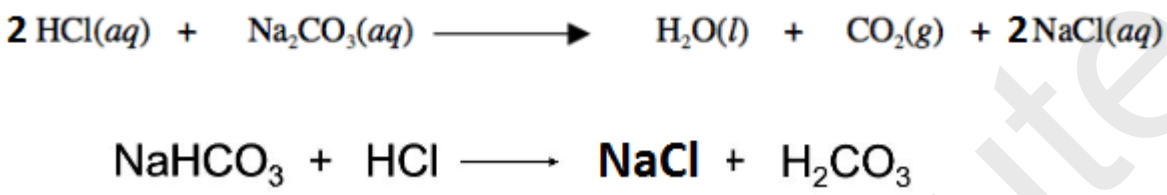
and $Moles\ of\ NaHCO_3 = \frac{1 - 0.558}{84} = 0.0053$



It is clear that for 1 mol of Na_2CO_3 2 mol of HCl is required, similarly for 1 mol of $NaHCO_3$ 1 mol of HCl is required.

So number of moles required of HCl = 2(0.00526) + 0.0053 = 0.01578 mol

It is given that molarity of HCl is 0.1 which means 0.1 mol of HCl in 1l of solution.

Thus required volume :

$= \frac{0.01578}{0.1} = 0.1578\ l = 157.8\ mL$

Q2.7 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 :

According to question we have 2 solute,

Solute 1. : 25% of 300 g gives :

$\frac{25}{100} \times 300 = 75\ g$

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Solute 2. : 40% of 400 g gives :

$$\frac{40}{100} \times 400 = 160\ g$$

So total amount of solute = 75 + 160 = 235 g.

Thus mass percentage of solute is :

$$= \frac{235}{700} \times 100 = 33.5\%$$

and mass percentage of water $= 100 - 33.5 = 66.5\%$

Q2.8 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 :

For finding molality we need to find the moles of ethylene glycol.

Moles of ethylene glycol :

$$= \frac{222.6}{62} = 3.59\ mol$$

We know that :

$$Molality = \frac{Moles\ of\ ethylene\ glycol}{Mass\ of\ water} \times 100$$

$$= \frac{3.59}{200} \times 100 = 17.95\ m$$

Now for **molarity :-**

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Total mass of solution = 200 + 222.6 = 422.6 g

Volume of solution

$$= \frac{422.6}{1.072} = 394.22\ mL$$

So molarity :-

$$= \frac{3.59}{394.22} \times 1000 = 9.11\ M$$

Q2.9(i) 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):

express this in percent by mass

Answer :

We know that 15 ppm means 15 parts per million.

Required percent by mass :

$$= \frac{Mass\ of\ chlorofoam}{Total\ mass} \times 100$$

$$= \frac{15}{10^6} \times 100 = 1.5 \times 10^{-3}\%$$

Q2.9(ii) 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):

determine the molality of chloroform in the water sample.

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Answer :

Moles of chloroform :

$$= \frac{15}{119.5} = 0.1255\ mol$$

Mass of water is 10^6 . (Since contamination is 15 ppm)

So molality will be :

$$= \frac{0.1255}{10^6} \times 1000 = 1.255 \times 10^{-4}\ m$$

Q2.10 What role does the molecular interaction play in a solution of alcohol and water?

Answer :

Both alcohol and water individually have strong hydrogen bonds as their force of attraction. When we mix alcohol with water they form solution due to the formation of hydrogen bonds but they are weaker as compared to hydrogen bonds of pure water or pure alcohol.

Thus this solution shows a positive deviation from the ideal behaviour.

Q2.11 Why do gases always tend to be less soluble in liquids as the temperature is raised?

Answer:

It is known that dissolution of gas in a liquid is an exothermic process. So, by Le Chatelier principle we know that equilibrium shifts backwards as we increase

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temperature in case of exothermic process. Thus gases always tend to be less soluble in liquids as the temperature is raised.

Q2.12 State Henry's law and mention some important applications.

Answer :

According to Henry's law at a constant temperature, the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas present above the surface of solution or liquid.

i.e., $p = k_h x$, Here k_h is Henry's law constant.

Some of its applications are as follows:-

(a) We can increase the solubility of CO_2 in soft drinks, the bottle is sealed under high pressure.

(b) To avoid bends (due to blockage of capillaries) and the toxic effects of high concentrations of nitrogen in the blood, the tanks used by scuba divers are filled with air diluted with helium (11.7% helium, 56.2% nitrogen and 32.1% oxygen).

(c) The partial pressure of oxygen is less at high altitudes than that at the ground level. This leads to low concentrations of oxygen in the blood and tissues climbers. Due to low blood oxygen, climbers become weak and unable to think clearly which are symptoms of a condition known as anoxia.

Q2.13 The partial pressure of ethane over a solution containing $6.56 \times 10^{-3}\ g$ of ethane is 1 bar. If the solution contains $5.00 \times 10^{-2} g$ of ethane, then what shall be the partial pressure of the gas?

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Answer :

Using Henry's Law we can write,

$m = k.P$

Putting value in this equation, we get :

$6.56 \times 10^{-3} = k \times 1$

So, the magnitude of k is 6.56×10^{-3}.

Now, we will again use the above equation for $m = 5.0 \times 10^{-2}\ g$.

So the required partial pressure is :-

$$p = \frac{m}{k} = \frac{5.0 \times 10^{-2}}{6.56 \times 10^{-3}}$$

or $p = 7.62\ bar$

Q2.14 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?

Answer :

Positive and negative deviation: **-** A non-ideal solution is defined as a solution which does not obey Raoult's law over the entire range of concentration i.e., $\Delta_{Mix}H \neq 0$ and $\Delta_{Mix}V \neq 0$. The vapour pressure of these solutions is either higher or lower than that expected by Raoult's law. If vapour pressure is higher, the solution shows a positive deviation and if it is lower, it shows a negative deviation from Raoult's law.

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Enthalpy relation to positive and negative deviation can be understood from the following example:-

Consider a solution made up of two components - A and B. In the pure state the intermolecular force of attraction between them are A-A and B-B. But when we mix the two, we get a binary solution with molecular interaction A-B.

If A-B interaction is weak than A-A and B-B then enthalpy of reaction will be positive thus reaction will tend to move in a backward direction. Hence molecules in binary solution will have a higher tendency to escape. Thus vapour pressure increases and shows positive deviation from the ideal behaviour.

Similarly, for negative deviation, A-B interaction is stronger than that of A-A and B-B.

Q2.15 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 :

It is given that 2% of aq. solution. This means 2 g of non-volatile solute in 98 g of H_2O.

Also the vapour of water at normal boiling point = 1.013 bar.

Using Raoult's law :

$$\frac{p_w^\circ - p}{p_w^\circ} = \frac{w_2.M_1}{w_1.M_2}$$

So we get :

$$M_2 = \frac{2 \times 18 \times 1.013}{0.009 \times 98} = 41.35\ g\ mol^{-1}$$

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Thus the molar mass of non-volatile solute is 41.35 unit.

Q2.16 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 :

Vapour pressure of heptane = $p_h^\circ = 105.2\ KPa$

and vapour pressure of octane = $p_o^\circ = 46.8\ KPa$

Firstly we will find moles of heptane and octane so that we can find vapour pressure of each.

Molar mass of heptane = 7(12) + 16(1) = 100 unit.

and molar mass of octane = 8(12) + 18(1) = 114 unit.

So moles of heptane :

$$\frac{26}{100} = 0.26$$

and moles of octane :

$$\frac{35}{114} = 0.31$$

Mole fraction of heptane = 0.456 and mole fraction of octane = 0.544

Now we will find the partial vapour pressure:-

(i) of heptane :- $p_h = 0.456 \times 105.2 = 47.97\ KPa$

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(ii) of octane :- $p_o = 0.544 \times 46.8 = 25.46\ KPa$

So total pressure of solution = $p_h + p_o$

= 47.97 + 25.46 = 73.43 KPa

Q2.17 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 :

It is asked the vapour pressure of 1 molal solution which means 1 mol of solute in 1000 g H_2O.

Moles in 1000g of water = 55.55 mol. (Since the molecular weight of H_2O is 18)

Mole fraction of solute :

$$\frac{1}{1+55.55} = 0.0177$$

Applying the equation :

$$\frac{p_w^\circ - p}{p_w^\circ} = x_2$$

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

or $p = 12.083\ KPa$

Thus the vapour pressure of the solution is 12.083 KPa

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Q2.18 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 the initial vapour pressure of octane = p_o° .

After adding solute to octane, the vapour pressure becomes :

$$= \frac{80}{100} \times p_o^\circ = 0.8p_o^\circ$$

Moles of octane :

$$= \frac{114}{114} = 1 \left(Molar\ mass\ of\ octane = 8(12) + 18(1) = 114\ g\ mol^{-1}\right)$$

Using Raoult's law we get :

$$\frac{p_o^\circ - p}{p_o^\circ} = x_2$$

or $$\frac{p_o^\circ - 0.8p_o^\circ}{p_o^\circ} = \frac{\frac{W}{40}}{\frac{W}{40} + 1}$$

or $w = 10\ g$

Thus required mass of non-volatile solute = 10g.

Q2.19(i) A solution containing 30 g of non-volatile solute exactly in $90\ g$ of water has a vapour pressure of $2.8\ k\ Pa$ at $298\ K.$. Further, $18\ g$ of water is then added to the solution and the new vapour pressure becomes $2.9\ kPa$ at $298\ K.$. Calculate:

molar mass of the solute

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Answer :

In this question we will find molar mass of solute by using Raoult's law .

Let the molar mass of solute is M.

Initially we have 30 g solute and 90 g water.

Moles of water :

$$\frac{90}{18} = 5\ mol$$

By Raoult's law we have :-

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

or $$\frac{p_w^\circ - 2.8}{p_w^\circ} = \frac{\frac{30}{M}}{5 + \frac{30}{M}} = \frac{30}{5M + 30}$$

or $$\frac{p_w^\circ}{2.8} = \frac{5M + 30}{5M}$$ ---------- (i)

Now we have added 18 g of water more, so the equation becomes:

Moles of H_2O :

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

Putting this in above equation we obtain :-

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

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$$\text{or } \frac{p_w^\circ}{2.9} = \frac{6M + 30}{6M}$$ ----------(ii)

From equation (i) and (ii) we get

M = 23 u

So the molar mass of solute is 23 units.

Q2.19(ii) 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

vapour pressure of water at 298 K.

Answer :

In the previous part we have calculated the value of molar mass the Raoul's law equation.

We had :-

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

Putting M = 23 u in the above equation we get,

$$\frac{p_w^\circ}{2.8} = \frac{5(23) + 30}{5(23)} = \frac{145}{115}$$

or $p_w^\circ = 3.53$

Thus vapour pressure of water = 3.53 kPa.

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Q2.20 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 :

It is given that freezing point of pure water is 273.15 K.

So, elevation of freezing point = 273.15 - 271 = 2.15 K

5% solution means 5 g solute in 95 g of water.

Moles of cane sugar :

$$= \frac{5}{342} = 0.0146$$

Molality :

$$= \frac{0.0146}{0.095} = 0.1537$$

We also know that - $\Delta T_f = k_f \times m$

or $k_f = \frac{\Delta T_f}{m} = 13.99\ K\ Kg\ mol^{-1}$

Now we will use the above procedure for glucose.

5% of glucose means 5 g of gluocse in 95 g of H_2O.

Moles of glucose :

$$\frac{5}{180} = 0.0278$$

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Thus molality :

$$= \frac{.0278}{0.095} = 0.2926\ mol\ kg^{-1}$$

So, we can find the elevation in freezing point:

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

$$= 13.99 \times 0.2926 = 4.09\ K$$

Thus freezing point of glucose solution is 273.15 - 4.09 = 269.06 K.

Q2.21 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 this question we will use the formula :

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

Firstly for compound AB_2 :-

$$M_B = \frac{K_f \times W_b \times 1000}{w_A \times \Delta T_f}$$

or $$= \frac{5.1 \times 1 \times 1000}{20 \times 2.3} = 110.87\ g/mol$$

Similarly for compound AB_4 :-

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$$M_B = \frac{5.1 \times 1 \times 1000}{20 \times 1.3} = 196\ g/mol$$

If we assume atomic weight of element A to be x and of element B to be y, then we have :-

x + 2y = 110.87 ---------- (i)

x + 4y = 196 ---------- (ii)

Solving both the equations, we get :-

x = 25.59 ; y = 42.6

Hence atomic mass of element A is 25.59u and atomic mass of element B is 42.6u.

Q2.22 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 :

According to given conditions we have same solution under same temperature. So we can write :

$$\frac{\Pi_1}{C_1} = \frac{\Pi_2}{C_2} \left(\Pi = C.R.T\ ; RT = \frac{\Pi}{C} \right)$$

So, if we put all the given values in above equation, we get

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

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or $C_2 = \frac{1.52 \times 36}{4.98 \times 180} = 0.061M$

Hence the required concentration is 0.061 M.

Q2.23(i) Suggest the most important type of intermolecular attractive interaction in the following pairs.

n-hexane and n-octane

Answer :

Since both the compounds are alkanes so their mixture has van der Waal force of attraction between compounds.

Q2.23(ii) Suggest the most important type of intermolecular attractive interaction in the following pairs.

$I_2 \; and \; CCl_4$

Answer :

The binary mixture of these compounds has van der Waal force of attraction between them.

Q2.23(iii) Suggest the most important type of intermolecular attractive interaction in the following pairs.

$NaClO_4 \; and \; Water$

Answer :

The given compounds will have ion-dipole interaction between them.

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Q2.23(iv) Suggest the most important type of intermolecular attractive interaction in the following pairs.

$methanol\ and\ acetone$

Answer :

Methanol has -OH group and acetone has ketone group. So there will be hydrogen bonding between them.

Q2.23(v) Suggest the most important type of intermolecular attractive interaction in the following pairs.

acetonitrile (CH_3CN) and acetone (C_3H_6O).

Answer :

They will have dipole-dipole interaction since both are polar compounds.

Q2.24 Based on solute-solvent interactions, arrange the following in order of increasing solubility in n-octane and explain.

Cyclohexane, KCl, CH3OH, CH3CN.

Answer :

The order will be : Cyclohexane > CH_3CN > CH_3OH > KCl

In this, we have used the fact that like dissolves like.

Since cyclohexane is an alkane so its solubility will be maximum.

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Q2.25(i) Amongst the following compounds, identify which are insoluble, partially soluble and highly soluble in water?

phenol

Answer :

We know the fact that like dissolves in like.

Since phenol is had both polar and non-polar group so it is partially soluble in water.

Q2.25(ii) Amongst the following compounds, identify which are insoluble, partially soluble and highly soluble in water?

toluene

Answer :

Since toluene is a non-polar compound i.e., it doesn't have any polar group so it is insoluble in water. (because water is a polar compound)

Q2.25(iii) Amongst the following compounds, identify which are insoluble, partially soluble and highly soluble in water?

formic acid

Answer :

Since the -OH group in formic acid (polar) can form H-bonds with water thus it is highly soluble in water.

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Q2.25(iv) Amongst the following compounds, identify which are insoluble, partially soluble and highly soluble in water?

ethylene glycol

Answer :

Ethylene glycol is an organic compound but is polar in nature. Also, it can form H-bonds with water molecules, thus it is highly soluble in water.

Q2.25(v) Amongst the following compounds, identify which are insoluble, partially soluble and highly soluble in water?

chloroform

Answer :

Chloroform is a non-polar compound so it is insoluble in water.

Q2.25(vi) Amongst the following compounds, identify which are insoluble, partially soluble and highly soluble in water?

pentanol

Answer :

Pentanol has both polar and non-polar groups so it is partially soluble in water.

Q2.26 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 molality of $Na^{+}\ ions$ in the lake.

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Answer :

We know that, Molality :

$$Molality = \frac{Moles\ of\ solute}{Mass\ of\ solvent\ in\ Kg}$$

So, for moles of solute we have :

$$Moles\ of\ Na^{+} = \frac{92}{23} = 4$$

Thus, molality :

$$= \frac{4}{1} = 4$$

Molality of Na $^{+}$ ions is 4m.

Q2.27 If the solubility product of CuS is 6×10^{-16} , calculate the maximum molarity of CuS in aqueous solution.

Answer :

We are given, $K_{sp} = 6 \times 10^{-16}$

The dissociation equation of CuS is given by :-

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

So, the equation becomes :- $K_{sp} = Cu^{2+} \times S^{2-}$

or $K_{sp} = s \times s = s^2$

or $s = 2.45 \times 10^{-8}\ M$

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Thus maximum molarity of solution is $2.45 \times 10^{-8}\ M$.

Q2.28 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 :

Total mass of solution = Mass of aspirin + Mass of acetonitrile = 6.5 + 450 = 456.5 g.

We know that :

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

So, $Mass\ percentage = \frac{6.5}{456.5} \times 100 = 1.42\%$

Thus the mass percentage of aspirin is 1.42%

Q2.29 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 :

We are given with molality of the solution, so we need to find the moles of Nalorphene.

Molar mass of nalorphene = 19(12) + 21(1) + 1(14) + 3(16) = 311u.

So moles of nalorphene :

$$\frac{1.5 \times 10^{-3}}{311} = 4.82 \times 10^{-6}\ moles$$

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Molality :

$$= \frac{No.\ of\ moles\ of\ solute}{Mass\ of\ solvent\ in\ kg}$$

or $1.5 \times 10^{-3} = \frac{4.82 \times 10^{-6}}{w}$

or $w = 3.2 \times 10^{-3}\ Kg$

So the required weight of water is 3.2 g.

Q2.30 Calculate the amount of benzoic acid (C_6H_5COOH) required for preparing $250\ mL$ of $0.15\ M$ solution in methanol.

Answer :

Molar mass of benzoic acid = 7(12) + 6(1) + 2(16) = 122u.

We are given with the molarity of solution.

$$Molarity = \frac{Moles\ of\ solute}{Volume\ of\ solution\ in\ litre}$$

or $0.15 = \frac{Moles\ of\ solute}{\frac{250}{1000}}$

or $Moles\ of\ solute = \frac{0.15 \times 250}{1000} = 0.0375\ mol$

So mass of benzoic acid :

$$= Moles\ of\ benzoic\ acid \times Molar\ mass$$

$$= 0.0375 \times 122 = 4.575\ g$$

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Hence the required amount of benzoic acid is 4.575 g.

Q2.31 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 :

We know that depression in freezing point of water will depend upon the degree of ionisation.

The degree of ionisation will be highest in the case of trifluoroacetic acid as it is most acidic among all three.

The order of degree of ionisation on the basis if acidic nature will be:- Trifluoroacetic acid > Trichloroacetic acid > Acetic acid.

So the depression in freezing point will be reverse of the above order.

Q2.32 Calculate the depression in the freezing point of water when $10\ g$ of $CH_3CH_2CHCICOOH$ is added to $250\ g$ of water.

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

Answer :

Firstly we will find the Vant's Hoff factor the dissociation of given compound.

	$CH_3CH_2CHClCOOH \leftrightarrow CH_3CH_2CHClCOO^- + H^+$		
Initial conc.	C mol L^{-1}	0	0
At equilibrium	$C(1-\alpha)$	$C\alpha$	$C\alpha$

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So we can write, $K_a = \frac{(Ca \times Ca)}{C(1-a)}$

or $K_a = Ca^2 (\because a << 1)$

or $a = \sqrt{\frac{K_a}{C}}$

Putting values of K_a and C in the last result, we get :

$a = 0.0655$

At equilibrium i = 1 - a + a + a = 1 + a = 1.0655

Now we need to find the moles of the given compound $CH_3CH_2CHClCOOH$.

So, moles =

$\frac{10}{122.5} = 0.0816\ mol$

Thus, molality of the solution :

$= \frac{0.0816 \times 1000}{250} = 0.3265\ m$

Now we will use :

$\Delta T_f = i\ K_f\ m$

or $= 1.065 \times 1.86 \times 0.3265 = 0.6467\ K$

Q2.33 $19.5\ g$ of CH_2FCOOH is dissolved in $500\ g$ of water. The depression in the freezing point of water observed is $1.0^\circ C$. Calculate the van't Hoff factor and dissociation constant of fluoroacetic acid.

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Answer :

Firstly we need to calculate molality in order to get vant's hoff factor.

So moles of CH_2FCOOH :

$$\frac{19.5}{78} = 0.25$$

We need to assume volume of solution to be nearly equal to 500 mL. (as 500 g water is present)

Now, we know that : $\Delta T_f = i\ K_f\ m$

or $i = \frac{1}{0.93} = 1.0753$

Now for dissociation constant :-

$$a = i - 1 = 1.0753 - 1 = 0.0753$$

and, $K_a = \frac{Ca^2}{1-a}$

Put values of C and a in the above equation, we get :

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

Q2.34 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 :

Firstly we will find number of moles of both water and glucose.

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Moles of glucose :

$$= \frac{25}{180} = 0.139\ mol$$

$$(Molar\ mass\ of\ glucose = 6(12) + 12(1) + 6(16) = 180\ g\ mol^{-1})$$

and moles of water :

$$= \frac{450}{18} = 25\ mol\ (Molar\ mass\ of\ water = 18\ g\ mol^{-1})$$

Now,

$$\frac{p_w^\circ - p}{p_w^\circ} = \frac{n_g}{n_g + n_w}$$

or $$\frac{17.535 - p}{17.535} = \frac{0.139}{0.139 + 25}$$

or $p = 17.44\ mm\ of\ Hg$

Thus vapour pressure of water after glucose addition = 17.44 mm of Hg

Q2.35 Henry's law constant for the molality of methane in benzene at $298\ K$ is $4.27 \times 10^5 mm\ Hg$. Calculate the solubility of methane in benzene at $298\ K$ under $760\ mm\ Hg$.

Answer :

We know that : $P = k \times C$

We are given value of P and k, so C can be found.

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

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Hence solubility of methane in benzene is 178×10^{-5} .

Q2.36 $100\ g$ 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 the solution is $475\ Torr$.

Answer :

For calculating partial vapour pressure we need to calculate mole fractions of components.

So number of moles of liquid A :

$$= \frac{100}{140} = 0.714$$

and moles of liquid B :

$$= \frac{1000}{180} = 5.556$$

Mole fraction of A (x_A) :

$$= \frac{0.714}{0.714 + 5.556} = 0.114$$

and mole fraction of B (x_B) :

$$= \frac{5.556}{0.714 + 5.556} = 0.866$$

Now, $P_{total} = P_A + P_B$

or $P_{total} = P_A^{\circ} x_A + P_B^{\circ} x_B$

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or $475 = P_A^\circ \times 0.114 \ + 500 \times 0.886$

or $P_A^\circ = 280.7 \ torr$

Thus vapour pressure in solution due to A = $P_A^\circ x_A$

$= 280.7 \times 0.114 = 32 \ torr$

Q2.37 Vapour pressures of pure acetone and chloroform at $328 \ K$ are $741.8 \ mm \ Hg$ and $638.8 \ mm \ Hg$ respectively. Assuming that they form ideal solution over the entire range of composition, plot ptotal, pchloroform, and pacetone as a function of xacetone. The experimental data observed for different compositions of mixture 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
$p_{chloroform}/mm \ Hg$	632.8	548.1	469.4	359.7	257.7	193.6	161.2	120.7
p_{total}	632.8	603	579.5	562.1	580.4	599.5	615.5	64.18

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

Answer :

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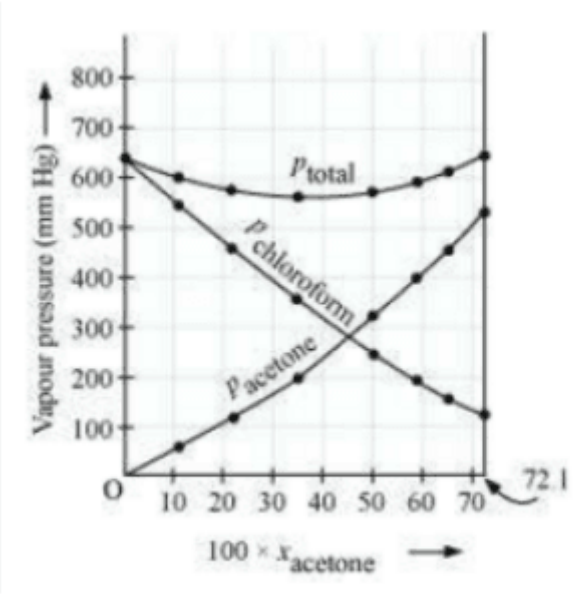

it has negative deviation from the ideal solution.

Q2.38 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 $100\ g$ toluene.

Answer :

Firstly, we will find the no. of moles of the given compounds.

No. of moles of benzene :

$$= \frac{80}{78} = 1.026\ mol$$

and the no. of moles of toluene :

$$= \frac{100}{92} = 1.087\ mol$$

.

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Now we will find mol fraction of both:-

Mole fraction of benzene :-

$$= \frac{1.026}{1.026 + 1.087} = 0.486$$

and mole fraction of toluene :

$$= 1 - 0.486 = 0.514$$

Now,

$P_{total} = P_b + P_t$

or $= 50.71 \times 0.486 \ + 32.06 \times 0.514 \ = 24.65 + 16.48$

or $= 41.13 \ mm \ of \ Hg$

Hence mole fraction of benzene in vapour phase is given by :

$$= \frac{24.65}{41.13} = 0.60$$

Q2.39 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 $3.30 \times 107 \ mm$ and $6.51 \times 107 \ mm$ respectively, calculate the composition of these gases in water

Answer :

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We have been given that the water is in equilibrium with air at a pressure of 10 atm or 7600 mm of Hg.

So the partial pressure of oxygen :

$$\frac{20}{100} \times 7600 = 1520 \ mm \ of \ Hg$$

and partial pressure of nitrogen :

$$\frac{79}{100} \times 7600 = 6004 \ mm \ of \ Hg$$

Now, by Henry's Law :

$$P = K_h.x$$

For oxygen :

$$x = \frac{1520}{3.30 \times 10^7} = 4.61 \times 10^{-5}$$

For nitrogen :

$$x = \frac{6004}{6.51 \times 10^7} = 9.22 \times 10^{-5}$$

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

Q2.41 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 $27^\circ C$.

Answer :

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We know that osmotic pressure :

$$\Pi = i\ (\frac{n}{v})\ R\ T$$

or $$\Pi = i\ (\frac{w}{M\ v})\ R\ T$$

We have been given the values of osmotic pressure, V, i and T.

So the value of w can be found.

$$w = \frac{0.75 \times 111 \times 2.5}{2.47 \times 0.0821 \times 300} (M = 1 \times 40 + 2 \times 35.5 = 111\ g\ mol^{-1})$$

$$= 3.42\ g$$

Hence 3.42 g $CaCl_2$ is required.

Q2.41 Determine the osmotic pressure of a solution prepared by dissolving $25\ mg$ of K_2SO_4 in $2\ litre$ of water at $25^{\circ}C$, assuming that it is completely dissociated.

Answer :

Dissociation of K_2SO_4 is as follows :-

$$K_2SO_{4(s)} \rightarrow 2\ K^+_{(aq)} + SO_4^{2-}{}_{(aq)}$$

It is clear that 3 ions are produced, so the value of i will be 3.

Molecular weight of K_2SO_4 = 2(39) + 1(32) + 4(16) = 174u.

$$\Pi = i\ C\ R\ T$$

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Putting all the values :-

$$\Pi = \frac{3 \times 25 \times 10 - 3 \times 0.082 \times 298}{174 \times 2}$$

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

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