Chapter 1: Solutions

Exercises — Exercises

Q1 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 pure substances whose composition can be varied within certain limits.

The substance present in a smaller amount is called the solute and the one present in a larger amount is called the solvent.

Nine types of solutions are possible based on the physical states of solute and solvent:

1. Gas in Gas — the solute and solvent are both gases. Example: air (O_2 dissolved in N_2).
2. Liquid in Gas — liquid solute dispersed in gaseous solvent. Example: water vapour in air.
3. Solid in Gas — solid solute in gaseous solvent. Example: camphor vapours in $N_2(g)$, iodine vapour in air.
4. Gas in Liquid — gaseous solute dissolved in liquid solvent. Example: CO_2 dissolved in water (soda water), O_2 dissolved in water.
5. Liquid in Liquid — liquid solute dissolved in liquid solvent. Example: ethanol dissolved in water.
6. Solid in Liquid — solid solute dissolved in liquid solvent. Example: glucose dissolved in water, NaCl in water.
7. Gas in Solid — gas dissolved in a solid. Example: hydrogen dissolved in palladium.
8. Liquid in Solid — liquid dispersed in a solid matrix. Example: amalgam of mercury with sodium.
9. Solid in Solid — solid solute dissolved in solid solvent. Example: copper dissolved in gold (alloy).

Answer: Nine types of solutions: gas-in-gas (e.g., air), liquid-in-gas (e.g., water vapour in air), solid-in-gas (e.g., camphor in N_2), gas-in-liquid (e.g., CO_2 in water), liquid-in-liquid (e.g., ethanol in water), solid-in-liquid (e.g., NaCl in water), gas-in-solid (e.g., H_2 in Pd), liquid-in-solid (e.g., Hg in Na amalgam), solid-in-solid (e.g., Cu in Au alloy).

Q2 Give an example of a solid solution in which the solute is a gas.

SOLUTION

A solid solution is one in which the solvent is in the solid state.
When the solute is a gas dissolved in a solid solvent, the result is a gas-in-solid solution.

Example: Solution of hydrogen gas (H_2) in palladium (Pd) metal.
Hydrogen gas is adsorbed and dissolved within the lattice of palladium, forming a solid solution.

Another example: O_2 dissolved in titanium metal.

Answer: Hydrogen (H_2) dissolved in palladium (Pd) is a classic example of a solid solution where the solute is a gas.

Q3 Define the following terms:
(i) Mole fraction
(ii) Molality
(iii) Molarity
(iv) Mass percentage.

SOLUTION

(i) Mole Fraction (x):
The mole fraction of a component in a solution is the ratio of the number of moles of that component to the total number of moles of all components in the solution.
For a binary solution of A and B: $x_A = n_A / (n_A + n_B)$, where n_A and n_B are moles of A and B respectively.
Mole fractions are dimensionless and $x_A + x_B = 1$.

(ii) Molality (m):
Molality is defined as the number of moles of solute dissolved per kilogram of solvent.
m = (moles of solute) / (mass of solvent in kg)

Unit: mol kg^{-1} (or m). Molality is independent of temperature.

(iii) Molarity (M):
Molarity is defined as the number of moles of solute dissolved per litre (dm^3) of solution.
M = (moles of solute) / (volume of solution in litres)
Unit: mol L^{-1} (or M). Molarity changes with temperature since volume changes.

(iv) Mass Percentage (w/w %):
Mass percentage of a component is the mass of that component per 100 g of the solution.
Mass % of A = [mass of A / (mass of A + mass of B)] × 100
It is dimensionless (or expressed as %). Mass percentage is temperature-independent.

Answer: (i) Mole fraction: x_A = n_A/(n_A+n_B); dimensionless, sum = 1. (ii) Molality: moles of solute per kg of solvent; mol kg^{-1}; temperature-independent. (iii) Molarity: moles of solute per litre of solution; mol L^{-1}; temperature-dependent. (iv) Mass percentage: mass of component per 100 g of solution; temperature-independent.

Q4 Concentrated nitric acid used in laboratory work is 68% nitric acid by mass in aqueous solution. What should be the molarity of such a sample of the acid if the density of the solution is 1.504 g mL^{-1}?

SOLUTION

Given:
Mass percentage of HNO_3 = 68%
Density of solution = 1.504 g mL^{-1}
Molar mass of HNO_3 = 1 + 14 + 3×16 = 63 g mol^{-1}

Consider 1 L (1000 mL) of solution:
Mass of solution = volume × density = 1000 mL × 1.504 g mL^{-1} = 1504 g

Mass of HNO_3 in 1 L = 1504 × (68/100) = 1022.72 g

Moles of HNO_3 = 1022.72 / 63 = 16.23 mol

Since these moles are present in 1 L of solution:
Molarity = 16.23 mol L^{-1}

Answer: Molarity of concentrated HNO_3 = 16.23 mol L^{-1}

Q5 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?

SOLUTION

Given: 10% w/w glucose solution; density = 1.2 g mL^{-1}
M(glucose, $C_6H_{12}O_6$) = 180 g mol^{-1}; M(water) = 18 g mol^{-1}

In 100 g of solution: mass of glucose = 10 g, mass of water = 90 g

Moles of glucose = 10/180 = 0.0556 mol
Moles of water = 90/18 = 5.000 mol

Molality:
m = moles of glucose / mass of water (kg) = 0.0556 / 0.090 = 0.617 mol kg^{-1}

Mole Fractions:
Total moles = 0.0556 + 5.000 = 5.0556
x(glucose) = 0.0556 / 5.0556 = 0.0110
x(water) = 5.000 / 5.0556 = 0.9890

Molarity (density = 1.2 g mL^{-1}):
Mass of 1 L solution = 1000 × 1.2 = 1200 g
Mass of glucose in 1 L = 1200 × 0.10 = 120 g
Moles of glucose = 120/180 = 0.667 mol
Molarity = 0.667 mol L^{-1}

Answer: Molality = 0.617 mol kg^{-1}; x(glucose) = 0.0110, x(water) = 0.9890; Molarity = 0.667 mol L^{-1}

Q6 How many mL of 0.1 M HCl are required to react completely with 1 g mixture of Na_2CO_3 and $NaHCO_3$ containing equimolar amounts of both?

SOLUTION

Given: 1 g mixture of Na_2CO_3 and $NaHCO_3$ in equimolar amounts.
M(Na_2CO_3) = 106 g mol^{-1}; M($NaHCO_3$) = 84 g mol^{-1}

Let the number of moles of each = n mol.
Total mass: 106n + 84n = 1 g
190n = 1 => n = 1/190 = 0.005263 mol

Reactions with HCl:
$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$ (requires 2 mol HCl per mol Na_2CO_3)

$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$ (requires 1 mol HCl per mol $NaHCO_3$)

Moles of HCl required:
= 2n (for Na_2CO_3) + 1n (for $NaHCO_3$)
= 3n = 3 × (1/190) = 3/190 = 0.01579 mol

Volume of 0.1 M HCl:
V = moles / molarity = 0.01579 / 0.1 = 0.1579 L = 157.9 mL

Answer: Volume of 0.1 M HCl required = 157.9 mL

Q7 A solution is obtained by mixing 300 g of 25% solution and 400 g of 40% solution by mass. Calculate the mass percentage of the resulting solution.

SOLUTION

Mass of solute in 300 g of 25% solution = 300 × 25/100 = 75 g
Mass of solute in 400 g of 40% solution = 400 × 40/100 = 160 g

Total mass of solute = 75 + 160 = 235 g
Total mass of resulting solution = 300 + 400 = 700 g

Mass percentage = (235/700) × 100 = 33.57%

Answer: Mass percentage of the resulting solution = 33.57%

Q8 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?

SOLUTION

Given: mass of ethylene glycol = 222.6 g; mass of water = 200 g; density = 1.072 g mL^{-1}
$M(C_2H_6O_2) = 2\times12 + 6\times1 + 2\times16 = 62$ g mol^{-1}

Moles of ethylene glycol = 222.6 / 62 = 3.590 mol

Molality:
m = 3.590 mol / (200/1000 kg) = 3.590 / 0.200 = 17.95 mol kg^{-1}

Molarity:
Total mass of solution = 222.6 + 200 = 422.6 g
Volume of solution = 422.6 / 1.072 g mL^{-1} = 394.22 mL = 0.39422 L

Molarity = 3.590 / 0.39422 = 9.11 mol L^{-1}

Answer: Molality = 17.95 mol kg^{-1}; Molarity = 9.11 mol L^{-1}

Q9 A sample of drinking water was found to be severely contaminated with chloroform ($CHCl_3$) supposed to be a carcinogen. The level of contamination was 15 ppm (by mass):
(i) express this in percent by mass
(ii) determine the molality of chloroform in the water sample.

SOLUTION

Given: chloroform contamination = 15 ppm by mass
$M(CHCl_3) = 12 + 1 + 3\times35.5 = 119.5$ g mol^{-1}

15 ppm by mass means 15 g $CHCl_3$ per 10^6 g solution (or per 10^6 g water, since $CHCl_3$ << water).

(i) Percent by mass:
$\% = (15 / 10^6) \times 100 = 1.5 \times 10^{-3}$ %

(ii) Molality:
In 10^6 g solution: 15 g $CHCl_3$ and approximately $(10^6 - 15) \approx 10^6$ g water ≈ 1000 kg water
Moles of $CHCl_3$ = 15 / 119.5 = 0.1255 mol
Molality = 0.1255 mol / 1000 kg = 1.25×10^{-4} mol kg^{-1}

Answer: (i) 1.5×10^{-3} % by mass; (ii) Molality = 1.25×10^{-4} mol kg^{-1}

Q10 What role does the molecular interaction play in a solution of alcohol and water?

SOLUTION

Alcohol (e.g., ethanol, C_2H_5OH) and water (H_2O) are both polar molecules capable of forming hydrogen bonds.

In pure ethanol: O-H...O hydrogen bonds exist between ethanol molecules.
In pure water: O-H...O hydrogen bonds exist between water molecules.

When ethanol is mixed with water:
- New O-H...O hydrogen bonds form between ethanol and water molecules.
- These ethanol-water interactions are weaker than the original water-water and ethanol-ethanol hydrogen bonds.

Effect on solution properties:

- The intermolecular forces in the mixture are slightly weaker than in the pure components.
- This results in a slight positive deviation from Raoult's law.
- The volume of the mixture is slightly less than the sum of volumes of pure components (volume contraction).
- The enthalpy of mixing (Delta_mix H) is slightly positive (endothermic mixing).

However, because ethanol and water both form hydrogen bonds, they are completely miscible and the deviations are not large.

Answer: Ethanol-water interactions (H-bonds) are weaker than pure water-water and ethanol-ethanol H-bonds, causing slight positive deviation from Raoult's law, small volume contraction, and small positive ΔmixH.

Q11 Why do gases always tend to be less soluble in liquids as the temperature is raised?

SOLUTION

The dissolution of a gas in a liquid is an exothermic process:
Gas + Solvent -> Solution (Delta H < 0)

When a gas dissolves in a liquid, gas molecules lose kinetic energy by interacting with solvent molecules, releasing heat.

By Le Chatelier's principle:
When temperature is increased, the equilibrium shifts in the direction that absorbs heat — i.e., in the reverse direction (toward the gas phase).

Therefore, increasing temperature causes gas molecules to escape from the solution back into the gas phase, reducing solubility.

In thermodynamic terms, at higher temperatures the gas molecules have greater kinetic energy, allowing them to overcome intermolecular attractions with the solvent and escape more readily.

This is why carbonated beverages lose their CO_2 faster when warm than when cold.

Answer: Gas dissolution is exothermic. By Le Chatelier's principle, raising temperature shifts equilibrium toward the gas phase, decreasing solubility. Higher temperature gives gas molecules more kinetic energy to escape the solvent.

Q12 State Henry's law and mention some important applications.

SOLUTION

Henry's Law Statement:

At a constant temperature, the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas above the liquid.

Mathematically: $p = K_H \times x$

where p = partial pressure of the gas, x = mole fraction of the gas in solution, and K_H = Henry's law constant (units: pressure, e.g. bar or atm).

Important characteristics:

- K_H is different for each gas and each solvent, and increases with temperature (gases become less soluble at higher T).
- Henry's law is obeyed at low pressures and for gases that do not react chemically with the solvent.

Important Applications:

1. Carbonated beverages: CO_2 is dissolved in soft drinks under high pressure. When the bottle is opened, pressure drops and CO_2 escapes (effervescence).

2. Deep-sea diving / Scuba diving: At high pressures underwater, more N_2 dissolves in the blood. If a diver ascends too quickly, dissolved N_2 forms bubbles in the blood, causing 'bends' (decompression sickness). Divers use O_2-He mixtures to avoid this problem since He has a much lower K_H in blood.

3. Respiration at high altitudes: At high altitudes the partial pressure of O_2 is low, so less O_2 dissolves in blood, causing altitude sickness.

4. Functioning of lungs: In alveoli, O_2 under partial pressure dissolves in blood (high-pO_2) and CO_2 is released from blood (low-pCO_2 in lungs).

Answer: Henry's law: $p = K_H \times x$ (partial pressure proportional to mole fraction in solution). Applications: (1) CO_2 in carbonated drinks, (2) prevention of 'bends' in scuba diving using He-O_2 mix, (3) altitude sickness (low pO_2 at high altitude), (4) gas exchange in lungs.

Q13 The partial pressure of ethane over a solution containing 6.56×10^{-3} g of ethane is 1 bar. If the solution contains 5.00×10^{-2} g of ethane, then what shall be the partial pressure of the gas?

SOLUTION

By Henry's law, for a dilute solution at constant temperature and constant volume of solvent:
$p \propto$ moles dissolved $\propto$ mass dissolved (same molar mass)

Given:
$m_1 = 6.56 \times 10^{-3}$ g of ethane, $p_1 = 1$ bar
$m_2 = 5.00 \times 10^{-2}$ g of ethane, $p_2 = ?$

$p_2 / p_1 = m_2 / m_1$
$p_2 = p_1 \times (m_2/m_1) = 1 \times (5.00 \times 10^{-2}) / (6.56 \times 10^{-3})$
$p_2 = 1 \times 7.622 = 7.62$ bar

Answer: Partial pressure of ethane = 7.62 bar

Q14 What is meant by positive and negative deviations from Raoult's law and how is the sign of DeltamixH related to positive and negative deviations from Raoult's law?

SOLUTION

Raoult's Law for an ideal solution: p_total = x_A × p*_A + x_B × p*_B
where p*_A and p*_B are vapour pressures of pure components.

Positive Deviation:
When the observed vapour pressure is GREATER than that predicted by Raoult's law:
p_total > x_A p*_A + x_B p*_B

Cause: Solute-solvent (A-B) interactions are WEAKER than solute-solute (A-A) and solvent-solvent (B-B) interactions.
Result: Molecules escape more readily into vapour phase => higher vapour pressure.
Delta_mix H > 0 (endothermic mixing) — energy is absorbed to break stronger A-A and B-B interactions.
Example: ethanol + water; acetone + carbon disulfide.

Negative Deviation:

When the observed vapour pressure is LESS than that predicted by Raoult's law: $p_{total} < x_A p^*_A + x_B p^*_B$

Cause: Solute-solvent (A-B) interactions are STRONGER than A-A and B-B interactions.
Result: Molecules have less tendency to escape => lower vapour pressure.
$\Delta_{mix} H < 0$ (exothermic mixing) — energy is released when stronger A-B interactions form.
Example: chloroform + acetone (H-bond between $CHCl_3$ and C=O); HNO_3 + water.

Summary:
Positive deviation => $\Delta_{mix} H > 0$ (endothermic)
Negative deviation => $\Delta_{mix} H < 0$ (exothermic)

Answer: Positive deviation: $p_{obs} >$ Raoult prediction; A-B forces weaker than A-A and B-B; $\Delta mixH > 0$ (endothermic). Negative deviation: $p_{obs} <$ Raoult prediction; A-B forces stronger; $\Delta mixH < 0$ (exothermic).

Q15 An aqueous solution of 2% non-volatile solute exerts a pressure of 1.004 bar at the normal boiling point of the solvent. What is the molar mass of the solute?

SOLUTION

At the normal boiling point of water, p^*(water) = 1 atm = 1.013 bar.
Observed vapour pressure p = 1.004 bar.

By Raoult's law:
$(p^* - p) / p^* = x_{solute}$
$x_{solute} = (1.013 - 1.004) / 1.013 = 0.009 / 1.013 = 0.00888$

For 2% w/w solution: 100 g solution contains 2 g solute and 98 g water.
Moles of water = 98 / 18 = 5.444 mol

For dilute solution: $x_{solute} \approx n_{solute} / n_{water}$
$n_{solute} = x_{solute} \times n_{water} = 0.00888 \times 5.444 = 0.0484$ mol

Molar mass of solute:
M = mass / moles = 2 / 0.0484 = 41.3 g mol^{-1}

Answer: Molar mass of the solute = 41.3 g mol^{-1}

Q16 Heptane and octane form an ideal solution. At 373 K, the vapour pressures of the two liquid components are 105.2 kPa and 46.8 kPa respectively. What will be the vapour pressure of a mixture of 26.0 g of heptane and 35 g of octane?

SOLUTION

Given:
p*(heptane) = 105.2 kPa; M(heptane, C_7H_{16}) = 100 g mol^{-1}
p*(octane) = 46.8 kPa; M(octane, C_8H_{18}) = 114 g mol^{-1}
Mass of heptane = 26.0 g; mass of octane = 35.0 g

Moles of heptane = 26.0 / 100 = 0.260 mol
Moles of octane = 35.0 / 114 = 0.307 mol
Total moles = 0.260 + 0.307 = 0.567 mol

Mole fractions:
x(heptane) = 0.260 / 0.567 = 0.4585
x(octane) = 0.307 / 0.567 = 0.5415

Partial pressures (Raoult's law, ideal solution):
p(heptane) = x(heptane) × p*(heptane) = 0.4585 × 105.2 = 48.24 kPa
p(octane) = x(octane) × p*(octane) = 0.5415 × 46.8 = 25.34 kPa

Total vapour pressure:
p_total = 48.24 + 25.34 = 73.58 kPa

Answer: Total vapour pressure of the mixture = 73.58 kPa

Q17 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.

SOLUTION

Given: p*(water) = 12.3 kPa; molality = 1 mol kg^{-1}

1 molal solution: 1 mol solute per 1000 g (55.56 mol) of water.
Moles of water = 1000/18 = 55.56 mol

Mole fraction of water:
x(water) = 55.56 / (55.56 + 1) = 55.56 / 56.56 = 0.9823

By Raoult's law:
p(water) = x(water) × p*(water) = 0.9823 × 12.3 = 12.08 kPa

Answer: Vapour pressure of the 1 molal solution = 12.08 kPa

Q18 Calculate the mass of a non-volatile solute (molar mass 40 g mol^{-1}) which should be dissolved in 114 g octane to reduce its vapour pressure to 80%.

SOLUTION

Given: Vapour pressure is reduced to 80% => p = 0.80 p*
M(solute) = 40 g mol^{-1}; mass of octane = 114 g
M(octane) = 114 g mol^{-1}

By Raoult's law:
(p* - p) / p* = x_solute
(p* - 0.80p*) / p* = 0.20 = x_solute

Moles of octane = 114/114 = 1.000 mol

x_solute = n_solute / (n_solute + n_octane)
0.20 = n_solute / (n_solute + 1)
0.20(n_solute + 1) = n_solute
0.20 = n_solute(1 - 0.20) = 0.80 × n_solute
n_solute = 0.20/0.80 = 0.25 mol

Mass of solute = 0.25 × 40 = 10 g

Answer: Mass of non-volatile solute required = 10 g

Q19 A solution containing 30 g of non-volatile solute exactly in 90 g of water has a vapour pressure of 2.8 kPa at 298 K. Further, 18 g of water is then added to the solution and the new vapour pressure becomes 2.9 kPa at 298 K. Calculate:
(i) molar mass of the solute
(ii) vapour pressure of water at 298 K.

SOLUTION

Let M = molar mass of solute, n_s = moles of solute = 30/M, p* = VP of pure water.

State 1: 30 g solute in 90 g water.
n_water,1 = 90/18 = 5 mol
By Raoult's law: (p* - 2.8)/p* = n_s/(n_s + 5) ... (1)
=> 2.8/p* = 5/(n_s + 5) ... (A)

State 2: 30 g solute in (90 + 18) = 108 g water.
n_water,2 = 108/18 = 6 mol
By Raoult's law: (p* - 2.9)/p* = n_s/(n_s + 6) ... (2)

=> 2.9/p* = 6/(n_s + 6) ... (B)

Dividing (A) by (B):
2.8/2.9 = [5/(n_s + 5)] / [6/(n_s + 6)]
2.8/2.9 = 5(n_s + 6) / [6(n_s + 5)]

Cross-multiply:
2.8 × 6 × (n_s + 5) = 2.9 × 5 × (n_s + 6)
16.8(n_s + 5) = 14.5(n_s + 6)
16.8 n_s + 84 = 14.5 n_s + 87
2.3 n_s = 3
n_s = 3/2.3 = 1.304 mol

(i) Molar mass of solute:
M = 30/n_s = 30/1.304 = 23.0 g mol^{-1}

(ii) Vapour pressure of water:
From equation (A): p* = 2.8 × (n_s + 5) / 5 = 2.8 × (1.304 + 5) / 5
p* = 2.8 × 6.304 / 5 = 17.651 / 5 = 3.53 kPa

Answer: (i) Molar mass of solute = 23.0 g mol^{-1}; (ii) Vapour pressure of pure water at 298 K = 3.53 kPa

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Q20 A 5% solution (by mass) of cane sugar in water has freezing point of 271 K. Calculate the freezing point of 5% glucose in water if freezing point of pure water is 273.15 K.

SOLUTION

Given: 5% cane sugar has FP = 271 K; FP of pure water = 273.15 K.
M(cane sugar, $C_{12}H_{22}O_{11}$) = 342 g mol^{-1}; M(glucose, $C_6H_{12}O_6$) = 180 g mol^{-1}

Step 1: Find Kf using cane sugar data.
In 100 g solution: 5 g cane sugar + 95 g water
n(cane sugar) = 5/342 = 0.01462 mol
Molality = 0.01462/0.095 = 0.1539 mol kg^{-1}

ΔTf(cane sugar) = 273.15 - 271 = 2.15 K
Kf = ΔTf / m = 2.15 / 0.1539 = 13.97 K kg mol^{-1}

Step 2: Apply Kf to 5% glucose solution.
In 100 g solution: 5 g glucose + 95 g water
n(glucose) = 5/180 = 0.02778 mol

Molality = 0.02778/0.095 = 0.2924 mol kg^{-1}

ΔTf(glucose) = Kf × m = 13.97 × 0.2924 = 4.085 K

Freezing point of glucose solution:
FP = 273.15 - 4.085 = 269.06 K

Answer: Freezing point of 5% glucose solution = 269.06 K

Q21 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.

SOLUTION

Given: Kf(benzene) = 5.1 K kg mol^{-1}; mass of benzene = 20 g = 0.020 kg

Using: M = Kf × w_solute / (ΔTf × w_solvent)

Molar mass of AB_2:
$M(AB_2)$ = 5.1 × 1 / (2.3 × 0.020) = 5.1 / 0.046 = 110.87 g mol^{-1}

Molar mass of AB_4:
$M(AB_4)$ = 5.1 × 1 / (1.3 × 0.020) = 5.1 / 0.026 = 196.15 g mol^{-1}

Setting up equations:
A + 2B = 110.87 ... (1)
A + 4B = 196.15 ... (2)

Subtracting (1) from (2):
2B = 196.15 - 110.87 = 85.28
B = 42.64 ≈ 42.64 u

Substituting back:
A = 110.87 - 2 × 42.64 = 110.87 - 85.28 = 25.59 u

Atomic mass of A ≈ 25.6 u (close to Mg = 24.3 u)
Atomic mass of B ≈ 42.6 u (close to Ca = 40.1 u)

Answer: Atomic mass of A ≈ 25.6 u; Atomic mass of B ≈ 42.6 u

Q22 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 bar at the same temperature, what would be its concentration?

SOLUTION

By van't Hoff equation: $\pi = CRT$

At constant temperature T, $\pi \propto C$.

Given: First solution: 36 g glucose in 1 L => $C_1 = 36/180 = 0.20$ mol L^{-1}, $\pi 1 = 4.98$ bar

Second solution: $\pi 2 = 1.52$ bar, $C_2 = ?$

$C_2 / C_1 = \pi 2 / \pi 1$

$C_2 = C_1 \times (\pi 2/\pi 1) = 0.20 \times (1.52/4.98)$

$C_2 = 0.20 \times 0.3052 = 0.0610$ mol L^{-1}

Answer: Concentration of the second solution = 0.0610 mol L^{-1}

Q23 Suggest the most important type of intermolecular attractive interaction in the following pairs.

(i) n-hexane and n-octane

(ii) I_2 and CCl_4

(iii) $NaClO_4$ and water

(iv) methanol and acetone

(v) acetonitrile (CH_3CN) and acetone (C_3H_6O).

SOLUTION

(i) n-hexane and n-octane:

Both are non-polar hydrocarbons (no dipole moment). The dominant intermolecular attraction is London dispersion forces (van der Waals forces).

(ii) I_2 and CCl_4:

Both are non-polar molecules (I_2 is non-polar; CCl_4 is tetrahedral and non-polar). The dominant interaction is London dispersion forces (van der Waals / induced dipole-induced dipole forces).

(iii) $NaClO_4$ and water:

$NaClO_4$ is an ionic compound (Na^+ and ClO_4^- ions). Water is a polar solvent. The dominant interaction is ion-dipole interactions — the ions are stabilised by the polar water molecules (hydration/solvation).

(iv) Methanol and acetone:
Methanol (CH_3OH) has an O-H group capable of donating an H-bond. Acetone (CH_3COCH_3) has a C=O group that can accept an H-bond. The dominant interaction is hydrogen bonding (O-H...O=C).

(v) Acetonitrile (CH_3CN) and acetone (C_3H_6O):
Both are polar molecules (CH_3CN has a C≡N dipole; acetone has a C=O dipole). Neither has an O-H or N-H to donate a classical H-bond. The dominant interaction is dipole-dipole interactions.

Answer: (i) London dispersion forces; (ii) London dispersion forces; (iii) Ion-dipole interactions; (iv) Hydrogen bonding (O-H...O=C); (v) Dipole-dipole interactions.

Q24 Based on solute-solvent interactions, arrange the following in order of increasing solubility in n-octane and explain. Cyclohexane, KCl, CH_3OH, CH_3CN.

SOLUTION

n-Octane is a non-polar hydrocarbon solvent. The principle 'like dissolves like' applies — non-polar or weakly polar solutes dissolve well in non-polar solvents, while ionic/strongly polar solutes do not.

Analysis of each solute:

KCl: ionic compound with strong interionic forces. n-Octane cannot solvate ions (no dipole). Essentially INSOLUBLE in n-octane.

CH_3OH (methanol): polar, capable of H-bonding with itself. Poor interaction with non-polar octane. Very sparingly soluble (very slightly soluble).

CH_3CN (acetonitrile): polar molecule (C≡N dipole) but no O-H or N-H for self-association. Has some limited interaction with octane. Slightly more soluble than methanol in n-octane.

Cyclohexane (C_6H_{12}): non-polar cycloalkane. Interacts well with n-octane through London dispersion forces. Highly soluble / miscible in all proportions with n-octane.

Order of increasing solubility in n-octane:
KCl < CH_3OH < CH_3CN < Cyclohexane

Answer: Increasing solubility in n-octane: KCl < CH_3OH < CH_3CN < Cyclohexane. (Non-polar cyclohexane is most soluble; ionic KCl is essentially insoluble.)

Q25 Amongst the following compounds, identify which are insoluble, partially soluble and highly soluble in water?
(i) phenol
(ii) toluene
(iii) formic acid
(iv) ethylene glycol
(v) chloroform
(vi) pentanol.

SOLUTION

Water is a polar, H-bonding solvent. Solubility depends on solute-water interactions.

(i) Phenol (C_6H_5OH):
Has an -OH group capable of H-bonding with water, but also a large non-polar benzene ring. Partially soluble in water (forms a two-phase system above ~8.3 g/100 mL at 20°C).

(ii) Toluene ($C_6H_5CH_3$):
Non-polar hydrocarbon; no groups for H-bonding with water. Essentially insoluble in water.

(iii) Formic acid (HCOOH):
Polar molecule; can form H-bonds with water through both C=O and O-H groups. Highly soluble (completely miscible with water) in all proportions.

(iv) Ethylene glycol ($HOCH_2CH_2OH$):
Two -OH groups forming strong H-bonds with water. Highly soluble (completely miscible with water).

(v) Chloroform ($CHCl_3$):
Slightly polar (net dipole due to three Cl atoms), but cannot form H-bonds as a donor. Insoluble in water (or very sparingly soluble; miscible with organic solvents).

(vi) Pentanol ($C_5H_{11}OH$):
Has one -OH group (H-bonding), but also a long non-polar C_5 chain. The non-polar part dominates. Partially soluble in water.

Answer: Highly soluble: formic acid (iii), ethylene glycol (iv). Partially soluble: phenol (i), pentanol (vi). Insoluble: toluene (ii), chloroform (v).

Q26 If the density of some lake water is 1.25 g mL^{-1} and contains 92 g of Na^+ ions per kg of water, calculate the molarity of Na^+ ions in the lake.

SOLUTION

Given: density = 1.25 g mL^{-1}; Na^+ content = 92 g per kg water.
$M(Na^+) = 23\ g\ mol^{-1}$

Moles of Na^+ per kg water = 92/23 = 4.00 mol

Mass of this solution (1 kg water + 92 g Na^+):
Total mass = 1000 + 92 = 1092 g

Volume of this solution:
V = mass/density = 1092/1.25 = 873.6 mL = 0.8736 L

Molarity of Na^+ = moles / volume (L) = 4.00 / 0.8736 = 4.58 mol L^{-1}

Answer: Molarity of Na^+ ions = 4.58 mol L^{-1}

Q27 If the solubility product of CuS is 6×10^{-16}, calculate the maximum molarity of CuS in aqueous solution.

SOLUTION

CuS dissociates in water as:
$CuS(s) \rightleftharpoons Cu^{2+}(aq) + S^{2-}(aq)$

Let solubility = s mol L^{-1}
$[Cu^{2+}] = s,\ [S^{2-}] = s$

$K_{sp} = [Cu^{2+}][S^{2-}] = s \times s = s^2$
$s^2 = 6 \times 10^{-16}$
$s = \sqrt{(6 \times 10^{-16})} = \sqrt{6} \times 10^{-8} = 2.449 \times 10^{-8}$ mol L^{-1}

Maximum molarity of CuS = 2.45×10^{-8} mol L^{-1}

Answer: Maximum molarity of CuS = 2.45×10^{-8} mol L^{-1}

Q28 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.

SOLUTION

Given: mass of aspirin ($C_9H_8O_4$) = 6.5 g; mass of acetonitrile (CH_3CN) = 450 g

Total mass of solution = 6.5 + 450 = 456.5 g

Mass percentage of aspirin:
w/w % = (mass of aspirin / total mass of solution) × 100
= (6.5 / 456.5) × 100
= 1.42%

Answer: Mass percentage of aspirin in acetonitrile = 1.42%

Q29 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.

SOLUTION

Given: dose = 1.5 mg = 1.5×10^{-3} g; molality of solution = 1.5×10^{-3} mol kg^{-1}
$M(C_{19}H_{21}NO_3) = 19\times12 + 21\times1 + 14 + 3\times16 = 228 + 21 + 14 + 48 = 311$ g mol^{-1}

Moles of nalorphene in the dose:
$n = 1.5 \times 10^{-3} / 311 = 4.823 \times 10^{-6}$ mol

In a 1.5×10^{-3} m solution: 1.5×10^{-3} mol nalorphene per 1000 g water.
Mass of nalorphene per 1000 g water = $1.5 \times 10^{-3} \times 311 = 0.4665$ g
Mass of solution per 1.5×10^{-3} mol nalorphene = 1000 + 0.4665 = 1000.47 g

Mass of solution needed for $n = 4.823 \times 10^{-6}$ mol nalorphene:
$= (n / 1.5 \times 10^{-3}) \times 1000.47$
$= (4.823 \times 10^{-6} / 1.5 \times 10^{-3}) \times 1000.47$
$= 3.215 \times 10^{-3} \times 1000.47$
= 3.22 g

Answer: Mass of 1.5×10^{-3} m nalorphene solution required for the 1.5 mg dose = 3.22 g

Q30 Calculate the amount of benzoic acid (C_6H_5COOH) required for preparing 250 mL of 0.15 M solution in methanol.

SOLUTION

Given: Volume = 250 mL = 0.250 L; Molarity = 0.15 M

$M(C_6H_5COOH) = 7\times12 + 6\times1 + 2\times16 = 84 + 6 + 32 = 122\ g\ mol^{-1}$

Moles of benzoic acid required:

$n = Molarity \times Volume = 0.15 \times 0.250 = 0.0375\ mol$

Mass of benzoic acid:

$m = n \times M = 0.0375 \times 122 = 4.575\ g$

Answer: Mass of benzoic acid required = 4.575 g

Q31 The depression in freezing point of water observed for the same amount of acetic acid, trichloroacetic acid and trifluoroacetic acid increases in the order given above. Explain briefly.

SOLUTION

The three acids dissociate in water as weak acids:

$HA \rightleftharpoons H^+(aq) + A^-(aq)$

Depression in freezing point: $\Delta Tf = i \times Kf \times m$

For the same molality, ΔTf is larger when i is larger, i.e., when the degree of dissociation (α) is greater.

$i = 1 + \alpha$ for a monobasic acid.

Acid strengths and dissociation:

Acetic acid (CH_3COOH): $K_a \approx 1.8 \times 10^{-5}$ — weakly dissociated, smallest α.

Trichloroacetic acid (CCl_3COOH): The three Cl atoms are strongly electron-withdrawing (inductive effect), destabilising the O-H bond and stabilising CCl_3COO^-, making it more acidic than CH_3COOH. $K_a \approx 2 \times 10^{-1}$ — much more dissociated.

Trifluoroacetic acid (CF_3COOH): F is more electronegative than Cl. The three F atoms exert a stronger electron-withdrawing inductive effect than three Cl atoms, further stabilising CF_3COO^- and weakening the O-H bond. $K_a \approx 5.9 \times 10^{-1}$ — most dissociated, closest to strong acid.

Therefore, for the same mass of acid:

Degree of dissociation: $CH_3COOH < CCl_3COOH < CF_3COOH$
van't Hoff factor i: $CH_3COOH < CCl_3COOH < CF_3COOH$
ΔTf: $CH_3COOH < CCl_3COOH < CF_3COOH$

This matches the observed order.

Answer: ΔTf increases in the order $CH_3COOH < CCl_3COOH < CF_3COOH$ because the electron-withdrawing power of the halogen substituents increases (Cl < F), increasing acid strength and degree of dissociation (α), which increases the van't Hoff factor i and hence $\Delta Tf = i \cdot Kf \cdot m$.

Q32 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 = 1.86 K kg mol^{-1}.

SOLUTION

M($CH_3CH_2CHClCOOH$): $C_4H_7ClO_2 = 4\times12 + 7\times1 + 35.5 + 2\times16 = 48 + 7 + 35.5 + 32 = 122.5$ g mol^{-1}

Moles of acid = 10 / 122.5 = 0.08163 mol
Molality c = 0.08163 / (250/1000) = 0.3265 mol kg^{-1}

Since Ka = 1.4×10^{-3} and c = 0.3265 mol kg^{-1}, alpha << 1 (weak acid).
From $Ka = c \times \alpha^2 / (1 - \alpha) \approx c \times \alpha^2$:
$\alpha \approx \sqrt{(Ka/c)} = \sqrt{(1.4 \times 10^{-3} / 0.3265)} = \sqrt{(0.004288)} = 0.0655$

More exact (solving quadratic):
$c\alpha^2 + Ka\alpha - Ka = 0$
$\alpha = [-Ka + \sqrt{(Ka^2 + 4Kac)}] / (2c) = 0.0634$

van't Hoff factor: $i = 1 + \alpha = 1 + 0.0634 = 1.0634$

$\Delta Tf = i \times Kf \times c = 1.0634 \times 1.86 \times 0.3265 = 0.6459$ K

Answer: Depression in freezing point $\Delta Tf = 0.646$ K

Q33 19.5 g of CH_2FCOOH is dissolved in 500 g of water. The depression in the freezing point of water observed is 1.00°C. Calculate the van't Hoff factor and dissociation constant of fluoroacetic acid.

SOLUTION

$M(CH_2FCOOH) = 2\times12 + 3\times1 + 19 + 2\times16 = 24 + 3 + 19 + 32 = 78\ g\ mol^{-1}$

Moles of fluoroacetic acid = 19.5/78 = 0.250 mol
Molality = 0.250 / (500/1000) = 0.500 mol kg^{-1}

Step 1: ΔTf if no dissociation (ideal):
$\Delta T_f(ideal) = K_f \times m = 1.86 \times 0.500 = 0.930\ K$

Step 2: van't Hoff factor:
$i = \Delta T_f(observed) / \Delta T_f(ideal) = 1.00 / 0.930 = 1.075$

Step 3: Degree of dissociation:
For $CH_2FCOOH \rightleftharpoons CH_2FCOO^- + H^+$ (monobasic acid)
$i = 1 + \alpha \Rightarrow \alpha = i - 1 = 1.075 - 1 = 0.075$

Step 4: Dissociation constant:
$CH_2FCOOH \rightleftharpoons H^+ + CH_2FCOO^-$
At equilibrium: $[H^+] = [CH_2FCOO^-] = c\alpha = 0.500 \times 0.075 = 0.0375\ mol\ L^{-1}$ (using molality ≈ molarity)
$[CH_2FCOOH] = c(1-\alpha) = 0.500 \times 0.925 = 0.4625\ mol\ L^{-1}$

$K_a = [H^+][CH_2FCOO^-] / [CH_2FCOOH]$
$K_a = (0.500 \times 0.075^2) / (1 - 0.075)$
$K_a = (0.500 \times 0.005625) / 0.925$
$K_a = 0.0028125 / 0.925 = 3.04 \times 10^{-3}$

Answer: van't Hoff factor i = 1.075; Dissociation constant $K_a = 3.04 \times 10^{-3}$

Q34 Vapour pressure of water at 293 K is 17.535 mm Hg. Calculate the vapour pressure of water at 293 K when 25 g of glucose is dissolved in 450 g of water.

SOLUTION

Given: $p^*(water) = 17.535$ mm Hg at 293 K
Mass of glucose = 25 g, $M(glucose) = 180\ g\ mol^{-1}$
Mass of water = 450 g, $M(water) = 18\ g\ mol^{-1}$

Moles of glucose = 25/180 = 0.1389 mol
Moles of water = 450/18 = 25.00 mol

Mole fraction of water:
x(water) = 25.00 / (25.00 + 0.1389) = 25.00 / 25.139 = 0.9945

By Raoult's law:
p(water) = x(water) × p*(water) = 0.9945 × 17.535 = 17.44 mm Hg

Lowering of vapour pressure:
Δp = 17.535 - 17.44 = 0.096 mm Hg

Answer: Vapour pressure of water in the glucose solution = 17.44 mm Hg (lowering = 0.096 mm Hg)

Q35 Henry's law constant for the molality of methane in benzene at 298 K is 4.27 × 10^5 mm Hg. Calculate the solubility of methane in benzene at 298 K under 760 mm Hg.

SOLUTION

Given: $K_H = 4.27 \times 10^5$ mm Hg; Partial pressure of methane p = 760 mm Hg

By Henry's law: $p = K_H \times x$
where x is the mole fraction of methane in benzene.

$x(\text{methane}) = p / K_H = 760 / (4.27 \times 10^5)$
$x(\text{methane}) = 1.780 \times 10^{-3}$

This mole fraction represents the solubility of methane in benzene under 760 mm Hg.

Answer: Mole fraction (solubility) of methane in benzene = 1.78×10^{-3}

Q36 100 g of liquid A (molar mass 140 g mol^{-1}) was dissolved in 1000 g of liquid B (molar mass 180 g mol^{-1}). The vapour pressure of pure liquid B was found to be 500 torr. Calculate the vapour pressure of pure liquid A and its vapour pressure in the solution if the total vapour pressure of the solution is 475 Torr.

SOLUTION

Given: mass of A = 100 g, M_A = 140 g mol^{-1}; mass of B = 1000 g, M_B = 180 g mol^{-1}
p^*_B = 500 torr; p_{total} = 475 Torr

Moles of A = 100/140 = 0.7143 mol
Moles of B = 1000/180 = 5.556 mol

Total moles = 0.7143 + 5.556 = 6.270 mol

Mole fractions:
x_A = 0.7143/6.270 = 0.1139
x_B = 5.556/6.270 = 0.8861

Partial pressure of B in solution:
p_B = x_B × p*_B = 0.8861 × 500 = 443.0 torr

Partial pressure of A in solution:
p_A = p_total - p_B = 475 - 443.0 = 32.0 torr

Vapour pressure of pure A:
p_A = x_A × p*_A
p*_A = p_A / x_A = 32.0 / 0.1139 = 280.9 torr

Answer: Vapour pressure of pure liquid A = 280.9 torr; Vapour pressure of A in the solution = 32.0 torr

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Q37 Vapour pressures of pure acetone and chloroform at 328 K are 741.8 mm Hg and 632.8 mm Hg respectively. Assuming that they form ideal solution over the entire range of composition, plot p_total, p_chloroform, and p_acetone as a function of x_acetone. The experimental data observed for different compositions of mixture is given in the table. Plot this data also on the same graph paper. Indicate whether it has positive deviation or negative deviation from the ideal solution.

SOLUTION

Ideal solution (Raoult's law):
p_acetone(ideal) = x_acetone × 741.8 mm Hg
p_chloroform(ideal) = (1 - x_acetone) × 632.8 mm Hg
p_total(ideal) = p_acetone(ideal) + p_chloroform(ideal)

At x_acetone = 0: p_total = 632.8 mm Hg (pure chloroform)
At x_acetone = 1: p_total = 741.8 mm Hg (pure acetone)
These are the endpoints of the ideal straight-line (p_total vs x_acetone).

Experimental data verification (at x_acetone = 0.118):
Ideal p_acetone = 0.118 × 741.8 = 87.5 mm Hg; Experimental = 54.9 mm Hg
Ideal p_chloroform = 0.882 × 632.8 = 558.1 mm Hg; Experimental = 548.1 mm Hg
Ideal p_total = 645.6 mm Hg; Experimental = 54.9 + 548.1 = 603.0 mm Hg

Since experimental p_total < ideal p_total at all intermediate compositions, the

system shows NEGATIVE DEVIATION from Raoult's law.

Explanation: Chloroform ($CHCl_3$) can donate a hydrogen bond to the C=O group of acetone. This $CHCl_3$...$O=C(CH_3)_2$ interaction is STRONGER than the chloroform-chloroform or acetone-acetone interactions in the pure liquids. Molecules are held more strongly in solution, reducing their tendency to escape into the vapour phase, hence lower vapour pressure than ideal.

Delta_mix H < 0 (exothermic mixing), confirming negative deviation.

Answer: Negative deviation from Raoult's law. The experimental vapour pressures are LOWER than the ideal values because H-bonding ($CHCl_3$...$O=C(CH_3)_2$) between chloroform and acetone is stronger than interactions in the pure components.

Q38 Benzene and toluene form ideal solution over the entire range of composition. The vapour pressure of pure benzene and toluene at 300 K are 50.71 mm Hg and 32.06 mm Hg respectively. Calculate the mole fraction of benzene in vapour phase if 80 g of benzene is mixed with 100 g of toluene.

SOLUTION

Given: p*_benzene = 50.71 mm Hg, p*_toluene = 32.06 mm Hg at 300 K
M(benzene, C_6H_6) = 78 g mol^{-1}; M(toluene, C_7H_8) = 92 g mol^{-1}

Moles of benzene = 80/78 = 1.026 mol
Moles of toluene = 100/92 = 1.087 mol
Total moles = 1.026 + 1.087 = 2.113 mol

Mole fractions in liquid phase:
x_benzene = 1.026/2.113 = 0.4855
x_toluene = 1.087/2.113 = 0.5145

Partial pressures (ideal solution):
p_benzene = 0.4855 × 50.71 = 24.62 mm Hg
p_toluene = 0.5145 × 32.06 = 16.50 mm Hg
p_total = 24.62 + 16.50 = 41.12 mm Hg

Mole fraction of benzene in vapour phase:
y_benzene = p_benzene / p_total = 24.62 / 41.12 = 0.5988 ≈ 0.599

Answer: Mole fraction of benzene in vapour phase = 0.599

Q39 The air is a mixture of a number of gases. The major components are oxygen and nitrogen with approximate proportion of 20% to 79% by volume at 298 K. The water is in equilibrium with air at a pressure of 10 atm. At 298 K, if the Henry's law constants for oxygen and nitrogen at 298 K are 3.30×10^7 mm and 6.51×10^7 mm respectively, calculate the composition of these gases in water.

SOLUTION

Total pressure of air-water system = 10 atm.
Volume % ≈ mole % for ideal gases.

Partial pressures:
$p_{O_2} = 0.20 \times 10$ atm $= 2.0$ atm $= 2.0 \times 760 = 1520$ mm Hg
$p_{N_2} = 0.79 \times 10$ atm $= 7.9$ atm $= 7.9 \times 760 = 6004$ mm Hg

By Henry's law: $x = p / K_H$

Mole fraction of O_2 in water:
$x(O_2) = p_{O_2} / K_H(O_2) = 1520 / (3.30 \times 10^7) = 4.606 \times 10^{-5}$

Mole fraction of N_2 in water:
$x(N_2) = p_{N_2} / K_H(N_2) = 6004 / (6.51 \times 10^7) = 9.223 \times 10^{-5}$

Answer: Mole fraction of O_2 in water = 4.61×10^{-5}; Mole fraction of N_2 in water = 9.22×10^{-5}

Q40 Determine the amount of $CaCl_2$ (i = 2.47) dissolved in 2.5 litre of water such that its osmotic pressure is 0.75 atm at 27°C.

SOLUTION

Given: i = 2.47, V = 2.5 L, π = 0.75 atm, T = 27 + 273 = 300 K
$R = 0.0821$ L atm mol^{-1} K^{-1}
$M(CaCl_2) = 40 + 2 \times 35.5 = 111$ g mol^{-1}

By van't Hoff equation: $\pi = i \times C \times R \times T$
$C = \pi / (i \times R \times T) = 0.75 / (2.47 \times 0.0821 \times 300)$
$C = 0.75 / 60.87 = 0.01232$ mol L^{-1}

Moles of $CaCl_2$ in 2.5 L:
$n = C \times V = 0.01232 \times 2.5 = 0.03080$ mol

Mass of $CaCl_2$:
$m = n \times M = 0.03080 \times 111 = 3.42$ g

Answer: Mass of $CaCl_2$ required = 3.42 g

Q41 Determine the osmotic pressure of a solution prepared by dissolving 25 mg of K_2SO_4 in 2 litre of water at 25°C, assuming that it is completely dissociated.

SOLUTION

Given: mass of K_2SO_4 = 25 mg = 25×10^{-3} g; V = 2 L; T = 25 + 273 = 298 K
$M(K_2SO_4) = 2\times39 + 32 + 4\times16 = 78 + 32 + 64 = 174$ g mol^{-1}

K_2SO_4 dissociates completely: $K_2SO_4 \rightarrow 2K^+ + SO_4^{2-}$
3 particles per formula unit => i = 3

Moles of $K_2SO_4 = 25 \times 10^{-3} / 174 = 1.437 \times 10^{-4}$ mol
Concentration: $C = 1.437 \times 10^{-4} / 2 = 7.184 \times 10^{-5}$ mol L^{-1}

Osmotic pressure:
$\pi = i \times C \times R \times T$
$R = 0.08314$ L bar mol^{-1} K^{-1}
$\pi = 3 \times 7.184 \times 10^{-5} \times 0.08314 \times 298$
$\pi = 3 \times 7.184 \times 10^{-5} \times 24.78$
$\pi = 5.34 \times 10^{-3}$ bar

Answer: Osmotic pressure of the K_2SO_4 solution = 5.34×10^{-3} bar