Liquid Solution

Question1

Calculate the solubility of gas in solvent at $25^{\circ}\mathrm{C}$ and 0.8 atm if Henry's law constant for solvent is 6.8×10^{-4} moldm^{-3} atm^{-1}.

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Options:

A.

$5.88 \times 10^{-4} \mathrm{M}$

B.

$6.12 \times 10^{-4} \mathrm{M}$

C.

$5.44 \times 10^{-4} \mathrm{M}$

D.

$6.48 \times 10^{-4} \mathrm{M}$

Answer: C

Solution:

We are asked to calculate the solubility of a gas in a solvent using Henry's law:

$C = k_H \cdot P$

Where:

- C = concentration of dissolved gas (mol dm^{-3})
- k_H = Henry's law constant (mol dm^{-3} atm^{-1})
- P = partial pressure of the gas (atm)

Step 1: Write down given values

- $k_H = 6.8 \times 10^{-4}\,\mathrm{mol\,dm^{-3}\,atm^{-1}}$
- $P = 0.8\,\mathrm{atm}$

Step 2: Apply formula

$C = (6.8 \times 10^{-4})(0.8)$

$C = 5.44 \times 10^{-4}\,\mathrm{mol\,dm^{-3}}$

Step 3: Compare with options

The correct value is:

Option C: $5.44 \times 10^{-4}\mathrm{M}$ ✅

Final Answer:

$\boxed{5.44 \times 10^{-4}\,\mathrm{M}}$

Question2

Calculate osmotic pressure of 0.1 M aqueous solution of an electrolyte at 300 K if van't Hoff factor is 1.125.

$$\left[\mathrm{R} = 0.0821\ \mathrm{atmdm^3\ K^{-1}\ mol^{-1}}\right]$$

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Options:

A.

2.15 atm

B.

2.41 atm

C.

2.77 atm

D.

3.25 atm

Answer: C

Solution:

Step 1: Recall formula

$\pi = iCRT$

Where:

- π = osmotic pressure (atm)
- i = van't Hoff factor = 1.125
- C = molarity = 0.1 M = 0.1 mol/dm^3
- R = 0.0821 L·atm·K^{-1}·mol^{-1}
- T = 300 K

Step 2: Substitute values

$\pi = (1.125)(0.1)(0.0821)(300)$

First, multiply step by step:

$0.1 \times 0.0821 = 0.00821$

$0.00821 \times 300 = 2.463$

$2.463 \times 1.125 = 2.770$

Step 3: Final Answer

$\pi \approx 2.77$ atm

✅ Correct option: **Option C (2.77 atm)**

Question3

What type of solution is the H_2 in palladium?

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Options:

A.

Gas as solute and solid as solvent

B.

Gas as solute and liquid as solvent

C.

Gas as solvent and solid as solute

D.

Liquid as solute and gas as solvent

Answer: A

Solution:

- Palladium is a solid metal that can absorb a large volume of hydrogen gas.
- The hydrogen atoms diffuse into the interstitial spaces of the palladium lattice, forming what is called an **interstitial solid solution**.
- Here, palladium is the solvent (solid state), and hydrogen is the solute (gas initially, dissolved as atoms inside the metal lattice).

Correct Option:

Option A: Gas as solute and solid as solvent ✅

Final Answer: The solution of H_2 in palladium is a *gas in solid* solution (Interstitial solid solution).

Question4

Calculate cryoscopic constant of solvent if depression in freezing point of 0.18 m solution of nonvolatile solute is 0.2 K

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Options:

A.

$1.73 \text{ K kg mol}^{-1}$

B.

$1.50 \text{ K kg mol}^{-1}$

C.

$1.11 \text{ K kg mol}^{-1}$

D.

$1.92 \text{ K kg mol}^{-1}$

Answer: C

Solution:

We are asked to calculate the **cryoscopic constant** K_f of a solvent.

Formula:

$\Delta T_f = K_f m$

where

- ΔT_f = depression in freezing point
- K_f = cryoscopic constant
- m = molality

Given:

- $\Delta T_f = 0.20\,\text{K}$
- $m = 0.18\,\text{mol kg}^{-1}$

Calculation:

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

$K_f = \frac{0.20}{0.18}$

$K_f \approx 1.11\,\text{K kg mol}^{-1}$

Final Answer:

$\boxed{1.11\,\text{K kg mol}^{-1}}$

✅ Correct option: **Option C**

Question5

Which of the following aqueous solutions having same molality exhibits maximum boiling point elevation? (Assume complete dissociation)

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Options:

A.

KCl

B.

NaCl

C.

$AlCl_3$

D.

$BaCl_2$

Answer: C

Solution:

The solution having more number of particles will exhibits maximum boiling point elevation. Suppose the concentration of each substance is 1 m . Then,

$$\text{i) } \underset{1\text{ m}}{KCl} \longrightarrow \underset{1\text{ m}}{K^+} + \underset{1\text{ m}}{Cl^-}$$

Total particles in solution $= 2$ mol

$$\text{ii) } \underset{1\text{ m}}{NaCl} \longrightarrow \underset{1\text{ m}}{Na^+} + \underset{1\text{ m}}{Cl^-}$$

Total particles in solution $= 2$ mol

$$\text{iii) } \underset{1\text{ m}}{AlCl_3} \longrightarrow \underset{1\text{ m}}{Al^{3+}} + \underset{3\text{ m}}{3\,Cl^-}$$

Total particles in solution $= 4$ mol

$$\text{iv) } \underset{1\text{ m}}{BaCl2} \longrightarrow \underset{1\text{ m}}{Ba^{2+}} + \underset{2\text{ m}}{2Cl^-}$$

Total particles in solution $= 3$ mol

Hence, $AlCl_3$ solution gives more number of particles and exhibit maximum boiling point elevation.

Question6

Calculate the mole fraction of pure liquid B in solution if total vapour pressure of solution, vapour pressure of pure liquid A and vapour pressure of pure liquid B are 500 mm Hg , 400 mm Hg and 575 mm Hg respectively at given temperature.

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Options:

A.

0.43

B.

0.57

C.

0.62

D.

0.38

Answer: B

Solution:

Given:

- Total vapour pressure of solution, $P_{total} = 500 \, \text{mmHg}$
- Vapour pressure of pure liquid A, $P_A^0 = 400 \, \text{mmHg}$
- Vapour pressure of pure liquid B, $P_B^0 = 575 \, \text{mmHg}$

Let the mole fractions of A and B in liquid phase be:

x_A and x_B

with

$x_A + x_B = 1$

Raoult's Law:

$P_{total} = x_A P_A^0 + x_B P_B^0$

Substitute values:

$500 = (1 - x_B)(400) + x_B(575)$

Solve:

$500 = 400 - 400x_B + 575x_B$

$500 = 400 + 175x_B$

$100 = 175x_B$

$x_B = \frac{100}{175} = 0.571$

Final Answer:

The mole fraction of liquid **B** is:

$\boxed{0.57}$

✅ Correct Option: **B (0.57)**

Question7

What is the numerical value of osmotic pressure of 1 M urea solution if numerical value of osmotic pressure of 0.5 M urea solution is ' x '?

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Options:

A. x

B. $\frac{x}{2}$

C. $2x$

D. $3x$

Answer: C

Solution:

Osmotic pressure $\pi = \mathrm{i} \times \mathrm{M} \times \mathrm{R} \times \mathrm{T}$

For urea $i = 1$

As, $\pi \propto \mathrm{M}$

$$\therefore \quad \frac{\pi_{1\mathrm{M}}}{\pi_{0.5\mathrm{M}}} = \frac{\mathrm{M}_1}{\mathrm{M}_2} = \frac{1}{0.5}$$
$$\therefore \quad \text{For } 0.5\mathrm{M}, \pi_{0.5\mathrm{M}} = x$$
$$\therefore \quad \pi_{1\mathrm{M}} = \frac{1}{0.5} \times x = \frac{1}{0.5} \times x = 2x$$

Question8

Calculate vapour pressure of volatile liquid A at given temperature if mole fraction and vapour pressure of volatile liquid B are 0.4 and 900 mm Hg respectively [$\mathrm{P}_{\text{total}} = 600$ mmHg]

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Options:

A. 450 mm Hg

B. 560 mm Hg

C. 500 mm Hg

D. 400 mm Hg

Answer: D

Solution:

For the two volatile liquids A and B

Using Dalton's law of partial pressure,

$$\begin{aligned} P_{\text{total}} &= P_A + P_B \\ P_{\text{total}} &= P_A^0 x_1 + P_B^0 x_2 \end{aligned}$$

since mole fraction of B, $x_B = 0.4$

$\therefore$ mole fraction of A, $x_A = 1 - x_B$

$$x_A = 1 - 0.4$$
$$x_A = 0.6$$

$$\begin{aligned} P_{\text{total}} &= P_A^0 x_A + P_B^0 x_B \\ 600 &= P_A^0 \times 0.6 + 900 \times 0.4 \\ 600 - 360 &= 0.6 P_A^0 \\ 240 &= 0.6 P_A^0 \end{aligned}$$

$\therefore \quad P_A^0 = \frac{240}{0.6} = 400 \text{ mmHg}$

Question9

Calculate the molality of the solution of nonvolatile solute if it freezes at $-0.36^{\circ}\mathrm{C}$.

$\left[K_f \textbf{ for solvent} = 1.86 \mathrm{~K~kg~mol}^{-1}\right]$

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Options:

A.

$0.218 \mathrm{~mol~kg}^{-1}$

B.

$0.193 \mathrm{~mol~kg}^{-1}$

C.

$0.401 \mathrm{~mol~kg}^{-1}$

D.

$0.520 \text{ mol kg}^{-1}$

Answer: B

Solution:

Step 1. Recall formula:

$\Delta T_{\mathrm{f}} = K_f m$

where

- $\Delta T_{\mathrm{f}} = T_f^{\circ} - T_f$ (decrease in freezing point),
- K_f is the cryoscopic constant,
- m is the molality.

Step 2. Calculate the depression in freezing point:

The solvent (probably water) has normal freezing point $0^{\circ}\mathrm{C}$. New freezing point = $-0.36^{\circ}\mathrm{C}$.

$\Delta T_{\mathrm{f}} = 0 - (-0.36) = 0.36\,\mathrm{K}$

Step 3. Substitute into formula:

$m = \frac{\Delta T_{\mathrm{f}}}{K_f} = \frac{0.36}{1.86}$

Compute:

$m \approx 0.1935\,\text{mol kg}^{-1}$

Step 4. Select closest option:

That corresponds to **Option B:** $0.193\,\text{mol kg}^{-1}$.

✅ **Final Answer:**

$\boxed{0.193\,\text{mol kg}^{-1}\ \text{(Option B)}}$

Question10

Calculate the vapour pressure of solution if relative lowering of vapour pressure and vapour pressure of pure solvent are 0.018 and 18 mm Hg respectively at 300 K .

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Options:

A. 18.32 mm Hg

B. 17.08 mm Hg

C. 17.68 mm Hg

D. 18.60 mm Hg

Answer: C

Solution:

$$\begin{aligned}
& \frac{\Delta P}{P_1^0} = 0.018 \\
& \Delta P = 0.018 \times P_1^0 \\
& \Delta P = 0.018 \times 18 = 0.324 \mathrm{mmHg} \\
\because \quad & \Delta P = P_1^0 - P_1 \\
& 0.324 = 18 - P_1 \\
& P_1 = 18 - 0.324 = 17.68 \text{ mmHg}
\end{aligned}$$

Question11

Calculate the number of moles of nonvolatile solute dissolved in 0.5 kg solvent if molal elevation constant for solvent is
$2 \text{ kg K mol}^{-1} [\Delta T_b = 0.8 \text{ K}]$

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Options:

A. 0.1

B. 0.2

C. 0.3

D. 0.4

Answer: B

Solution:

Molal elevation constant relation:

$\Delta T_b = K_b \cdot m$

where

- $\Delta T_b = 0.8 \text{ K}$
- $K_b = 2 \text{ K kg mol}^{-1}$
- $m = \text{molality} = \dfrac{n}{\text{kg of solvent}}$
- solvent mass = 0.5 kg.

Step 1: Calculate molality

$m = \frac{\Delta T_b}{K_b} = \frac{0.8}{2} = 0.4 \text{ mol kg}^{-1}$

Step 2: Relating molality and moles of solute

$m = \frac{n}{\text{kg solvent}} \quad \Rightarrow \quad n = m \times \text{kg solvent}$

$n = 0.4 \times 0.5 = 0.2 \text{ mol}$

✅ Final Answer:

The number of moles of solute = **0.2 mol**

Correct Option: B (0.2)

Question12

Arrange the following equimolar solutions according to increasing order of osmotic pressure [Assume complete ionisation]

i) KCl

ii) $BaCl_2$

iii) $AlCl_3$

iv) $Al_2(SO_4)_3$

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Options:

A.

$BaCl_2 < Al_2(SO_4)_3 < KCl < AlCl_3$

B.

$Al_2(SO_4)_3 < KCl < BaCl_2 < AlCl_3$

C.

$KCl < BaCl_2 < AlCl_3 < Al_2(SO_4)_3$

D.

$AlCl_3 < BaCl_2 < Al_2(SO_4)_3 < KCl$

Answer: C

Solution:

Osmotic pressure is a colligative property that depends on number of particles in solution. The solution having more number of particles will have large osmotic pressure.

Suppose the concentration of each substance is 1 m . Then,

$$\text{i) } \underset{1\text{ m}}{KCl} \longrightarrow \underset{1\text{ m}}{K^{+}} + \underset{1\text{ m}}{Cl^{-}}$$

Total particles in solution $= 2$ mol

$$\text{ii) } \underset{1\text{ m}}{BaCl_2} \longrightarrow \underset{1\text{ m}}{Ba^{2+}} + \underset{2\text{ m}}{2Cl^{-}}$$

Total particles in solution $= 3$ mol

$$\text{iii) } \underset{1\text{ m}}{AlCl_3} \longrightarrow \underset{1\text{ m}}{Al^{3+}} + \underset{3\text{ m}}{3\,Cl^{-}}$$

Total particles in solution $= 4$ mol

$$\text{iv) } Al_2(SO_4)_3 \longrightarrow \underset{2\ m}{2Al^{3+}} + \underset{3\ m}{3SO_4^{2-}}$$
$$\underset{1\ m}{}$$

Total particles in solution $= 5$ mol

Hence, increasing order of osmotic pressure is $KCl < BaCl_2 < AlCl_3 < Al_2(SO_4)_3$

Question13

Which from following pairs of solutions in water exhibits same osmotic pressure at same temperature?

[molar mass of urea $= 60\ \text{g mol}^{-1}$, sucrose $= 342\ \text{g mol}^{-1}$]

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Options:

A. $3\ \text{g L}^{-1}$ urea and $17.1\ \text{g L}^{-1}$ sucrose

B. $6\ \text{g L}^{-1}$ urea and $17.1\ \text{g L}^{-1}$ sucrose

C. $3\ \text{g L}^{-1}$ urea and $34.2\ \text{g L}^{-1}$ sucrose

D. $6\ \text{g L}^{-1}$ urea and $8.6\ \text{g L}^{-1}$ sucrose

Answer: A

Solution:

$\pi = \frac{W_2}{M_2} \times \frac{1}{V} RT$ i.e. $\pi \propto \frac{W_2}{M_2}$

For same osmotic pressure, solutions must be isotonic.

$$\pi_{\text{area}} = \pi_{\text{sucrose}}$$
$$\frac{W_{\text{Urea}}}{M_{\text{Urea}}} = \frac{W_{\text{Suroxe}}}{M_{\text{Surorose}}}$$

For option (A)

$\frac{\mathrm{W_{Urea}}}{\mathrm{M_{Ura}}} = \frac{3}{60} = 0.05$ mol and $\frac{\mathrm{W_{Sucrose}}}{\mathrm{M_{Surose}}} = \frac{17.1}{342} = 0.05$ mol

$\therefore$ Solutions of $3\ \text{g L}^{-1}$ urea and $17.1\ \text{g L}^{-1}$ sucrose in water exhibit same osmotic pressure.

Question14

A solution of 5 g nonvolatile solute in 50 g water decreases its freezing point by 0.2 K . Calculate the molar mass of solute if $\mathrm{K_f}$ of water is $1.86 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$.

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Options:

A. $840 \mathrm{~g} \mathrm{~mol}^{-1}$

B. $930 \mathrm{~g} \mathrm{~mol}^{-1}$

C. $960 \mathrm{~g} \mathrm{~mol}^{-1}$

D. $870 \mathrm{~g} \mathrm{~mol}^{-1}$

Answer: B

Solution:

$$
\begin{aligned}
& \Delta \mathrm{T_f}=0.2 \mathrm{~K} \\
& \quad \mathrm{M_2}=\frac{\mathrm{K_f} \times \mathrm{W_2} \times 1000}{\Delta \mathrm{~T_f} \times \mathrm{W_1}}=\frac{1.86 \times 5 \times 1000}{0.2 \times 50} \\
& \qquad=930 \mathrm{~g} \mathrm{~mol}^{-1} \\
& \therefore \quad \text{Molar mass of solute }=930 \mathrm{~g} \mathrm{~mol}^{-1}
\end{aligned}
$$

Question15

Calculate the molality of the solution containing nonvolatile solute if boiling point elevation of solution is 0.39 K .

[$\mathrm{K_b}$ of water $=0.52 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$]

Options:

A. 0.52 mol kg^{-1}

B. 0.65 mol kg^{-1}

C. 0.75 mol kg^{-1}

D. 0.86 mol kg^{-1}

Answer: C

Solution:

Formula:

$\Delta T_b = K_b \cdot m$

Where:

- ΔT_b = elevation in boiling point
- K_b = molal boiling point constant
- m = molality of the solution

Step 1: Substitute known values

$\Delta T_b = 0.39 \, \mathrm{K}, \quad K_b = 0.52 \, \mathrm{K \, kg \, mol}^{-1}$

$0.39 = 0.52 \times m$

Step 2: Solve for molality

$m = \frac{0.39}{0.52}$

$m \approx 0.75 \, \mathrm{mol \, kg}^{-1}$

✅ Correct Answer: **Option C: 0.75 mol kg^{-1}**

Question16

Calculate the cryoscopic constant of solvent if depression in freezing point of 0.4 m solution of nonvolatile solute is 1.8 K .

Options:

A. $4.0 \text{ K kg mol}^{-1}$

B. $4.5 \text{ K kg mol}^{-1}$

C. $5.1 \text{ K kg mol}^{-1}$

D. $5.7 \text{ K kg mol}^{-1}$

Answer: B

Solution:

Formula:

$\Delta T_f = K_f \cdot m$

where

- ΔT_f = depression in freezing point
- K_f = cryoscopic constant
- m = molality of solute

Given:

$\Delta T_f = 1.8\,\text{K}$

$m = 0.4\,\text{mol kg}^{-1}$

Step 1: Rearrange the formula

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

Step 2: Substitute values

$K_f = \frac{1.8}{0.4} = 4.5\,\text{K·kg·mol}^{-1}$

✅ **Final Answer:**

$\boxed{4.5\,\text{K·kg·mol}^{-1}}$

Correct option: **B**

Question17

Which from following mixtures obeys Raoult's law?

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Options:

A. Phenol and aniline

B. Chloroform and acetone

C. Ethanol and acetone

D. Benzene and toluene

Answer: D

Solution:

Raoult's law is obeyed by **ideal solutions**, i.e. mixtures where intermolecular interactions between unlike molecules (A–B) are similar to those between like molecules (A–A and B–B).

Non-ideal behavior arises when strong hydrogen bonding or specific interactions change the vapor pressure significantly.

- **Phenol and aniline** → Both form strong intermolecular hydrogen bonds. Thus, show *negative deviation* from Raoult's law. Not ideal.
- **Chloroform and acetone** → They form strong hydrogen-bond-like interactions (chloroform H with acetone O). Large *negative deviation*. Not ideal.
- **Ethanol and acetone** → Interaction between hydrogen-bond donor (ethanol) and acceptor (acetone) alters escaping tendency of molecules. Deviation from Raoult's law. Not ideal.
- **Benzene and toluene** → Both are non-polar hydrocarbons of similar molecular size and intermolecular forces (weak van der Waals). This pair behaves nearly ideally.

✅ **Correct Answer: Option D – Benzene and toluene**

Question18

Calculate the osmotic pressure of 0.2 M aqueous solution of electrolyte at 300 K . If van't Hoff factor is

$$1.6\left[\mathrm{R}=0.0821\ \mathrm{atmdmK}^{-1}\ \mathrm{mol}^{-1}\right].$$

Options:

A. 7.21 atm

B. 7.88 atm

C. 8.81 atm

D. 8.32 atm

Answer: B

Solution:

Formula:

$\pi = iMRT$

Where:

- i = van't Hoff factor
- M = molarity of electrolyte solution = 0.2 M
- $R = 0.0821\,\mathrm{L{\cdot}atm{\cdot}K^{-1}{\cdot}mol^{-1}}$
- $T = 300\,\mathrm{K}$

Step 1: Insert values

$\pi = (1.6)(0.2)(0.0821)(300)$

Step 2: Simplify

$0.2 \times 300 = 60$

$1.6 \times 60 = 96$

$96 \times 0.0821 \approx 7.8816$

Final Answer:

$\pi \approx 7.88\,\mathrm{atm}$

✅ **Correct Option: B (7.88 atm)**

Question19

Determine the expected value of ΔT_f for $1 \mathrm{~mCaCl}_2$ solution if 1 m urea solution has $\Delta \mathrm{T}_{\mathrm{f}}$ value ' x ' K .

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Options:

A. x K

B. $2x$ K

C. $3x$ K

D. $\frac{x}{2}$ K

Answer: C

Solution:

Step 1. Recall freezing-point depression formula

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

where:

- i = van 't Hoff factor (number of particles produced per formula unit of solute),
- K_f = cryoscopic constant (depends only on solvent),
- m = molality.

Step 2. Reference system (urea)

- Urea is a **non-electrolyte** (does not dissociate).

$i = 1.$

For $m = 1$:

$\Delta T_f(\text{urea}) = 1 \cdot K_f \cdot 1 = K_f$

But we are told this value is x.

Therefore:

$K_f = x.$

Step 3. Case of CaCl_2

- Calcium chloride dissociates (ideally) as:

$\mathrm{CaCl}_2 \rightarrow \mathrm{Ca}^{2+} + 2\,\mathrm{Cl}^-$

So total number of ions per unit = 3.

Thus, $i = 3$.

Step 4. Freezing-point depression for $CaCl_2$

For $m = 1$:

$\Delta T_f(\mathrm{CaCl}_2) = i \cdot K_f \cdot m = 3 \cdot x \cdot 1 = 3x.$

✅ **Final Answer:**

$\Delta T_f = 3x\,\mathrm{K}$

Correct Option: C $3x$ K

Question20

Calculate van't Hoff factor of aqueous solution of 0.18 m electrolyte that freezes at $-0.54^{\circ}\mathrm{C}$.

$\left(\mathrm{K_f}\ \textbf{for solvent} = 1.86\ \mathrm{K\ kg\ mol}^{-1}\right)$

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Options:

A. 1.126

B. 2.449

C. 1.612

D. 2.150

Answer: C

Solution:

Step 1. Recall the freezing point depression formula

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

where

- ΔT_f = depression in freezing point
- K_f = cryoscopic constant of solvent
- m = molality of solution
- i = van't Hoff factor

Step 2. Insert known values

- Experimental freezing point depression:

$\Delta T_f = 0.54\,\mathrm{K}$

(since solution freezes at $-0.54°C$, while pure solvent freezes at $0°C$)

- $K_f = 1.86\,\mathrm{K \cdot kg \cdot mol^{-1}}$
- $m = 0.18\,\mathrm{mol \cdot kg^{-1}}$

Step 3. Solve for i

$i = \frac{\Delta T_f}{K_f \cdot m}$

$i = \frac{0.54}{1.86 \times 0.18}$

First compute denominator:

$1.86 \times 0.18 = 0.3348$

Now compute:

$i = \frac{0.54}{0.3348} \approx 1.612$

Final Answer:

The van't Hoff factor is

$\boxed{1.612}$

✅ Correct Option: **C (1.612)**

Question21

Calculate the molality of solution of nonvolatile solute if boiling point elevation is 1.75 K .

[$\mathrm{K_b}$ for solvent $= 3\ \mathrm{K\ kg\ mol^{-1}}$]

Options:

A. 0.480 m

B. 0.33 m

C. 0.58 m

D. 0.63 m

Answer: C

Solution:

We are asked to calculate the **molality** of a solution given:

- Boiling point elevation, $\Delta T_b = 1.75\,\mathrm{K}$
- K_b of solvent = 3 K·kg·mol^{-1}

Step 1: Recall formula

$\Delta T_b = K_b \times m$

where

- ΔT_b = elevation in boiling point
- K_b = ebullioscopic constant
- m = molality

Step 2: Solve for molality

$m = \frac{\Delta T_b}{K_b} = \frac{1.75}{3}$

$m = 0.583\,\mathrm{mol/kg}$

Step 3: Match nearest option

That is approximately **0.58 m**.

✅ **Correct Answer: Option C – 0.58 m**

Question22

Calculate the relative lowering of vapour pressure of solution containing 3 g urea in 50 g water. [molar mass of urea $= 60 \text{ g mol}^{-1}$]

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Options:

A. 0.013

B. 0.025

C. 0.018

D. 0.028

Answer: C

Solution:

We are tasked with calculating the **relative lowering of vapour pressure (RLVP):**

$$\frac{\Delta p}{p^0} = \frac{n_{\text{solute}}}{n_{\text{solvent}}}$$

Step 1: Moles of solute (urea)

$$\text{Mass} = 3 \text{ g}, \quad M = 60 \text{ g mol}^{-1}$$

$$n_{\text{solute}} = \frac{3}{60} = 0.05 \text{ mol}$$

Step 2: Moles of solvent (water)

$$\text{Mass} = 50 \text{ g}, \quad M = 18 \text{ g mol}^{-1}$$

$$n_{\text{solvent}} = \frac{50}{18} \approx 2.78 \text{ mol}$$

Step 3: Calculate RLVP

$$\frac{\Delta p}{p^0} = \frac{n_{\text{solute}}}{n_{\text{solvent}}} = \frac{0.05}{2.78} \approx 0.018$$

✅ **Final Answer:**

Option C: 0.018

Question23

Which from following solutions exhibits minimum boiling point elevation under identical conditions? (Assume complete dissociation)

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Options:

A. 0.2 m KCl

B. 0.1 m NaCl

C. 1 $mAlCl_3$

D. 0.05 $mMgCl_2$

Answer: D

Solution:

Step 1: Recall the formula

Boiling point elevation:

$\Delta T_b = i\,K_b\,m$

where:

- m = molality
- i = van't Hoff factor (number of particles produced per formula unit)
- K_b = ebullioscopic constant (same for all since solvent is same)

So, under identical solvent conditions:

$\Delta T_b \propto i \times m$

Step 2: Calculate i for each solute (assuming complete dissociation)

- **KCl** $\rightarrow K^+ + Cl^- \rightarrow i = 2$
- **NaCl** $\rightarrow Na^+ + Cl^- \rightarrow i = 2$
- **$AlCl_3$** $\rightarrow Al^{3+} + 3Cl^- \rightarrow i = 4$
- **$MgCl_2$** $\rightarrow Mg^{2+} + 2Cl^- \rightarrow i = 3$

Step 3: Compute effective molality $i \times m$

- **Option A (0.2 m KCl):** $im = 2 \times 0.2 = 0.4$

- **Option B (0.1 m NaCl):** $im = 2 \times 0.1 = 0.2$
- **Option C (1.0 m $AlCl_3$):** $im = 4 \times 1 = 4.0$
- **Option D (0.05 m $MgCl_2$):** $im = 3 \times 0.05 = 0.15$

Step 4: Identify minimum

Smallest product $im = 0.15$ (for $MgCl_2$).

✅ **Answer: Option D (0.05 m $MgCl_2$)** exhibits minimum boiling point elevation.

Question24

When 0.01 mole of nonvolatile solute is dissolved in certain solvent calculate the mass of solvent in kg if $\Delta \mathrm{T_b} = 0.6\ \mathrm{K}$ and $\mathrm{K_b}$ for solvent= $2\ \mathrm{K\ kg\ mol}^{-1}$

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Options:

A. 0.014 kg

B. 0.028 kg

C. 0.033 kg

D. 0.045 kg

Answer: C

Solution:

- Moles of solute = $n = 0.01$ mol
- Elevation in boiling point = $\Delta T_b = 0.6$ K
- Molal elevation constant = $K_b = 2\ \mathrm{K\ kg\ mol}^{-1}$
- Mass of solvent = ? (in kg)

Step 1: Recall the formula for elevation of boiling point

$\Delta T_b = K_b m$

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

Step 2: Express molality

$m = \frac{0.01}{W}$

where W = mass of solvent in kg.

Step 3: Substitute into the formula

$0.6 = 2 \times \frac{0.01}{W}$

$0.6 = \frac{0.02}{W}$

Step 4: Solve for W

$W = \frac{0.02}{0.6} = 0.033 \text{ kg}$

✔ Correct Answer:

Option C: 0.033 kg

Question25

Which from following mixtures exhibits positive deviation from Raoult's law?

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Options:

A. Ethanol and acetone

B. Benzene and toluene

C. Chloroform and acetone

D. Phenol and aniline

Answer: A

Solution:

The solutions in which solute-solvent intermolecular attractions are weaker than those between solute-solute molecules and solvent solvent molecules, exhibit positive deviations. The solution of ethanol and acetone shows positive deviation from the Raoult's law.

Question26

Calculate the osmotic pressure of 0.5 M aqueous solution of nonvolatile solute at 300 K .

$$\left[\mathrm{R}=0.0821 \text{ atmdm } \mathrm{K}^{-1} \mathrm{~mol}^{-1}\right]$$

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Options:

A. 9.51 atm

B. 12.32 atm

C. 15.60 atm

D. 6.75 atm

Answer: B

Solution:

Formula for Osmotic Pressure:

$\pi = MRT$

Where:

- M = molarity = 0.5 mol/L
- R = gas constant = 0.0821 L·atm·K^{-1}·mol^{-1}
- T = temperature in Kelvin = 300 K

Substitution:

$\pi = (0.5) \times (0.0821) \times (300)$

First calculate step by step:

$0.0821 \times 300 = 24.63$

$\pi = 0.5 \times 24.63 = 12.315 \text{ atm}$

Final Answer:

$\pi \approx 12.32\,\text{atm}$

✅ Correct Option: **B. 12.32 atm**

Question27

Calculate the number of moles of nonvolatile solute dissolved in 0.3 kg solvent if $\Delta \mathrm{T_b} = 0.3\ \mathrm{K}$ and $\mathrm{K_b}$ for solvent is $1.8\ \mathrm{K\ kg\ mol^{-1}}$.

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Options:

A. 0.051

B. 0.044

C. 0.062

D. 0.073

Answer: A

Solution:

We are given:

$\Delta T_b = K_b \cdot m$

where m is the molality of the solution.

Step 1: Rearrange for molality

$m = \frac{\Delta T_b}{K_b}$

$m = \frac{0.3}{1.8} = 0.1667\ \text{mol/kg}$

Step 2: Relate molality to moles of solute

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

where n = moles of solute.

The solvent mass is given: $0.3\,\text{kg}$.

$n = m \times \text{mass of solvent}$

$n = 0.1667 \times 0.3 = 0.0500\,\text{mol}$

Step 3: Compare with options

- Approx. value = **0.051 mol**

👉 Correct Answer: **Option A (0.051)**

Question28

Calculate the boiling point elevation of solution if 15 g urea is dissolved in 1000 g water. $[\mathrm{K_b}$ for water $= 0.52\ \mathrm{K\ kg\ mol^{-1}}$; molar mass of urea $= 60\ \mathrm{g\ mol^{-1}}]$

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Options:

A. 0.13 K

B. 0.24 K

C. 0.38 K

D. 0.54 K

Answer: A

Solution:

We are asked to calculate the **boiling point elevation** (ΔT_b).

Step 1: Recall the formula

$\Delta T_b = K_b \times m$

where

- K_b = ebullioscopic constant = $0.52\,\text{K kg mol}^{-1}$
- m = molality of solution

Step 2: Calculate moles of solute (urea)

Moles of urea $= \frac{\text{Mass of solute}}{\text{Molar mass}} = \frac{15}{60} = 0.25\,\text{mol}$

Step 3: Calculate mass of solvent in kg

Mass of solvent (water) $= 1000\,\text{g} = 1.0\,\text{kg}$

Step 4: Calculate molality

$m = \frac{0.25\,\text{mol}}{1.0\,\text{kg}} = 0.25\,\text{mol/kg}$

Step 5: Apply elevation formula

$\Delta T_b = 0.52 \times 0.25 = 0.13\,\text{K}$

✅ **Final Answer: Option A → 0.13 K**

Question29

Arrange the following solutions according to decreasing order of osmotic pressure under similar condition of temperature and assuming complete dissociation.

I.	0.2 mKCl
II.	$0.3\ \text{mMgSO}_4$
III.	$0.1\ \text{mBaCl}_2$
IV.	$0.5\ \text{mAl}_2(\text{SO}_4)_3$

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Options:

A. IV > II > I > III

B. III > IV > I > II

C. I > III > IV > II

D. II > III > I > IV

Answer: A

Solution:

Osmotic pressure is a colligative property that depends on number of particles in solution. The solution having more number of particles will have larger osmotic pressure.

Solution	Dissociation	Molality of ions after dissociation
0.2 mKCl	$KCl \longrightarrow K^{+} + Cl^{-}$	$2 \times 0.2 = 0.4$ m
0.3 m $MgSO_4$	$MgSO_4 \xrightarrow{Mg^{2+}} +SO_4^{2-}$	$2 \times 0.3 = 0.6$ m
0.1 m $BaCl_2$	$BaCl_2 \xrightarrow[Ba^{2+}]{} +2Cl^{-}$	$3 \times 0.1 = 0.3$ m
0.5 m $Al_2(SO_4)_3$	$Al_2(SO_4)_3 \longrightarrow 2Al^{3+} + 3SO_4^{2-}$	$5 \times 0.5 = 2.5$ m

Hence, $Al_2(SO_4)_3$ solution gives more number of particles and has the highest osmotic pressure among the given. So, the correct order of increasing osmotic pressure is

IV > II > I > III.

Question30

Calculate the relative lowering of vapour pressure of solution containing 0.56 g nonvolatile solute in 100 g water [molar mass of solute $= 60 \text{ g mol}^{-1}$]

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Options:

A. 0.0024

B. 0.0120

C. 0.0017

D. 0.0221

Answer: C

Solution:

We are asked to calculate the **relative lowering of vapour pressure (RLVP)**.

Formula:

$$\frac{\Delta p}{p^0} = \frac{n_{\text{solute}}}{n_{\text{solvent}}}$$

Step 1: Calculate moles of solute

$$n_{\text{solute}} = \frac{\text{mass of solute}}{\text{molar mass of solute}}$$

$$n_{\text{solute}} = \frac{0.56}{60} = 0.00933\,\text{mol}$$

Step 2: Calculate moles of solvent (water)

Molar mass of water $= 18\,\text{g/mol}$.

Mass of solvent = $100\,\text{g}$.

$$n_{\text{solvent}} = \frac{100}{18} \approx 5.556\,\text{mol}$$

Step 3: Relative lowering of vapour pressure

$$\frac{\Delta p}{p^0} = \frac{n_{\text{solute}}}{n_{\text{solvent}}} = \frac{0.00933}{5.556}$$

$$\frac{\Delta p}{p^0} \approx 0.00168 \approx 0.0017$$

✅ Final Answer:

Option C — 0.0017

Question31

Which of the following aqueous solutions exhibits lowest freezing point depression assuming complete dissociation?

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Options:

A. 0.1 m NaCl

B. 0.05 m$MgSO_4$

C. $0.08 \mathrm{~mAlPO}_4$

D. $0.06 \mathrm{~mAl}_2(\mathrm{SO}_4)_3$

Answer: B

Solution:

Depression of freezing point is colligative property which depends on number of particles present in solution. Lesser is the number of particles, less will be the depression.

Solution	Dissociation	Molality of ions after dissociation
$0.1 \mathrm{~mNaCl}$	$\mathrm{NaCl} \rightarrow$ $\mathrm{Na}^{+}+\mathrm{Cl}^{-}$	$0.1 \times 2=0.2 \mathrm{~m}$
$0.05 \mathrm{~mMgSO}_4$	$\mathrm{MgSO}_4 \rightarrow$ $\mathrm{Mg}^{2+}+\mathrm{SO}_4^{2-}$	$0.05 \times 2=0.1 \mathrm{~m}$
$0.08 \mathrm{~mAlPO}_4$	$\mathrm{AlPO}_4 \rightarrow$ $\mathrm{Al}^{3+}+\mathrm{PO}_4^{3-}$	$0.08 \times 2=0.16 \mathrm{~m}$
$0.06 \mathrm{~mAl}_2(\mathrm{SO}_4)_3$	$\mathrm{Al}_2(\mathrm{SO}_4)_3 \rightarrow$ $2 \mathrm{Al}^{3+}+3 \mathrm{SO}_4^{2-}$	$0.06 \times 5=0.3 \mathrm{~m}$

Question32

Calculate the temperature of 0.05 M sucrose solution in Kelvin if the osmotic pressure of the solution is 1.5 atm .

$$\left[\mathrm{R}=0.0821 \mathrm{dm}^3 \text { atm } \mathrm{K}^{-1} \mathrm{~mol}^{-1}\right]$$

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Options:

A. 370.2

B. 375.4

C. 380.6

D. 365.4

Answer: D

Solution:

Given:

Osmotic pressure, $\Pi = 1.5$ atm

Molarity, $C = 0.05 \text{ mol L}^{-1}$

Gas constant, $R = 0.0821 \text{ dm}^3 \text{ atm K}^{-1} \text{ mol}^{-1}$

According to the formula for osmotic pressure:

$\Pi = CRT$

We have to calculate temperature T in Kelvin.

So,

$T = \frac{\Pi}{C \cdot R}$

Substitute the given values:

$T = \frac{1.5}{0.05 \times 0.0821}$

Calculate the denominator:

$0.05 \times 0.0821 = 0.004105$

Now,

$T = \frac{1.5}{0.004105}$

$T \approx 365.4 \text{ K}$

Correct Option:

Option D: 365.4

Question33

Calculate molality of the solution containing nonvolatile solute if boiling point elevation of the solution is 0.2 K

$$\left[\mathrm{K_b} = 0.52 \text{ K kg mol}^{-1}\right]$$

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Options:

A. $0.162 \text{ mol kg}^{-1}$

B. $0.281 \text{ mol kg}^{-1}$

C. $0.385 \text{ mol kg}^{-1}$

D. $0.501 \text{ mol kg}^{-1}$

Answer: C

Solution:

Given data:

- Elevation in boiling point, $\Delta T_b = 0.2 \text{ K}$
- Molal elevation constant, $K_b = 0.52 \text{ K kg mol}^{-1}$

According to NCERT:

The elevation in boiling point is calculated using the formula:

$\Delta T_b = K_b \cdot m$

where m is the molality of the solution.

Step 1: Substitute the values

$0.2 = 0.52 \times m$

Step 2: Solve for m (molality)

$m = \frac{0.2}{0.52}$

Step 3: Calculate

$m = \frac{0.2}{0.52} = 0.3846 \text{ mol kg}^{-1}$

Rounding off to three significant figures,

$m \approx 0.385 \text{ mol kg}^{-1}$

Correct option is:

Option C: $0.385 \text{ mol kg}^{-1}$

Question34

Find the expected value of $\Delta \mathrm{T_b}$ for $1 \mathrm{~mAlCl_3}$ solution in water if solution of nonelectrolyte of same concentration has $\Delta \mathrm{T_b}$ value ' x ' K .

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Options:

A. x K

B. $2x$ K

C. $3x$ K

D. $4x$ K

Answer: D

Solution:

Given:

- ΔT_b for 1 m solution of a nonelectrolyte is x K.
- For 1 m $AlCl_3$ solution in water, find the expected value of ΔT_b.

Let us proceed step by step:

Step 1: Boiling Point Elevation Formula

For any solute,

$\Delta T_b = i \cdot K_b \cdot m$

- i is the van't Hoff factor.
- For a nonelectrolyte, $i = 1$.

Given, for nonelectrolyte:

$\Delta T_{b,\text{ nonelec}} = 1 \cdot K_b \cdot 1 = K_b$ (since $m = 1$)

But it is given that $\Delta T_{b,\text{ nonelec}} = x$ K, so $K_b = x$.

Step 2: Find i for $AlCl_3$

$AlCl_3$ dissociates in water as follows:

$AlCl_3 \rightarrow Al^{3+} + 3\,Cl^-$

So, **total particles produced** $= 1 + 3 = 4$

Therefore, $i = 4$

Step 3: Calculate ΔT_b for $AlCl_3$

$\Delta T_{b,\,AlCl_3} = i \cdot K_b \cdot m = 4 \cdot x \cdot 1 = 4x$

Step 4: Final Answer

So, the expected value is $4x$ K.

Correct option:

$\boxed{4x \text{ K}}$

(Option D)

Question35

Calculate vapour pressure of pure volatile liquid B at given temperature if mole faction of liquid B and vapour pressure of pure volatile liquid A are 0.4 and 400 mm Hg respectively.

$\left[\mathrm{P}_{\text{total}} \quad = 600 \text{ mmHg}\right]$

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Options:

A. 750 mm Hg

B. 800 mm Hg

C. 850 mm Hg

D. 900 mm Hg

Answer: D

Solution:

There are two volatile liquids A and B.

Using Dalton's law of partial pressure,

$\therefore \quad P_{\text{Total}} = P_A + P_B$

Applying Raoult's law,

$\therefore \quad P_{\text{Total}} = x_A P_A^0 + x_B P_B^0$

$\therefore \quad x_A = 1 - x_B = 1 - 0.4$

$\therefore \quad x_A = 0.6$

$P_A^0 = 400 \mathrm{mmHg}$

$P_{\text{Total}} = 600 \mathrm{mmHg}$

Now, $P_{\text{Total}} = 0.6 \times 400 + 0.4 \times P_B^0$

$\therefore \quad 600 = 0.6 \times 400 + 0.4 \times P_B^0$

$600 - (0.6 \times 400) = 0.4 \times P_B^0$

$\therefore \quad P_B^0 = \frac{600-(0.6 \times 400)}{0.4} = \frac{600-240}{0.4} = \frac{360}{0.4}$

$= 900 \mathrm{mmHg}$

$P_B^0 =$ vapour pressure of pure $B = 900 \mathrm{mmHg}$

Question36

Calculate the molality of nonvolatile solution if solution freezes at $-0.95^{\circ} \mathrm{C}$

$[K_f$ for water, $= 1.86 \mathrm{~K} \mathrm{~mol}^{-1}$, freezing point of water $= 0^{\circ} \mathrm{C}$]

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Options:

A. $0.032 \mathrm{~mol} \mathrm{~kg}^{-1}$

B. $0.041 \mathrm{~mol} \mathrm{~kg}^{-1}$

C. $0.051 \mathrm{~mol} \mathrm{~kg}^{-1}$

D. $0.065 \mathrm{~mol} \mathrm{~kg}^{-1}$

Answer: C

Solution:

Step 1: Determine the freezing point depression

The freezing point depression, ΔT_f, is the difference between the freezing point of the pure solvent ($0^\circ\mathrm{C}$ for water) and the freezing point of the solution ($-0.95^\circ\mathrm{C}$).

$$\Delta T_f = T_{f,\mathrm{solvent}} - T_{f,\mathrm{solution}}$$

$$\Delta T_f = 0^\circ\mathrm{C} - (-0.95^\circ\mathrm{C}) = 0.95^\circ\mathrm{C}$$

Step 2: Use the freezing point depression formula to find molality

The formula relating freezing point depression to molality (m) is $\Delta T_f = i \cdot K_f \cdot m$. For a nonvolatile solution (assuming a non-electrolyte solute), the van't Hoff factor (i) is 1. The cryoscopic constant (K_f) for water is $1.86\,\mathrm{K\,mol^{-1}}$. Since a change of $1^\circ\mathrm{C}$ is equal to a change of $1\,\mathrm{K}$, ΔT_f can be used directly in Kelvin.

$$\Delta T_f = K_f \cdot m$$

$$0.95\,\mathrm{K} = 1.86\,\mathrm{K\,mol^{-1}} \cdot m$$

Step 3: Calculate the molality

Rearrange the equation to solve for m:

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

$$m = \frac{0.95\,\mathrm{K}}{1.86\,\mathrm{K\,mol^{-1}}}$$

$$m \approx 0.051\,\mathrm{mol\,kg^{-1}}$$

Answer:

(C) $\mathbf{0.051\,mol\,kg^{-1}}$

Question37

Which of the following solutions exhibits highest freezing point depression?

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Options:

A. 0.1 m NaCl

B. 0.05 m$MgSO_4$

C. 1 m$AlPO_4$

D. 0.05 m$Al_2(SO_4)_3$

Answer: C

Solution:

$NaCl \longrightarrow Na^+ + Cl^-$ Total no. of particles $= 0.1 + 0.1 = 0.2$ m
$MgSO_4 \longrightarrow Mg^{2+} + SO_4^{2-}$ Total no. of particles $= 0.05 + 0.05 = 0.1$ m
$AlPO_4 \longrightarrow Al^{3+} + PO_4^{3-}$ Total no. of particles $= 1 + 1 = 2$ m
$Al_2(SO_4)_3 \longrightarrow 2Al^{3+} + 3SO_4^{2-}$ Total no. of particles $= 0.10 + 0.15 = 0.25$ m

1 m$AlPO_4$ solution has maximum no. of particles, hence it shows highest freezing point depression.

Question38

Calculate the concentration of dissolved gas in water at $25^\circ C$ if partial pressure of gas at same temperature is 0.15 atm .

$$\left[K_H = 0.15 \text{ moldm}^{-3} \text{ atm}^{-1}\right]$$

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Options:

A. 0.0261 M

B. 0.0182 M

C. 0.0293 M

D. 0.0225 M

Answer: D

Solution:

According to **Henry's Law**:

$C = K_H \cdot P$

where:

- C = concentration of dissolved gas (mol dm^{-3})
- K_H = Henry's law constant (mol dm^{-3} atm^{-1})
- P = partial pressure of the gas (atm)

Given:

- $K_H = 0.15 \text{ mol dm}^{-3} \text{ atm}^{-1}$
- $P = 0.15 \text{ atm}$

Step 1: Substitute values into the formula

$C = 0.15 \text{ mol dm}^{-3} \text{ atm}^{-1} \times 0.15 \text{ atm}$

Step 2: Calculate

$C = 0.0225 \text{ mol dm}^{-3}$

Step 3: Write the answer using suitable units and match options

The concentration of the dissolved gas is 0.0225 M.

Correct option:

Option D: 0.0225 M

Question39

Calculate the osmotic pressure of 0.03 mole of non electrolyte solute dissolved in 0.1dm^3 of water at $300 \text{ K}. \left[\text{R} = 0.082 \text{dm}^3 \text{ atm mol}^{-1} \text{ K}^{-1}\right]$

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Options:

A. 7.4 atm

B. 6.4 atm

C. 8.0 atm

D. 5.6 atm

Answer: A

Solution:

Given:

Number of moles of solute, $n = 0.03$ mole

Volume of solution, $V = 0.1\ \text{dm}^3$

Temperature, $T = 300$ K

Gas constant, $R = 0.082\ \text{dm}^3\ \text{atm mol}^{-1}\ \text{K}^{-1}$

According to NCERT, the formula for osmotic pressure (π) for a dilute solution is:

$\pi = CRT$

where,

$C = \text{molarity of the solution} = \dfrac{n}{V}$

Step 1: Calculate molarity (C)

$C = \frac{n}{V} = \frac{0.03}{0.1} = 0.3\ \text{mol dm}^{-3}$

Step 2: Substitute in the formula for osmotic pressure:

$\pi = CRT$

$\pi = 0.3 \times 0.082 \times 300$

Step 3: Calculate the value

$0.3 \times 0.082 = 0.0246$

$0.0246 \times 300 = 7.38$

So,

$\pi \approx 7.4$ atm

Correct answer: Option A (7.4 atm)

Question40

Identify from following salts so that the solubility of salt in water decreases with increase in temperature.

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Options:

A. NaBr

B. NaCl

C. $NaNO_3$

D. Na_2SO_4

Answer: D

Solution:

Solubility of most salts increases with increase in temperature. However, **some salts show decreased solubility with rise in temperature** if their dissolution is an exothermic process.

Let us consider each option:

1. **NaBr** (Sodium bromide)

- Solubility increases with temperature (endothermic dissolution).

1. **NaCl** (Sodium chloride)

- Solubility almost remains constant or increases slightly with temperature.

1. **$NaNO_3$** (Sodium nitrate)

- Solubility increases significantly with temperature.

1. **Na_2SO_4** (Sodium sulphate)

- Solubility **decreases** with increase in temperature after a certain point (above about $32.4°C$), because its dissolution is exothermic.

Correct Answer: Option D

$\boxed{Na_2SO_4}$

Question41

Calculate the molal elevation constant of solvent if boiling point of 0.12 m solution is 319.8 K (Boling point of solvent $= 319.5 \mathrm{~K}$)

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Options:

A. $2.0 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

B. $3.0 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

C. $2.5 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

D. $3.5 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

Answer: C

Solution:

Given:

- Boiling point of solution, $T_b = 319.8 \mathrm{~K}$
- Boiling point of solvent, $T_b^0 = 319.5 \mathrm{~K}$
- Molality of solution, $m = 0.12 \mathrm{~mol} \mathrm{~kg}^{-1}$

Step 1: Calculate elevation in boiling point

$\Delta T_b = T_b - T_b^0 = 319.8 \mathrm{~K} - 319.5 \mathrm{~K} = 0.3 \mathrm{~K}$

Step 2: Use the formula for elevation in boiling point

$\Delta T_b = K_b \times m$

Where K_b is the molal elevation constant (boiling point elevation constant).

Step 3: Substitute the values and solve for K_b

$0.3 = K_b \times 0.12$

$K_b = \frac{0.3}{0.12}$

$K_b = 2.5 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

Final Answer:

$\boxed{2.5 \text{ K kg mol}^{-1}}$

So the correct option is **C**.

Question42

Calculate the concentration of an aqueous solution of non electrolyte at 300 K if its osmotic pressure is 12 atm .

$$\left[\mathrm{R}=0.0821 \text{ atmdm}^{3} \mathrm{~K}^{-1} \mathrm{~mol}^{-1}\right]$$

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Options:

A. 0.371 M

B. 0.615 M

C. 0.487 M

D. 0.726 M

Answer: C

Solution:

Given:

- Osmotic pressure, $\pi = 12$ atm
- Temperature, $T = 300$ K
- $R = 0.0821$ atm·dm³·K⁻¹·mol⁻¹

Let concentration C (in mol/L or M) = ?

Step 1: Write the formula for osmotic pressure.

According to NCERT, for a non-electrolyte:

$\pi = CRT$

Step 2: Substitute the given values.

$12 = C \times 0.0821 \times 300$

Step 3: Rearrange the equation to solve for C.

$C = \frac{12}{0.0821 \times 300}$

Step 4: Calculate the denominator.

$0.0821 \times 300 = 24.63$

Step 5: Substitute and simplify.

$C = \frac{12}{24.63}$

$C = 0.487 \text{ mol/L (M)}$

Final Answer:

Option C — 0.487 M

Question43

Which from following compounds is least soluble in water at STP?

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Options:

A. C_2H_5OH

B. CH_3OH

C. CH_3NH_2

D. CH_4

Answer: D

Solution:

The $N - H$ bonds in amines are less polar than $O - H$ bond in alcohols. As a result, alcohol forms stronger H -bonds with water as compared to amines. Hence, alcohols are more soluble in water than amines, while alkanes are the least polar. Therefore, CH_4 is least soluble in water at STP.

Question44

Calculate the percent dissociation of 0.02 m solution if its freezing point depression is 0.046 K .

$\left[\mathrm{K_f}\ \textbf{for water} = 1.86\ \mathrm{K\ kg\ mol}^{-1}; \mathrm{n} = 2\right]$

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Options:

A. 12.3%

B. 23.6%

C. 35.00%

D. 48.1%

Answer: B

Solution:

Given data:

- Molality, $m = 0.02\ \mathrm{m}$
- Depression in freezing point, $\Delta T_f = 0.046\ \mathrm{K}$
- Cryoscopic constant, $K_f = 1.86\ \mathrm{K\,kg\,mol}^{-1}$
- Number of ions after dissociation, $n = 2$

Step 1: Write formula for depression in freezing point

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

where i is the van't Hoff factor.

Step 2: Calculate van't Hoff factor i

$i = \frac{\Delta T_f}{K_f \cdot m}$

Substitute values:

$i = \frac{0.046}{1.86 \times 0.02} = \frac{0.046}{0.0372} \approx 1.24$

Step 3: Relate i to degree of dissociation (α)

For a salt AB which dissociates as $AB \rightarrow A^+ + B^-$ $(n = 2$ ions):

$i = 1 + (n - 1)\alpha$

Here, $n = 2$:

$i = 1 + (2 - 1)\alpha = 1 + \alpha$

$\alpha = i - 1$

$\alpha = 1.24 - 1 = 0.24$

Step 4: Find percent dissociation

$\text{Percent dissociation} = \alpha \times 100 = 0.24 \times 100 = 24\%$

The closest option is:

$\boxed{23.6\%}$

Option B is correct.

Question45

Which from following mixtures obeys Raoult's law?

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Options:

A. Chloroform + acetone

B. Carbon disulfide + acetone

C. Benzene + toluene

D. Ethanol + acetone

Answer: C

Solution:

To determine which mixture obeys Raoult's law, recall:

- **Raoult's law** is obeyed by mixtures that are *ideal solutions*.
- *Ideal solutions* are formed when the intermolecular forces between like molecules (A-A and B-B) are nearly the same as those between unlike molecules (A-B).
- In ideal solutions, there is no change in enthalpy ($\Delta H_{mix} = 0$) and no change in volume ($\Delta V_{mix} = 0$) on mixing.

Let us examine each option:

(A) Chloroform + Acetone:

- These have strong hydrogen bonding between chloroform (C-H…O) and acetone.
- Intermolecular forces between unlike molecules are much stronger than between like molecules.
- Shows *deviation* from Raoult's law (specifically, negative deviation).

(B) Carbon disulfide + Acetone:

- Carbon disulfide (CS_2) is non-polar; acetone is polar.
- Forces between unlike molecules differ from those between like molecules.
- Shows *deviation* from Raoult's law.

(C) Benzene + Toluene:

- Both are non-polar and structurally similar (aromatic hydrocarbons).
- Intermolecular forces between like and unlike molecules are very similar.
- They form *ideal solutions* and *obey* Raoult's law.

(D) Ethanol + Acetone:

- Ethanol has hydrogen bonding; acetone is a polar aprotic solvent.
- Interactions between unlike molecules are different from like molecules.
- Shows *deviation* from Raoult's law.

Final Answer:

Option C: Benzene + Toluene obeys Raoult's law.

Question46

Find molar mass of nonvolatile solute when 20 g of it dissolved in 200 g water at 300 K . [Relative lowering of vapour pressure $= 0.02$]

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Options:

A. 120 g mol^{-1}

B. 110 g mol^{-1}

C. 90 g mol^{-1}

D. 100 g mol^{-1}

Answer: C

Solution:

Given:

Mass of solute, $w_2 = 20$ g

Mass of solvent (water), $w_1 = 200$ g

Relative lowering of vapour pressure, $\dfrac{\Delta p}{p^0} = 0.02$

Molar mass of solvent (water), $M_1 = 18$ g mol^{-1}

Let molar mass of solute be M_2 (to be found).

Step 1: Write formula for relative lowering of vapour pressure

According to Raoult's law (NCERT formula),

Relative lowering of vapour pressure $= \frac{\Delta p}{p^0} = \frac{n_2}{n_1+n_2}$

For dilute solutions, $n_2 \ll n_1$, so

$\frac{\Delta p}{p^0} \approx \frac{n_2}{n_1}$

Where n_2 = moles of solute, n_1 = moles of solvent.

Step 2: Write moles in terms of mass and molar mass

$n_2 = \frac{w_2}{M_2} \qquad n_1 = \frac{w_1}{M_1}$

$\Rightarrow \frac{\Delta p}{p^0} = \dfrac{\dfrac{w_2}{M_2}}{\dfrac{w_1}{M_1}}$

$= \frac{w_2}{M_2} \cdot \frac{M_1}{w_1}$

$= \frac{w_2 M_1}{w_1 M_2}$

Step 3: Substitute values and solve for M_2

$0.02 = \frac{20 \times 18}{200 \times M_2}$

$0.02 = \frac{360}{200 \times M_2}$

$0.02 \times 200 \times M_2 = 360$

$4M_2 = 360$

$M_2 = \frac{360}{4} = 90$ g mol^{-1}

Final answer:

Option C

$\boxed{90 \mathrm{~g} \mathrm{~mol}^{-1}}$

Question47

What type of solution is the iodine in air?

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Options:

A. Liquid in solid

B. Solid in gas

C. Solid in liquid

D. Liquid in gas

Answer: B

Solution:

The type of solution for iodine in air is:

Option B: Solid in gas

Iodine, when dispersed in air, forms a solution where small particles of solid iodine are distributed within the gaseous medium. In chemistry, this is known as a "solid in gas" solution.

Question48

Calculate the relative lowering of vapour pressure if vapour pressure of pure solvent and vapour pressure of solution at $25^{\circ} \mathrm{C}$ are 32 and 30 mm Hg respectively.

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Options:

A. 0.0721

B. 0.0552

C. 0.0625

D. 0.9375

Answer: C

Solution:

Relative lowering of vapour pressure

$= \frac{P^{\circ}-P}{P^{\circ}} = \frac{32-30}{32} = 0.0625 \text{ Hg}$

Question49

Calculate the molality of solution if its depression in freezing point is 0.18 K . $\left[K_f = 1.6 \text{ K kg mol}^{-1}\right]$

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Options:

A. 0.195 m

B. 0.174 m

C. 0.156 m

D. 0.113 m

Answer: D

Solution:

To calculate the molality of a solution with a given depression in freezing point, the formula relevant to freezing point depression is:

$\Delta T_f = K_f \cdot m$

where:

ΔT_f is the depression in freezing point,

K_f is the cryoscopic constant (freezing point depression constant),

m is the molality of the solution.

Given:

$\Delta T_f = 0.18$ K,

$K_f = 1.6$ K kg mol^{-1}.

We need to solve for m:

$$m = \frac{\Delta T_f}{K_f} = \frac{0.18 \text{ K}}{1.6 \text{ K kg mol}^{-1}} = 0.1125 \text{ mol kg}^{-1}$$

Since the calculated molality (0.1125 m) is closest to 0.113 m:

Option D: 0.113 m is the correct answer.

Question50

Identify the correct statement from following properties.

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Options:

A. The osmosis is a colligative property.

B. The vapour pressure of solution containing nonvolatile solute is lower than that of pure solvent at any given temperature.

C. The osmotic pressure of 0.1 M NaCl solution is lower than 0.1 M sucrose solution.

D. The boiling point of solution containing nonvolatile solute is lower than that of pure solvent.

Answer: B

Solution:

Option A: "The osmosis is a colligative property."

Strictly speaking, the *colligative property* in this context is **osmotic pressure**, not the general phenomenon of osmosis itself. Osmotic pressure depends on the number of solute particles (i.e., it is colligative), but saying “osmosis is a colligative property” is not the standard statement. Usually, we say “osmotic pressure” is a colligative property.

Option B: “The vapour pressure of solution containing nonvolatile solute is lower than that of pure solvent at any given temperature.”

This is a **true** statement. Nonvolatile solutes reduce the mole fraction of the solvent in the solution, thereby lowering its vapor pressure compared to the pure solvent (Raoult’s law).

Option C: “The osmotic pressure of 0.1 M NaCl solution is lower than 0.1 M sucrose solution.”

This is **false** because NaCl dissociates into ions (Na^+ and Cl^-). For a 0.1 M NaCl solution, the effective concentration of particles is higher than 0.1 M (taking the van ‘t Hoff factor $i \approx 2$), so its osmotic pressure is actually **higher** than that of a 0.1 M sucrose solution (which does not dissociate).

Option D: “The boiling point of solution containing nonvolatile solute is lower than that of pure solvent.”

This is **false** because the presence of a nonvolatile solute **increases** the boiling point (boiling point elevation).

Correct Statement

From the above analysis, the **only correct statement** is:

Option B: “The vapour pressure of solution containing nonvolatile solute is lower than that of pure solvent at any given temperature.”

Question51

Calculate van't Hoff factor of 0.15 M solution of electrolyte if it freezes at -0.5 K.

$$\left[\mathrm{K_f} = 1.86\ \mathrm{K\ kg\ mol}^{-1}\right]$$

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Options:

A. 1.12

B. 1.25

C. 1.45

D. 1.79

Answer: D

Solution:

$$\Delta T_f = T_f^0 - T_f = 0 - (-0.5\ \mathrm{K}) = 0.5\ \mathrm{K}$$
$$\Delta\ \mathrm{T_f} = \mathrm{iK_f\ m}$$
$$\therefore\quad \mathrm{i} = \frac{\Delta \mathrm{T}}{\mathrm{K_f\ m}} = \frac{0.5\ \mathrm{K}}{1.86\ \mathrm{K\ kg\ mol^{-1}} \times 0.15\ \mathrm{mol\ kg^{-1}}} = 1.79$$

Question52

Calculate the relative lowering of vapour pressure of solution containing 46 g of non volatile solute in 162 g of water at $20^{\circ}\mathrm{C}$. [Molar mass of nonvolatile solute $= 46\ \mathrm{g\ mol^{-1}}$]

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Options:

A. 0.89

B. 0.045

C. 0.11

D. 0.06

Answer: C

Solution:

The relative lowering of vapor pressure is given by Raoult's Law for solutions. It can be calculated using the formula:

$$\frac{\Delta P}{P_0} = \frac{n_{\text{solute}}}{n_{\text{solvent}} + n_{\text{solute}}}$$

where:

ΔP is the decrease in vapor pressure,

P_0 is the pure solvent vapor pressure,

n_{solute} and n_{solvent} are the number of moles of solute and solvent, respectively.

For solutions with only a small amount of solute compared to the solvent, we can approximate the above relation to:

$$\frac{\Delta P}{P_0} \approx \frac{n_{\text{solute}}}{n_{\text{solvent}}}$$

To calculate the moles of solute, we use the formula:

$$n_{\text{solute}} = \frac{\text{mass of solute}}{\text{molar mass of solute}}$$

Given:

Mass of solute = 46 g,

Molar mass of solute = 46 g/mol,

Thus:

$$n_{\text{solute}} = \frac{46 \text{ g}}{46 \text{ g/mol}} = 1 \text{ mol}$$

Similarly, for the solvent (water):

Given:

Mass of water = 162 g,

Molar mass of water = 18 g/mol (approximately for H_2O),

Thus:

$$n_{\text{solvent}} = \frac{162 \text{ g}}{18 \text{ g/mol}} = 9 \text{ mol}$$

Now, calculate the relative lowering of vapor pressure:

$$\frac{\Delta P}{P_0} = \frac{n_{\text{solute}}}{n_{\text{solvent}}} = \frac{1}{9} = 0.11$$

Therefore, the correct option is:

Option C

0.11

Question53

Calculate molar mass of a solute at 300 K if 400 mg of it is dissolved in 300 mL of water exerts osmotic pressure of 0.2 atm .

$$\left(\mathrm{R} = 0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1}\right)$$

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Options:

A. $90 \mathrm{~g} \mathrm{~mol}^{-1}$

B. $120 \mathrm{~g} \mathrm{~mol}^{-1}$

C. $164 \mathrm{~g} \mathrm{~mol}^{-1}$

D. $180 \mathrm{~g} \mathrm{~mol}^{-1}$

Answer: C

Solution:

$$\begin{aligned} \mathrm{M}_2 & =\frac{\mathrm{W}_2 \mathrm{RT}}{\pi \mathrm{~V}} \\ & =\frac{0.4 \mathrm{~g} \times 0.0821 \mathrm{dm}^3 \mathrm{atmK}^{-1} \mathrm{~mol}^{-1} \times 300 \mathrm{~K}}{0.2 \mathrm{~atm} \times 0.3 \mathrm{dm}^3} \\ & =164.2 \mathrm{~g} \mathrm{~mol}^{-1} \approx 164 \mathrm{~g} \mathrm{~mol}^{-1} \end{aligned}$$

Question54

What is the vapour pressure of a solution containing 0.1 mol of non volatile solute dissolved in 16.2 g water? $\left(\mathrm{P}_1^0=24 \mathrm{mmHg}\right.$, molar mass of water $18 \mathrm{~g} \mathrm{~mol}^{-1}$)

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Options:

A. 12.4 mm Hg

B. 15.7 mm Hg

C. 18.1 mm Hg

D. 21.6 mm Hg

Answer: D

Solution:

$$\frac{P_1^0 - P_1}{P_1^0} = \frac{n_2 M_1}{W_1}$$
$$\therefore \quad \frac{24 - P_1}{24} = \frac{0.1 \times 18}{16.2}$$
$$\therefore \quad P_1 = 21.33 \text{ mmHg} \approx 21.6 \text{ mmHg}$$

Question55

Carbonated water is an example of a solution of

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Options:

A. Gas in liquid

B. Gas in solid

C. Gas in gas

D. Solid in gas

Answer: A

Solution:

Option A : Gas in liquid

Carbonated water is an example of a solution where a gas (carbon dioxide, CO_2) is dissolved in a liquid (water, H_2O). This is a common type of solution known as a liquid-gas solution. When CO_2 is dissolved in water, it can form carbonic acid, which is responsible for the characteristic fizz and slight acidity of carbonated water. The solubility of carbon dioxide in water increases with higher pressure and decreases with higher temperature, which is why carbonated beverages are often stored under pressure and served cold to maintain carbonation.

Question56

What is the expected value of ΔT_f for $1.25 \text{ m} CaCl_2$ solution if 1.25 m sucrose solution has ΔT_f value x K ?

Options:

A. xK

B. $\frac{x}{2}$K

C. $2 \times$ K

D. $3x$ K

Answer: D

Solution:

The depression in freezing point, ΔT_f, is determined by the formula:

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

where:

i is the van't Hoff factor, which represents the number of particles the solute dissociates into in solution.

K_f is the molal freezing point depression constant.

m is the molality of the solution.

For 1.25 m sucrose solution:

Sucrose does not dissociate into ions. Hence, the van't Hoff factor i for sucrose is 1.

Therefore, $\Delta T_f = x$ K for this solution.

For 1.25 m $CaCl_2$ solution:

$CaCl_2$ dissociates into 3 ions: 1 Ca^{2+} and 2 Cl^- ions. Hence, the van't Hoff factor i for $CaCl_2$ is 3.

Therefore, the ΔT_f for the $CaCl_2$ solution is given by:

$\Delta T_f = 3 \times K_f \times 1.25 = 3x$ K

So, the expected value of ΔT_f for the $1.25\,\mathrm{m}\,CaCl_2$ solution is $3x$ K.

The correct option is **Option D:** $3x$ K.

Question57

Calculate Henry's law constant if solubility of gas in water at $25^{\circ}\mathrm{C}$ is 5.14×10^{-4} moldm^{-3} and partial pressure of the gas is 0.75 bar above solution.

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Options:

A. 6.85×10^{-4} $\mathrm{moldm}^{-3}\mathrm{bar}^{-1}$

B. $5.14 \times 10^{-4}\mathrm{moldm}^{-3}\mathrm{bar}^{-1}$

C. 1.028×10^{-4} $\mathrm{moldm}^{-3}\mathrm{bar}^{-1}$

D. 1.371×10^{-4} $\mathrm{moldm}^{-3}\mathrm{bar}^{-1}$

Answer: A

Solution:

According to Henry's Law, the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas above the liquid. The law is mathematically expressed as:

$C = k_H \cdot P$

where:

C is the solubility of the gas in the liquid (in molarity, mol/dm^3),

k_H is Henry's law constant (in mol/dm^3/bar),

P is the partial pressure of the gas (in bar).

Given:

Solubility, $C = 5.14 \times 10^{-4}\,\mathrm{mol/dm}^3$,

Partial pressure, $P = 0.75\,\mathrm{bar}$.

We can rearrange the formula to solve for the Henry's law constant:

$k_H = \frac{C}{P}$

Substitute the known values:

$k_H = \frac{5.14 \times 10^{-4}\,\mathrm{mol/dm}^3}{0.75\,\mathrm{bar}}$

Calculate:

$k_H = \frac{5.14 \times 10^{-4}}{0.75} = 6.85 \times 10^{-4}\,\mathrm{mol/dm}^3/\mathrm{bar}$

Thus, the correct option is:

Option A: 6.85×10^{-4} mol/dm^3/bar.

Question58

A solution of nonvolatile solute is obtained by dissolving 0.8 g in $0.3 \mathrm{dm}^3$ water has osmotic pressure 0.2 atm at 300 K . Calculate the molar mass of solute.

$$\left[\mathrm{R}=0.082 \text{ atmdm}^3 \mathrm{~K}^{-1} \mathrm{~mol}^{-1}\right]$$

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Options:

A. $300 \mathrm{~g} \mathrm{~mol}^{-1}$

B. $340 \mathrm{~g} \mathrm{~mol}^{-1}$

C. $328 \mathrm{~g} \mathrm{~mol}^{-1}$

D. $352 \mathrm{~g} \mathrm{~mol}^{-1}$

Answer: C

Solution:

$$\begin{aligned} \mathrm{M}_2 & =\frac{\mathrm{W}_2 \mathrm{RT}}{\pi \mathrm{V}} \\ & =\frac{0.8 \mathrm{~g} \times 0.082 \mathrm{dm}^3 \text{ atm } \mathrm{K}^{-1} \mathrm{~mol}^{-1} \times 300 \mathrm{~K}}{0.2 \text{ atm} \times 0.3 \mathrm{dm}^3} \\ & =328 \mathrm{~g} \mathrm{~mol}^{-1}\end{aligned}$$

Question59

In a solution, mole fraction of solute is 0.2 , when lowering in vapour pressure is 10 mm Hg . To get lowering of vapour pressure of 20 mm

Hg , mole fraction of solute in solution is

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Options:

A. 0.2

B. 0.4

C. 0.6

D. 0.8

Answer: B

Solution:

The lowering of vapor pressure in a solution is directly proportional to the mole fraction of the solute, according to Raoult's Law. Mathematically, the relation is given by:

$\Delta P = P_{\text{solvent}} \times X_{\text{solute}}$

where:

ΔP is the lowering of vapor pressure,

P_{solvent} is the vapor pressure of the pure solvent,

X_{solute} is the mole fraction of the solute.

Given that when the lowering in vapor pressure is 10 mm Hg, the mole fraction of the solute is 0.2. We can express this as:

$10 = P_{\text{solvent}} \times 0.2$

For the new condition where the lowering of vapor pressure is 20 mm Hg, we want to find the new mole fraction X'_{solute} :

$20 = P_{\text{solvent}} \times X'_{\text{solute}}$

Dividing the second equation by the first gives:

$\frac{20}{10} = \frac{P_{\text{solvent}} \times X'_{\text{solute}}}{P_{\text{solvent}} \times 0.2}$

Simplifying, we get:

$2 = \frac{X'_{\text{solute}}}{0.2}$

Therefore:

$X'_{\text{solute}} = 2 \times 0.2 = 0.4$

Thus, the mole fraction of the solute in the solution needed to achieve a lowering of vapor pressure of 20 mm Hg is 0.4. This corresponds to **Option B: 0.4**.

Question60

Which of the following solutions will not show flow of solvent in either direction when separated by semipermeable membrane?

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Options:

A. 6 g urea dm^{-3} and 85.5 g sucrose dm^{-3}

B. 15 g urea dm^{-3} and 34.2 g sucrose dm^{-3}

C. 6 g urea dm^{-3} and 34.2 g sucrose dm^{-3}

D. 15 g urea dm^{-3} and 171 g sucrose dm^{-3}

Answer: C

Solution:

$\mathrm{n}_{\text{ura}} = \frac{6}{60} = 0.1 \text{ mol and } \mathrm{n}_{\text{secrose}} = \frac{34.2}{342} = 0.1 \text{ mol}$

Now, $\mathrm{n}_{\text{urea}} = \mathrm{n}_{\text{sucrose}}$

$\therefore \quad \pi_{\text{urea}} = \pi_{\text{sucrose}}$

Hence, if these solutions are separated by a semipermeable membrane, there is no flow of solvent in either direction.

Question61

The solution containing $18 \mathrm{~gdm}^{-3}$ glucose (molar mass 180) in water and another containing $6 \mathrm{~gdm}^{-3}$ of solute A in water boils at same temperature. What is molar mass of A ?

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Options:

A. 54 $\mathrm{g\ mol^{-1}}$

B. 90 $\mathrm{g\ mol^{-1}}$

C. 120 $\mathrm{g\ mol^{-1}}$

D. 60 $\mathrm{g\ mol^{-1}}$

Answer: D

Solution:

The boiling point elevation is a colligative property, which means it depends on the number of solute particles in a solution rather than the identity of the particles. Since both the solutions boil at the same temperature, their boiling point elevations are equal. Therefore, we can equate their molality.

For glucose:

The concentration is given as $18\,\mathrm{g\,dm^{-3}}$, and the molar mass is $180\,\mathrm{g\,mol^{-1}}$.

Convert this into molarity:

Molarity of glucose $= \frac{18\text{ g dm}^{-3}}{180\text{ g mol}^{-1}} = 0.1\,\mathrm{mol\,dm^{-3}}$

For solute A:

The concentration is given as $6\,\mathrm{g\,dm^{-3}}$, and we need to find its molar mass M.

Convert this into molarity:

Molarity of solute A $= \frac{6\text{ g dm}^{-3}}{M\text{ g mol}^{-1}}$

Since the boiling point elevations are the same, the molarity of the two solutions must be equal:

$0.1 = \frac{6}{M}$

Solving for M:

$M = \frac{6}{0.1} = 60\,\mathrm{g\,mol^{-1}}$

Thus, the molar mass of solute A is $60\,\mathrm{g\,mol^{-1}}$.

The correct answer is option D: $60\,\mathrm{g\,mol^{-1}}$.

Question62

If P_1 partial pressure of a gas and x_1 is its mole fraction in a mixture, then correct relation between P_1 and x_1 is

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Options:

A. $\mathrm{P}_{\text{total}} = \mathrm{P}_1 \mathrm{x}_1$

B. $\mathrm{x}_1 = \frac{\mathrm{P}_1}{\mathrm{P}_{\text{total}}}$

C. $\mathrm{P}_{\text{total}} = 1 - \mathrm{P}_1 \mathrm{x}_1$

D. $P_{\text{total}} = P_1 (1 - x_1)$

Answer: B

Solution:

The correct relation between the partial pressure of a gas (P_1) and its mole fraction (x_1) in a mixture is given by Dalton's Law of Partial Pressures. According to this law, the partial pressure of a gas in a mixture is proportional to its mole fraction and the total pressure of the mixture. The expression can be written as:

$P_1 = x_1 \cdot P_{\text{total}}$

From this relationship, we can derive the mole fraction as:

$x_1 = \frac{P_1}{P_{\text{total}}}$

This corresponds to Option B:

$x_1 = \frac{P_1}{P_{\text{total}}}$

Question63

Calculate van't Hoff factor (i) of 0.2 m aqueous solution of an electrolyte if it freezes at -0.660 K. $\left(\mathrm{K_f} = 1.84\ \mathrm{K\ kg\ mol^{-1}}\right)$

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Options:

A. 1.97

B. 1.79

C. 0.242

D. 0.557

Answer: B

Solution:

To calculate the van't Hoff factor (i), we can use the formula for freezing point depression:

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

where:

ΔT_f is the depression in the freezing point,

i is the van't Hoff factor,

K_f is the freezing point depression constant,

m is the molality of the solution.

From the problem, we know:

$\Delta T_f = 0.660 \, \mathrm{K}$,

$K_f = 1.84 \, \mathrm{K\,kg\,mol^{-1}}$,

$m = 0.2 \, \mathrm{mol\,kg^{-1}}$.

We can rearrange the formula to solve for i:

$i = \frac{\Delta T_f}{K_f \cdot m}$

Substituting the given values:

$i = \frac{0.660}{1.84 \times 0.2}$

Calculate:

$i = \frac{0.660}{0.368} \approx 1.79$

The van't Hoff factor for this solution is approximately 1.79. Therefore, the correct option is **Option B: 1.79**.

Question64

Vapour pressure of a pure solvent is 550 mm of Hg . By addition of a non volatile solute it decreases to 510 mm of Hg . Calculate the mole fraction of solute in solution.

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Options:

A. 0.215

B. 0.072

C. 0.610

D. 0.512

Answer: B

Solution:

To calculate the mole fraction of the solute in the solution, we need to use Raoult's Law, which relates the vapor pressure of a solvent in a solution to its mole fraction and the vapor pressure of the pure solvent.

According to Raoult's Law, the vapor pressure of the solvent in a solution (P_{solution}) is given by:

$$P_{\text{solution}} = X_{\text{solvent}} \cdot P_{\text{pure solvent}}$$

where

P_{solution} is the vapor pressure of the solvent above the solution (510 mm of Hg),

X_{solvent} is the mole fraction of the solvent,

$P_{\text{pure solvent}}$ is the vapor pressure of the pure solvent (550 mm of Hg).

The mole fraction of the solvent (X_{solvent}) can be found using:

$$X_{\text{solvent}} = \frac{P_{\text{solution}}}{P_{\text{pure solvent}}}$$

Substitute the given values:

$$X_{\text{solvent}} = \frac{510}{550} = \frac{51}{55}$$

Now calculate:

$X_{\text{solvent}} \quad = \frac{51}{55} \approx 0.927$

Since the sum of the mole fractions of the solute and the solvent must equal 1, the mole fraction of the solute (X_{solute}) is:

$X_{\text{solute}} \quad = 1 - X_{\text{solvent}}$

Thus:

$X_{\text{solute}} \quad = 1 - 0.927 = 0.073$

Thus, the mole fraction of the solute in the solution is approximately 0.072, which corresponds to option B.

Question65

Identify the false statement among the following.

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Options:

A. Ideal solution obey Raoult's law over entire range of temperature and concentration.

B. For ideal solution $\Delta_{\text{mix}} \quad \mathrm{V} = 0$.

C. Non ideal solution do not obey Raoult's law over the entire range of concentration.

D. Vapour pressure of non ideal solution always lies between vapour pressure of pure components.

Answer: D

Solution:

Option D is the false statement.

In non-ideal solutions, the vapor pressure does not always lie between the vapor pressures of the pure components. Non-ideal solutions can have vapor pressures higher or lower than those predicted by Raoult's law, depending on the nature of the interactions between the molecules of the solute and solvent. This deviation occurs due to positive deviation (where interaction between different molecules is weaker than between similar molecules) or negative deviation (where interaction between different molecules is stronger than between similar molecules).

In summary, for non-ideal solutions:

Positive deviation results in a vapor pressure higher than predicted, which can exceed the vapor pressures of the pure components.

Negative deviation results in a vapor pressure lower than predicted, which can also fall below the vapor pressures of the individual pure components.

Question66

Calculate the solubility of a gas having partial pressure 0.15 bar at $25^{\circ}\mathrm{C}.\left[\mathrm{K_H} = 0.16 \text{ moldm}^{-3}\text{bar}^{-1}\right]$

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Options:

A. 2.4×10^{-2} mol dm^{-3}

B. 3.0×10^{-2} mol dm^{-3}

C. 2.7×10^{-2} mol dm^{-3}

D. 1.8×10^{-2} mol dm^{-3}

Answer: A

Solution:

The solubility of a gas in a liquid at a given temperature can be calculated using Henry's law, which is expressed as:

$S = K_H \cdot P$

where:

S is the solubility of the gas (in mol/dm^3),

K_H is the Henry's law constant (in mol/dm^3/bar),

P is the partial pressure of the gas (in bar).

For the given values:

$K_H = 0.16 \, \text{mol dm}^{-3} \, \text{bar}^{-1}$

$P = 0.15 \, \text{bar}$

Substitute the values into the formula:

$S = 0.16 \, \text{mol dm}^{-3} \, \text{bar}^{-1} \times 0.15 \, \text{bar} = 0.024 \, \text{mol dm}^{-3}$

Thus, the solubility of the gas under the given conditions is $2.4 \times 10^{-2}\,\mathrm{mol\,dm^{-3}}$.

Therefore, the correct option is:

Option A

$2.4 \times 10^{-2}\,\mathrm{mol\,dm^{-3}}$

Question67

Which from following ionic solids exhibits decrease in its solubility in water with increase of temperature?

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Options:

A. $NaCl$

B. KNO_3

C. $NaNO_3$

D. Na_2SO_4

Answer: D

Solution:

The ionic solid that exhibits a decrease in solubility in water with an increase in temperature is Na_2SO_4 (Option D).

Sodium sulfate (Na_2SO_4) is known for having a unique solubility behavior. Unlike most other salts, which typically increase in solubility with rising temperature, the solubility of Na_2SO_4 actually decreases after reaching a certain temperature. This is due to its retrograde solubility behavior.

This behavior can be partially explained by the complex ion interactions and the energy changes associated with the dissolution process, where the dissolution becomes less favorable at higher temperatures beyond a certain point.

The solubility curve for sodium sulfate has an unusual shape, and this phenomenon is particularly noticeable at temperatures above its decahydrate form's dissolution, which is around 32.4°C. Beyond this temperature, the solubility tends to decrease, contrasting with the typical solubility patterns observed in compounds like $NaCl$, KNO_3, or $NaNO_3$.

Question68

Calculate molar mass of nonvolatile solute if a solution containing 0.35 g solute in 100 g water has boiling point elevation 0.01 K $\left[K_b = 0.50 \text{ K kg mol}^{-1}\right]$

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Options:

A. 265 g mol^{-1}

B. 175 g mol^{-1}

C. 105 g mol^{-1}

D. 195 g mol^{-1}

Answer: B

Solution:

The boiling point elevation (ΔT_b) can be calculated using the formula:

$\Delta T_b = i \cdot K_b \cdot m$

where:

$\Delta T_b = 0.01$ K is the boiling point elevation,

$K_b = 0.50 \, \text{K kg mol}^{-1}$ is the ebullioscopic constant,

m is the molality of the solution,

i is the van't Hoff factor (for a nonvolatile, non-electrolyte solute, $i = 1$),

The molality m is defined as:

$m = \frac{\text{moles of solute}}{\text{kilograms of solvent}}$

Rearranging the formula to find the molality m:

$m = \frac{\Delta T_b}{i \cdot K_b} = \frac{0.01 \text{ K}}{1 \times 0.50 \text{ K kg mol}^{-1}} = 0.02 \, \text{mol kg}^{-1}$

Given that the mass of the solute is 0.35 g and the mass of the solvent (water) is 100 g or 0.1 kg, we can find the moles of solute:

$m = \frac{\text{moles of solute}}{0.1\text{ kg}} = 0.02\,\text{mol kg}^{-1}$

moles of solute $= 0.02\,\text{mol kg}^{-1} \times 0.1\,\text{kg} = 0.002\,\text{mol}$

The molar mass M of the solute can be calculated by:

$M = \frac{\text{mass of solute}}{\text{moles of solute}} = \frac{0.35\text{ g}}{0.002\text{ mol}} = 175\,\text{g mol}^{-1}$

Therefore, the molar mass of the nonvolatile solute is $175\,\text{g mol}^{-1}$.

Option B: $\boxed{175\,\text{g mol}^{-1}}$ is the correct answer.

Question69

0.2 molal aqueous solution of KCl freezes at -0.680°C. Calculate van't Hoff factor for this solution. $\left(\text{K}_\text{f} = 1.86\text{ K kg mol}^{-1}\right)$

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Options:

A. 1.22

B. 1.32

C. 1.42

D. 1.83

Answer: D

Solution:

To calculate the van't Hoff factor for the given solution, we can use the formula for freezing point depression:

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

where:

ΔT_f is the depression in freezing point,

i is the van't Hoff factor,

K_f is the cryoscopic constant ($1.86\text{ K kg mol}^{-1}$),

m is the molality of the solution (0.2 molal).

Given that the solution freezes at $-0.680°C$, the depression in freezing point is:

$\Delta T_f = 0 - (-0.680) = 0.680\,\text{K}$

Substituting the known values into the equation:

$0.680 = i \cdot 1.86 \cdot 0.2$

Solving for i:

$i = \frac{0.680}{1.86 \cdot 0.2}$

Now, calculating the value:

$i = \frac{0.680}{0.372} \approx 1.83$

Thus, the van't Hoff factor for this solution is approximately **1.83**. Therefore, the correct answer is Option D: **1.83**.

Question70

0.1 molal aqueous solution of glucose boils at $100.16°C$. What is boiling point of 0.5 molal aqueous solution of glucose?

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Options:

A. $100.80°C$

B. $100.16°C$

C. $100.10°C$

D. $20.8°C$

Answer: A

Solution:

For the 0.1 molal glucose solution:

$\Delta T_b = 100.16°C - 100°C = 0.16°C$

Since the solvent is the same for both solutions, the boiling point elevation constant, K_b, remains constant.

Using the boiling point elevation formula:

$\Delta T_b = K_b \times m$

We can derive the relationship:

$\frac{(\Delta T_b)_1}{(\Delta T_b)_2} = \frac{m_1}{m_2}$

For the 0.5 molal solution:

$(\Delta T_b)_2 = \frac{(\Delta T_b)_1 \times m_2}{m_1} = \frac{0.16 \times 0.5}{0.1} = 0.8^{\circ}\mathrm{C}$

Therefore, for the 0.5 molal solution, the change in boiling point, ΔT_b, is $0.8^{\circ}\mathrm{C}$.

Thus, the boiling point of the 0.5 molal glucose solution is:

$100^{\circ}\mathrm{C} + 0.8^{\circ}\mathrm{C} = 100.8^{\circ}\mathrm{C}$

Question71

A solution of non volatile solute has boiling point elevation 0.5 K . Calculate molality of solution $\left[\mathrm{K_b} = 2.40\ \mathrm{K\ kg\ mol^{-1}}\right]$.

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Options:

A. $0.12\ \mathrm{mol\ kg^{-1}}$

B. $0.16\ \mathrm{mol\ kg^{-1}}$

C. $0.21\ \mathrm{mol\ kg^{-1}}$

D. 0.28 mol kg

Answer: C

Solution:

The elevation in boiling point for a solution can be calculated using the formula for boiling point elevation:

$\Delta T_b = i \cdot K_b \cdot m$

where:

ΔT_b is the boiling point elevation,

i is the van't Hoff factor (for non-volatile solute, $i = 1$),

K_b is the ebullioscopic constant of the solvent,

m is the molality of the solution.

Given that:

$\Delta T_b = 0.5\,\mathrm{K}$

$K_b = 2.40\,\mathrm{K\ kg\ mol^{-1}}$

Since the solute is non-volatile, the van't Hoff factor $i = 1$. Substituting these values into the formula:

$0.5 = 1 \cdot 2.40 \cdot m$

Solving for m (molality):

$m = \frac{0.5}{2.40}$

Calculating:

$m = \frac{0.5}{2.40} = 0.2083\,\mathrm{mol\ kg^{-1}}$

Rounding to two decimal places, the molality is approximately:

$m = 0.21\,\mathrm{mol\ kg^{-1}}$

Therefore, the correct answer is **Option C:** $0.21\,\mathrm{mol\ kg^{-1}}$**.**

Question72

Calculate the molar mass of solute in a solution prepared by dissolving 1 gram in $0.3\ \mathrm{dm}^3$ solvent having osmotic pressure 0.2 atm at 300 K.

$$\left[\mathrm{R} = 0.082\ \mathrm{dm}^3\ \mathrm{atm\ K}^{-1}\ \mathrm{mol}^{-1}\right]$$

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Options:

A. $442\ \mathrm{g\ mol}^{-1}$

B. $372\ \mathrm{g\ mol}^{-1}$

C. $390\ \mathrm{g\ mol}^{-1}$

D. 410 g mol^{-1}

Answer: D

Solution:

The molar mass of the solute can be calculated using the formula for osmotic pressure:

$\Pi = \frac{n}{V} RT$

Where:

Π is the osmotic pressure (0.2 atm),

n is the number of moles of solute,

V is the volume of the solution in liters (0.3 dm^3),

R is the gas constant (0.082 dm^3 atm K^{-1} mol^{-1}),

T is the temperature in Kelvin (300 K).

Since the number of moles (n) can be expressed as mass (w) divided by molar mass (M), $n = \frac{w}{M}$, we can substitute this into the osmotic pressure formula:

$\Pi = \frac{w}{MV} RT$

Rearranging to solve for the molar mass (M), we get:

$M = \frac{wRT}{\Pi V}$

Substituting the given values into the equation:

$w = 1 \text{ g}$

$R = 0.082 \text{ dm}^3 \text{ atm K}^{-1} \text{ mol}^{-1}$

$T = 300 \text{ K}$

$\Pi = 0.2 \text{ atm}$

$V = 0.3 \text{ dm}^3$

$M = \frac{1 \times 0.082 \times 300}{0.2 \times 0.3}$

Simplifying this equation:

$M = \frac{24.6}{0.06}$

$M = 410 \text{ g mol}^{-1}$

Therefore, the molar mass of the solute is 410 g mol^{-1}, which corresponds to Option D.

Question73

Assuming complete ionisation, arrange the following solutions in order of increasing osmotic pressure.

a) $0.5 \mathrm{~m} \ \mathrm{Li_2SO_4}$

b) $0.5 \mathrm{~m} \ \mathrm{KCl}$

c) $0.5 \mathrm{~m} \ \mathrm{Al_2(SO_4)_3}$

d) $0.1 \mathrm{~m} \ \mathrm{BaCl_2}$

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Options:

A. d < b < a < c

B. b < d < c < a

C. a < d < c < b

D. c < d < b < a

Answer: A

Solution:

Osmotic pressure is a colligative property that depends on number of particles in solution. The solution having more number of particles will have larger osmotic pressure.

	Solution	Dissociation	
a)	$0.5 \mathrm{~m} \ \mathrm{Li_2SO_4}$	$\mathrm{Li_2SO_4 \rightarrow 2Li^+ + SO_4^{2-}}$	(1.5 m)
b)	$0.5 \mathrm{~m} \ \mathrm{KCl}$	$\mathrm{KCl \longrightarrow K^+ + Cl^-}$	(1.0 m)
c)	$0.5 \mathrm{~mAl_2(SO_4)_3}$	$\mathrm{Al_2(SO_4)_3 \rightarrow 2Al^{3+} + 3SO_4^{2-}}$	(2.5 m)
d)	$0.1 \mathrm{~mBaCl_2}$	$\mathrm{BaCl_2 \rightarrow Ba^{2+} + 2Cl^-}$	(0.3 m)

Hence, $\mathrm{Al_2(SO_4)_3}$ solution gives more number of particles and has the highest osmotic pressure among the given. So, the correct order of increasing osmotic pressure is $d < b < a < c$

Question74

Calculate the cryoscopic constant of solvent when 2.5 gram solute is dissolved in 35 gram solvent lowers its freezing point by 3 K. (molar mass of solute is 117 g mol^{-1})

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Options:

A. $3.11 \text{ K kg mol}^{-1}$

B. $3.56 \text{ K kg mol}^{-1}$

C. $5.52 \text{ K kg mol}^{-1}$

D. $4.91 \text{ K kg mol}^{-1}$

Answer: D

Solution:

The cryoscopic constant (K_f) of a solvent can be calculated using the formula for freezing point depression:

$\Delta T_f = K_f \cdot m$

where:

ΔT_f is the depression in the freezing point,

m is the molality of the solution.

The molality (m) is defined as the number of moles of solute per kilogram of solvent:

Calculate the moles of solute:

$\text{Moles of solute} = \frac{\text{mass of solute}}{\text{molar mass of solute}} = \frac{2.5 \text{ g}}{117 \text{ g/mol}} = \frac{2.5}{117} \text{ mol}$

Calculate the mass of solvent in kilograms:

$\text{Mass of solvent} = 35 \text{ g} = 0.035 \text{ kg}$

Calculate the molality:

$m = \frac{\text{Moles of solute}}{\text{Mass of solvent in kg}} = \frac{\frac{2.5}{117} \text{ mol}}{0.035 \text{ kg}}$

Substitute the values into the formula for freezing point depression to find K_f:

Rearrange the formula to solve for K_f:

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

Substitute the given $\Delta T_f = 3\,\mathrm{K}$:

$K_f = \frac{3}{\frac{2.5}{117} \times \frac{1}{0.035}}$

Calculate K_f:

$K_f = \frac{3}{\left(\frac{2.5}{117}\right) \times \frac{1}{0.035}} = \frac{3}{0.6121} \approx 4.91\,\mathrm{K\,kg\,mol^{-1}}$

The cryoscopic constant of the solvent is approximately $4.91\,\mathrm{K\,kg\,mol^{-1}}$.

Therefore, the answer is Option D: $4.91\,\mathrm{K\,kg\,mol^{-1}}$.

Question75

Identify the reason for the solubility of polar solute in polar solvent from the following.

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Options:

A. solute - solute interactions, solute solvent interactions and solvent - solvent interactions are of similar magnitude.

B. solute - solute interactions > solute solvent interactions > solvent - solvent interactions .

C. solvent- solvent interactions > solute solvent interactions > solute - solute interactions.

D. solute - solvent interactions > solute solute interactions > solvent - solvent interactions.

Answer: A

Solution:

To understand why a polar solute dissolves in a polar solvent, we consider the relative strengths of the intermolecular interactions involved:

Solute-solute interactions: Forces between solute molecules.

Solvent-solvent interactions: Forces between solvent molecules.

Solute-solvent interactions: Forces between solute molecules and solvent molecules.

For the solute to dissolve, the solute-solute and solvent-solvent interactions need to be effectively replaced by solute-solvent interactions. If the solute-solvent interactions are comparable in strength to the original interactions, then no large energy barrier is created, allowing the solute to mix uniformly with the solvent. This scenario is well described by the idea "like dissolves like."

Thus, when the interactions are of roughly equal magnitude, the process of dissolution is energetically favorable (or at least not too unfavorable). This is the underlying reason why polar solutes tend to dissolve in polar solvents.

Answer:

Option A: solute - solute interactions, solute solvent interactions and solvent - solvent interactions are of similar magnitude.

Question76

What is the osmotic pressure of solution prepared by dissolving 3 gram solute in $2\mathrm{dm}^3$ water at 300 K .(Molar mass of solute $= 60 \mathrm{~g} \mathrm{~mol}^{-1}$, $\mathrm{R} = 0.0821 \mathrm{dm}^3$ atm K mol^{-1})

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Options:

A. 0.76 atm

B. 0.62 atm

C. 0.51 atm

D. 0.84 atm

Answer: B

Solution:

The osmotic pressure (Π) of a solution can be calculated using the formula:

$\Pi = \frac{n}{V} RT$

where:

n is the number of moles of solute,

V is the volume of the solution in liters,

R is the universal gas constant,

T is the absolute temperature in Kelvin.

First, calculate the number of moles of solute (n):

$n = \frac{\text{mass of solute}}{\text{molar mass of solute}}$

Substituting the given values:

$n = \frac{3\text{ g}}{60\text{ g/mol}} = 0.05\text{ mol}$

The volume of the solution (V) is given as 2 dm^3, which is equivalent to 2 L.

Given that $R = 0.0821\text{ dm}^3\text{ atm K}^{-1}\text{ mol}^{-1}$ and $T = 300\text{ K}$.

Plug these values into the osmotic pressure formula:

$\Pi = \frac{0.05\text{ mol}}{2\text{ L}} \cdot 0.0821\text{ dm}^3\text{ atm K}^{-1}\text{ mol}^{-1} \cdot 300\text{ K}$

Simplify and calculate:

$\Pi = \frac{0.05}{2} \times 0.0821 \times 300$

$\Pi = 0.025 \times 24.63$

$\Pi = 0.61575\text{ atm}$

Thus, the osmotic pressure of the solution is approximately **0.62 atm**.

Therefore, the correct option is **Option B: 0.62 atm**.

Question77

Calculate the molar mass of solute when 4 g of it dissolved in 1dm^3 solvent has osmotic pressure 2 atm at 300 K . [R $\left[= 0.082\text{ dm}^3\text{ atm K mol}^{-1}\right]$

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Options:

A. 49.2 g mol^{-1}

B. 44.5 g mol^{-1}

C. 54.2 g mol^{-1}

D. 56.4 g mol^{-1}

Answer: A

Solution:

To determine the molar mass of the solute, the van't Hoff equation for osmotic pressure can be used, which is given by:

$\Pi = \frac{n}{V}RT$

where:

Π is the osmotic pressure (2 atm),

n is the number of moles of solute,

V is the volume of the solution in liters ($1 \text{ dm}^3 = 1 \text{ L}$),

R is the universal gas constant ($0.082 \text{ dm}^3 \text{ atm K}^{-1} \text{ mol}^{-1}$),

T is the temperature in Kelvin (300 K).

Rearranging the formula to find n, the number of moles:

$n = \frac{\Pi V}{RT}$

Substitute the values into the equation:

$n = \frac{2 \text{ atm } \times 1 \text{ L}}{0.082 \text{ dm}^3 \text{atm K}^{-1} \text{mol}^{-1} \times 300 \text{ K}}$

Calculate n:

$n = \frac{2}{0.082 \times 300}$

$n = \frac{2}{24.6}$

$n \approx 0.0813 \text{ mol}$

Now, calculate the molar mass (M) of the solute using:

$M = \frac{\text{mass of solute}}{n}$

The mass of the solute is given as 4 g. Thus,

$M = \frac{4 \text{ g}}{0.0813 \text{ mol}}$

$M \approx 49.2 \text{ g mol}^{-1}$

Thus, the molar mass of the solute is 49.2 g mol^{-1}, which corresponds to option A.

Question78

Calculate ΔT_f of aqueous 0.01 m formic acid if van't Hoff factor is 1.1

$$\left[K_f = 1.86 \text{ K Kg mol}^{-1}\right]$$

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Options:

A. 0.042 K

B. 0.020 K

C. 0.011 K

D. 0.033 K

Answer: B

Solution:

To calculate the freezing point depression (ΔT_f) for an aqueous solution of formic acid, we use the formula:

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

where:

i is the van't Hoff factor,

K_f is the cryoscopic constant,

m is the molality of the solution.

Given:

$i = 1.1$ (van't Hoff factor for formic acid),

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

$m = 0.01 \, \mathrm{mol \, kg^{-1}}$.

Plugging these values into the formula:

$\Delta T_f = 1.1 \cdot 1.86\,\mathrm{K\,kg\,mol^{-1}} \cdot 0.01\,\mathrm{mol\,kg^{-1}}$

Calculating the expression:

$\Delta T_f = 1.1 \times 1.86 \times 0.01 = 0.02046\,\mathrm{K}$

Rounded to three significant figures, this is approximately:

Option B: 0.020 K

Question79

What is the relation between the vapour pressure of solution, vapour pressure of solvent and its mole fraction in the solution?

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Options:

A. $\mathrm{P}_1 = \mathrm{P}_1^{\circ} x_1$

B. $\mathrm{P}_1^0 = \mathrm{P}_1 x_1$

C. $\mathrm{P}_1 = \mathrm{P}_1^0 \mathrm{x}_2$

D. $\mathrm{P}_1^0 = \mathrm{P}_1 \mathrm{x}_2$

Answer: A

Solution:

The relation between the vapor pressure of a solution, the vapor pressure of the pure solvent, and its mole fraction in the solution can be described by **Raoult's Law**. Raoult's Law states that the partial vapor pressure of a component in a solution is directly proportional to the mole fraction of the component in the solution and the vapor pressure of the pure component.

Raoult's Law for a solvent in a solution can be expressed as:

$P_1 = P_1^{\circ} x_1$

Where:

P_1 is the vapor pressure of the solvent in the solution.

P_1° is the vapor pressure of the pure solvent.

x_1 is the mole fraction of the solvent in the solution.

Therefore, the correct expression according to Raoult's Law is given by Option A:

$\mathrm{P}_1=\mathrm{P}_1^{\circ} x_1$

Question80

Calculate vapour pressure of a solution containing mixture of 2 moles of volatile liquid A and 3 moles of volatile liquid B at room temperature. $\left(\mathrm{P}_{\mathrm{A}}^{\circ}=420, \mathrm{P}_{\mathrm{B}}^{\circ}=610 \mathrm{~mm} \mathrm{~Hg}\right)$

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Options:

A. 600 mm Hg

B. 570 mm Hg

C. 534 mm Hg

D. 480 mm Hg

Answer: C

Solution:

To calculate the vapor pressure of a solution containing a mixture of two volatile liquids A and B, use Raoult's Law. According to Raoult's Law, the total vapor pressure (P_{total}) of an ideal solution is given by:

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

Where:

x_A and x_B are the mole fractions of components A and B, respectively.

P_A° and P_B° are the vapor pressures of pure components A and B, respectively.

Given:

$P_A^{\circ}=420 \mathrm{~mm} \mathrm{~Hg}$

$P_B^{\circ}=610 \mathrm{~mm} \mathrm{~Hg}$

Moles of A = 2

Moles of B = 3

First, calculate the mole fractions:

Total moles = moles of A + moles of B = 2 + 3 = 5

$x_A = \frac{\text{moles of A}}{\text{total moles}} = \frac{2}{5}$

$x_B = \frac{\text{moles of B}}{\text{total moles}} = \frac{3}{5}$

Substitute the values into Raoult's Law:

$P_{\text{total}} = \left(\frac{2}{5} \cdot 420\right) + \left(\frac{3}{5} \cdot 610\right)$

Calculate each term:

$\frac{2}{5} \cdot 420 = 168$

$\frac{3}{5} \cdot 610 = 366$

Add the partial pressures:

$P_{\text{total}} = 168 + 366 = 534 \text{ mm Hg}$

Thus, the vapor pressure of the solution is **534 mm Hg**.

Correct Option: C 534 mm Hg.

Question81

Calculate the molar mass of non volatile solute when 1 g of it is dissolved in 100 g solvent decreases its freezing point by 0.2 K . $\left[\mathrm{K_f} = 1.2 \mathrm{~K~kg~mol}^{-1}\right]$

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Options:

A. $55 \mathrm{~g~mol}^{-1}$

B. $60 \mathrm{~g~mol}^{-1}$

C. $65 \mathrm{~g~mol}^{-1}$

D. $70 \mathrm{~g~mol}^{-1}$

Answer: B

Solution:

To calculate the molar mass of the non-volatile solute, we can use the formula for freezing point depression:

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

Where:

ΔT_f is the change in freezing point, given as $0.2\,\mathrm{K}$.

i is the van't Hoff factor, which is 1 for non-electrolyte solutes.

K_f is the cryoscopic constant (freezing point depression constant), given as $1.2\,\mathrm{K\ kg\ mol}^{-1}$.

m is the molality of the solution.

First, solve for the molality:

$m = \frac{\Delta T_f}{i \cdot K_f}$

Substituting the known values:

$m = \frac{0.2}{1 \cdot 1.2} = \frac{0.2}{1.2} \approx 0.1667\,\mathrm{mol\ kg}^{-1}$

Molality (m) is defined as moles of solute per kilogram of solvent:

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

Let M be the molar mass of the solute in grams per mole. The moles of solute can be expressed as:

$\text{moles of solute} = \frac{1\,\mathrm{g}}{M}$

Given the mass of the solvent is $100\,\mathrm{g}$ or $0.1\,\mathrm{kg}$, replace into the formula:

$0.1667 = \frac{1/M}{0.1}$

Solve for M:

$0.1667 = \frac{1}{0.1M}$

$0.1667 \cdot 0.1M = 1$

$0.01667M = 1$

Therefore:

$M = \frac{1}{0.01667} \approx 60\,\mathrm{g\ mol}^{-1}$

Hence, the molar mass of the non-volatile solute is approximately $60\,\mathrm{g\ mol}^{-1}$. Thus, the correct option is **Option B**.

Question82

Which of the following solutions on complete dissociation exhibits maximum elevation in boiling point?

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Options:

A. 0.1 m KCl

B. 0.05 m NaCl

C. 0.1 m $BaCl_2$

D. 0.1 m $MgSO_4$

Answer: C

Solution:

	Solution	Moles of particles in 1 kg solution
(A)	0.1 m KCl	0.2
(B)	0.05 m NaCl	0.1
(C)	0.1 m $BaCl_2$	0.3
(D)	0.1 m $MgSO_4$	0.2

0.1 m $BaCl_2$ solution has maximum number of particles in solution, so it exhibits maximum elevation of boiling point.

Question83

The molal elevation boiling point constant for water is $0.513^{\circ}\mathrm{CKgmol}^{-1}$.Calculate boiling point of solution if 0.1 mole of sugar is dissolved in 200 g water?

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Options:

A. $100.513^{\circ}\mathrm{C}$

B. $100.256^{\circ}\mathrm{C}$

C. $100.0513^{\circ}\mathrm{C}$

D. $100.025^{\circ}\mathrm{C}$

Answer: B

Solution:

To find the boiling point of the solution, we need to determine the boiling point elevation, which is given by the formula:

$\Delta T_b = i \cdot K_b \cdot m$

where:

ΔT_b is the boiling point elevation,

i is the van't Hoff factor (which is 1 for sugar, a non-electrolyte),

K_b is the ebullioscopic constant or boiling point elevation constant ($0.513^{\circ}\mathrm{C}\,\mathrm{kg}\,\mathrm{mol}^{-1}$ for water),

m is the molality of the solution.

First, calculate the molality (m):

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

Given that 0.1 mole of sugar is dissolved in 200 g (0.2 kg) of water:

$m = \frac{0.1\text{ mol}}{0.2\text{ kg}} = 0.5\,\text{mol/kg}$

Now, calculate the boiling point elevation:

$\Delta T_b = 1 \cdot 0.513\text{ °C kg/mol} \cdot 0.5\,\text{mol/kg}$

$\Delta T_b = 0.2565\text{ °C}$

The normal boiling point of water is $100^{\circ}\mathrm{C}$. Thus, the boiling point of the solution is:

$T_{\text{solution}} \quad = 100^{\circ}\mathrm{C} + 0.2565^{\circ}\mathrm{C}$

$T_{\text{solution}} \quad \approx 100.256^{\circ}\mathrm{C}$

Therefore, the correct answer is:

Option B

$100.256^{\circ}\mathrm{C}$

Question84

What mass of solute (molar mass $58 \mathrm{~g} \mathrm{~mol}^{-1}$) is to be dissolved in $2.5 \mathrm{~dm}^3 \mathrm{H}_2 \mathrm{O}$ to generate osmotic pressure of 0.245 atm at 300 K ?

$\left(\mathrm{R}=0.0821 \mathrm{~dm}^3 \text{ atm } \mathrm{K}^{-1} \mathrm{~mol}^{-1}\right)$

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Options:

A. 1.0 gram

B. 0.72 gram

C. 1.44 gram

D. 1.75 gram

Answer: C

Solution:

$$
\begin{aligned}
& \Pi=C R T \\
& \Pi=\frac{W_2 \times R \times T}{M_2 \times V} \\
& \mathrm{W}_2=\frac{\Pi \times \mathrm{M}_2 \times \mathrm{V}}{\mathrm{R} \times \mathrm{T}} \\
& \mathrm{W}_2=\frac{0.245 \mathrm{~atm} \times 58 \mathrm{~g} \mathrm{~mol}^{-1} \times 2.5 \mathrm{dm}^3}{0.0821 \mathrm{dm}^3 \mathrm{~atm} \mathrm{~K}^{-1} \mathrm{~mol}^{-1} \times 300 \mathrm{~K}} \\
\therefore \quad & \text{Mass of solute }=1.44 \mathrm{~g}
\end{aligned}
$$

Question85

Which of the following equation correctly represents molar mass of a solute by knowing boiling point elevation?

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Options:

A. $\mathrm{M_2} = \frac{1000 \times \Delta \mathrm{T_b} \times \mathrm{W_2}}{\mathrm{K_b} \times \mathrm{W_1}}$

B. $\mathrm{M_2} = \frac{1000 \times \mathrm{K_b} \times \mathrm{W_1}}{\Delta\ \mathrm{T_b} \times \mathrm{W_2}}$

C. $\mathrm{M_2} = \frac{1000 \times \Delta \mathrm{T_b} \times \mathrm{W_1}}{\mathrm{K_b} \times \mathrm{W_2}}$

D. $\mathrm{M_2} = \frac{1000 \times \mathrm{K_b} \times \mathrm{W_2}}{\Delta\ \mathrm{T_b} \times \mathrm{W_1}}$

Answer: D

Solution:

To determine which equation correctly represents the molar mass (M_2) of a solute using the boiling point elevation (ΔT_b), we need to recall the relationship between these quantities.

Boiling point elevation (ΔT_b) is directly proportional to the molality (m) of the solution:

$\Delta T_b = K_b \times m$

where:

K_b is the ebullioscopic (boiling point elevation) constant of the solvent.

m is the molality of the solution.

Molality (m) is defined as:

$m = \frac{\text{moles of solute}}{\text{kilograms of solvent}} = \frac{W_2/M_2}{W_1/1000} = \frac{1000 \times W_2}{M_2 \times W_1}$

where:

W_2 is the mass of the solute in grams.

M_2 is the molar mass of the solute.

W_1 is the mass of the solvent in grams.

Substitute m back into the boiling point elevation equation:

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

Rearrange the equation to solve for M_2:

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

This equation corresponds to **Option D**.

Answer: Option D

Question86

A solution of non volatile solute has boiling point elevation 1.75 K .Calculate molality of solution $\left[\mathrm{K_b} = 3.5\ \mathrm{K\ kg\ mol^{-1}}\right]$

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Options:

A. $0.77\ \mathrm{mol\ kg^{-1}}$

B. $0.69\ \mathrm{mol\ kg^{-1}}$

C. $0.50\ \mathrm{mol\ kg^{-1}}$

D. $0.35\ \mathrm{mol\ kg^{-1}}$

Answer: C

Solution:

The elevation in boiling point is calculated using the formula:

$\Delta T_b = K_b \cdot m$

where:

ΔT_b is the boiling point elevation,

K_b is the ebullioscopic constant, and

m is the molality of the solution.

Given $\Delta T_b = 1.75\,\mathrm{K}$ and $K_b = 3.5\,\mathrm{K\ kg\ mol^{-1}}$, we can solve for the molality m.

Rearranging the formula to solve for m:

$m = \frac{\Delta T_b}{K_b}$

Substituting the given values:

$m = \frac{1.75\,\mathrm{K}}{3.5\,\mathrm{K\ kg\ mol^{-1}}}$

$m = 0.5\,\mathrm{mol\ kg^{-1}}$

Therefore, the molality of the solution is $\boxed{0.50 \text{ mol kg}^{-1}}$. This corresponds to Option C.

Question87

Which from the following statements is correct for aqueous solution of 6 g L^{-1} urea and $17 \cdot 12 \text{ g L}^{-1}$ of sucrose?

[Molar mass of urea $= 60 \text{ g mol}^{-1}$

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

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Options:

A. Osmotic pressure exhibited by urea solution is lower than that of sucrose solution.

B. Urea solution is hypertonic to sucrose solution.

C. These solutions are isotonic.

D. On doubling the concentration of sucrose solution it becomes hypertonic to urea solution.

Answer: B

Solution:

To compare the properties of the given solutions, calculate their molarity and use this information to determine the osmotic pressure, which is given by the formula:

$\Pi = iCRT$

where Π is the osmotic pressure, i is the van't Hoff factor (1 for non-electrolytes like urea and sucrose), C is the molarity, R is the ideal gas constant, and T is the temperature in Kelvin.

Calculations

Molarity of Urea Solution:

Given mass of urea $= 6 \text{ g L}^{-1}$.

Molar mass of urea $= 60 \text{ g mol}^{-1}$.

Molarity (C_{urea}):

$$C_{\text{urea}} = \frac{6\ \text{g L}^{-1}}{60\ \text{g mol}^{-1}} = 0.1\ \text{mol L}^{-1}$$

Molarity of Sucrose Solution:

Given mass of sucrose $= 17.12\ \text{g L}^{-1}$.

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

Molarity (C_{sucrose}):

$$C_{\text{sucrose}} = \frac{17.12\ \text{g L}^{-1}}{342\ \text{g mol}^{-1}} \approx 0.05\ \text{mol L}^{-1}$$

Comparison of Osmotic Pressures

Since both solutions are at the same temperature and are non-electrolytes, the osmotic pressure directly depends on molarity.

Urea solution has a molarity of $0.1\ \text{mol L}^{-1}$.

Sucrose solution has a molarity of $0.05\ \text{mol L}^{-1}$.

Therefore, the osmotic pressure of the urea solution is higher than that of the sucrose solution.

Conclusion

Option A is incorrect because the urea solution has a higher osmotic pressure than the sucrose solution.

Option B is correct because urea solution, being more concentrated, is hypertonic to the sucrose solution.

Option C is incorrect because the solutions have different osmotic pressures.

Option D is incorrect because doubling the concentration of the sucrose solution to $0.1\ \text{mol L}^{-1}$ would make it isotonic to the original urea solution, not hypertonic.

Therefore, the correct answer is **Option B: Urea solution is hypertonic to sucrose solution**.

Question88

Calculate the molar mass of nonvolatile solute when 1.5 g of it is dissolved in 90 g solvent decreases its freezing point by 0.25 K. $\left[\mathrm{K_f} = 1.2\ \mathrm{K\ kg\ mol^{-1}}\right]$

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Options:

A. 72 g mol^{-1}

B. 80 g mol^{-1}

C. 88 g mol^{-1}

D. 96 g mol^{-1}

Answer: B

Solution:

To calculate the molar mass of the nonvolatile solute, we can use the formula for freezing point depression:

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

Where:

ΔT_f is the change in freezing point,

i is the van't Hoff factor, which is 1 for nonionizing solutes,

K_f is the cryoscopic constant,

m is the molality of the solution.

Given:

$\Delta T_f = 0.25\,\mathrm{K}$,

$K_f = 1.2\,\mathrm{K\ kg\ mol}^{-1}$,

Mass of solute = 1.5 g,

Mass of solvent = 90 g $= 0.090\,\mathrm{kg}$.

First, rearrange the equation to solve for molality (m):

$m = \frac{\Delta T_f}{i \cdot K_f}$

Substitute the known values:

$m = \frac{0.25}{1 \times 1.2} = \frac{0.25}{1.2} \approx 0.2083\,\mathrm{mol\ kg}^{-1}$

Next, use the definition of molality:

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

Substitute the values:

$0.2083 = \frac{\frac{1.5}{M}}{0.090}$

Here, M is the molar mass of the solute. Rearrange to solve for M:

$M = \frac{1.5}{0.2083 \times 0.090}$

Calculate:

$M = \frac{1.5}{0.01875} \approx 80 \mathrm{~g} \mathrm{~mol}^{-1}$

Therefore, the molar mass of the solute is $80 \mathrm{~g} \mathrm{~mol}^{-1}$, which corresponds to option B.

Question89

Identify false statement form following.

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Options:

A. The boiling point of solution containing non volatile solute is always higher than pure solvent.

B. At any temperature the vapour pressure of solution containing non volatile solute is lower than that of pure solvent.

C. The boiling point of liquid is the temperature at which its vapour pressure equals atmospheric pressure.

D. The molal elevation constant is the boiling point elevation produced by 1 molar solution.

Answer: D

Solution:

Option A:

> "The boiling point of solution containing non-volatile solute is always higher than pure solvent."

This is **true** (colligative property: elevation of boiling point).

Option B:

> "At any temperature, the vapour pressure of solution containing non-volatile solute is lower than that of pure solvent."

This is **true** (Raoult's law: addition of non-volatile solute lowers vapour pressure).

Option C:

> "The boiling point of liquid is the temperature at which its vapour pressure equals atmospheric pressure."

This is **true** (definition of boiling point).

Option D:

> "The molal elevation constant is the boiling point elevation produced by 1 molar solution."

This is **false.**

Reason: The molal elevation constant (Kb) is defined as the elevation in boiling point produced when **1 mole of solute is dissolved in 1 kg of solvent**, i.e. a **1 molal solution**, not "1 molar" solution (which is 1 mole solute in 1 L solution—different concentration unit).

✅ **Answer: The false statement is Option D**

Question90

Calculate van't Hoff factor of 0.2 m aquesous solution of an electrolyte if it freezes at -0.7 K $\left[\mathrm{K_f} = 1.86 \text{ K kg mol}^{-1}\right]$

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Options:

A. 1.304

B. 1.123

C. 1.432

D. 1.882

Answer: D

Solution:

To calculate the van't Hoff factor (i) of the solution, the formula for the depression in freezing point is used:

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

where:

ΔT_f is the depression in freezing point,

i is the van't Hoff factor,

K_f is the cryoscopic constant or freezing point depression constant,

m is the molality of the solution.

Given:

$\Delta T_f = 0.7\,\mathrm{K}$

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

$m = 0.2\,\mathrm{m}$

Rearranging the equation to solve for i, we have:

$i = \frac{\Delta T_f}{K_f \cdot m}$

Substituting in the given values:

$i = \frac{0.7}{1.86 \times 0.2}$

Calculate the expression:

Calculate the denominator:

$1.86 \times 0.2 = 0.372$

Divide the depression in freezing point by the result:

$i = \frac{0.7}{0.372} \approx 1.882$

The van't Hoff factor of the solution is approximately 1.882, which corresponds to Option D.

Question91

Calculate the molar mass of non volatile solute when 5 g of it is dissolved in 50 g solvent, boils at $119.6^{\circ}\mathrm{C}$. $\left[\mathrm{K_b} = 3.2\ \mathrm{K\ kg\ mol^{-1}}\right.$, boiling point of pure solvent $= 118^{\circ}\mathrm{C}$].

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Options:

A. $180\ \mathrm{g\ mol^{-1}}$

B. $210\ \mathrm{g\ mol^{-1}}$

C. $200\ \mathrm{g\ mol^{-1}}$

D. $190\ \mathrm{g\ mol^{-1}}$

Answer: C

Solution:

$$\Delta T_b = T_b - T_b^0 = 119.6 - 118 = 1.6^\circ C = 1.6\ K$$
$$\Delta T_b = \frac{1000\ K_b\ W_2}{M_2\ W_1}$$
$$M_2 = \frac{1000\ K_b\ W_2}{\Delta\ T_b\ W_1} = \frac{1000 \times 3.2\ K\ kg\ mol^{-1} \times 5\ g}{1.6\ K \times 50\ g}$$
$$= 200\ g\ mol^{-1}$$

Question92

Identify an example of solution that consists of solid as solute and liquid as solvent.

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Options:

A. Sea water

B. Bronze

C. Carbonated water

D. Chloroform in nitrogen

Answer: A

Solution:

An example of a solution that consists of a solid as the solute and a liquid as the solvent is:

Option A: Sea water

In sea water, the solid solute is various salts (primarily sodium chloride), which dissolve in the liquid solvent, water.

Question93

Calculate the molality of solution of non volatile solute having depression in freezing point 0.93 K and cryoscopic constant of solvent $1.86 \text{ K kg mol}^{-1}$**.**

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Options:

A. 0.3 mol kg^{-1}

B. 0.5 mol kg^{-1}

C. 0.4 mol kg^{-1}

D. 0.6 mol kg^{-1}

Answer: B

Solution:

The depression in freezing point (ΔT_f) of a solution can be calculated using the formula:

$\Delta T_f = K_f \times m$

where:

ΔT_f is the depression in freezing point,

K_f is the cryoscopic constant of the solvent,

m is the molality of the solution.

Given:

$\Delta T_f = 0.93 \, \text{K}$

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

We can rearrange the formula to solve for molality (m):

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

Substitute the given values:

$m = \frac{0.93 \text{ K}}{1.86 \text{ K kg mol}^{-1}}$

Calculate:

$m = \frac{0.93}{1.86} \mathrm{~mol~kg}^{-1} = 0.5 \mathrm{~mol~kg}^{-1}$

The molality of the solution is $0.5 \mathrm{~mol~kg}^{-1}$, which corresponds to Option B.

Question94

A solution of non volatile solute is obtained by dissolving 2 g in 50 g benzene. Calculate the vapour pressure of solution if vapour pressure of pure benzene is 640 mmHg at $25^{\circ}\mathrm{C}$.[mol. mass of benzene $= 78 \mathrm{~g~mol}^{-1}$, mol. mass of solute $= 64 \mathrm{~g~mol}^{-1}$]

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Options:

A. 600.21 mm Hg

B. 604.52 mm Hg

C. 608.64 mm Hg

D. 612.83 mm Hg

Answer: C

Solution:

Relative lowering in vapour pressure

$$\frac{\Delta \mathrm{P}}{\mathrm{P}^0} = \frac{\mathrm{P}^0 - \mathrm{P}}{\mathrm{P}^0} = \mathrm{x}_2 = \frac{\mathrm{W}_2 \mathrm{M}_1}{\mathrm{W}_1 \mathrm{M}_2}$$

$$\therefore \quad \frac{640 - \mathrm{P}}{640} = \frac{2 \times 78}{50 \times 64}$$

$$640 - \mathrm{P} = 640 \times 0.048$$

$$\mathrm{P} = 640 - 31.2$$

$$\mathrm{P} = 608.8 \mathrm{~mmHg}$$

Question95

What is Henry's law constant of a gas if solubility of gas in water at $25^{\circ}\mathrm{C}$ is $0.028 \mathrm{~mol~dm}^{-3}$?

[Partial pressure of gas $= 0.346$ bar]

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Options:

A. $0.081 \mathrm{~mol~dm}^{-3} \mathrm{bar}^{-1}$

B. $0.075 \mathrm{~mol~dm}^{-3} \mathrm{bar}^{-1}$

C. $0.093 \mathrm{~mol~dm}^{-3} \mathrm{bar}^{-1}$

D. $0.049 \mathrm{~mol~dm}^{-3} \mathrm{bar}^{-1}$

Answer: A

Solution:

According to Henry's Law, the solubility of a gas in a liquid at a particular temperature is directly proportional to the pressure of the gas above the liquid. The law can be mathematically expressed as:

$S = k_H \cdot P$

Where:

- S is the solubility of the gas (in $\mathrm{mol~dm}^{-3}$).
- k_H is Henry's law constant (in $\mathrm{mol~dm}^{-3} \mathrm{bar}^{-1}$).
- P is the partial pressure of the gas (in bar).

To find Henry's law constant for the given conditions, we need to rearrange the equation to solve for k_H:

$k_H = \frac{S}{P}$

Given that the solubility $S = 0.028 \mathrm{~mol~dm}^{-3}$ and the partial pressure $P = 0.346$ bar, we can plug these values into the equation:

$k_H = \frac{0.028 \mathrm{~mol~dm}^{-3}}{0.346 \mathrm{~bar}}$

To calculate k_H:

$k_H = \frac{0.028}{0.346} \mathrm{~mol~dm}^{-3} \mathrm{bar}^{-1}$

$k_H \approx 0.08092 \mathrm{~mol~dm}^{-3} \mathrm{bar}^{-1}$

If you review the options provided, the closest value to 0.08092 is option A, 0.081 mol $dm^{-3}bar^{-1}$. Therefore, the correct option is:

Option A 0.081 mol $dm^{-3}bar^{-1}$

Question96

A solution of nonvolatile solute is obtained by dissolving $15 \mathrm{~g}$ in $200 \mathrm{~mL}$ water has depression in freezing point $0.75 \mathrm{~K}$. Calculate the molar mass of solute if cryoscopic constant of water is $1.86 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$.

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Options:

A. $160 \mathrm{~g} \mathrm{~mol}^{-1}$

B. $172 \mathrm{~g} \mathrm{~mol}^{-1}$

C. $186 \mathrm{~g} \mathrm{~mol}^{-1}$

D. $198 \mathrm{~g} \mathrm{~mol}^{-1}$

Answer: C

Solution:

$$\begin{aligned} \mathrm{M}_2 & =\frac{1000 \mathrm{~K}_{\mathrm{f}} \mathrm{W}_2}{\Delta \mathrm{T}_{\mathrm{f}} \mathrm{W}_1} \\ & =\frac{1000 \times 1.86 \times 15}{0.75 \times 200} \\ & =186 \mathrm{~g} \mathrm{~mol}^{-1}\end{aligned}$$

Question97

What type of following solutions is the gasoline?

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Options:

A. Liquid as solute and liquid as solvent

B. Liquid as solute and solid as solvent

C. Solid as solute and liquid as solvent

D. Gas as solute and liquid as solvent

Answer: A

Solution:

Gasoline is a homogeneous mixture of liquid hydrocarbons. It is primarily used as a fuel in internal combustion engines. Given the nature of its composition, the solution itself is liquid, consisting of many different liquid organic compounds (hydrocarbons and additives) mixed together. Thus, each component can be considered a liquid solute that dissolves in a liquid solvent—other hydrocarbons that make up the bulk of the gasoline. Therefore, the correct option for the type of solution that gasoline represents is:

Option A: Liquid as solute and liquid as solvent

Gasoline does not fit the descriptions for options B, C, or D since it does not contain solid solutes or solvents, nor does it have gases dissolved in it under normal conditions to any significant extent that would classify it as a gas in liquid solution.

Question98

Calculate the relative lowering of vapour pressure if the vapour pressure of benzene and vapour pressure of solution of non-volatile solute in benzene are 640 mmHg and 590 mmHg respectively at same temperature.

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Options:

A. 0.078

B. 0.175

C. 0.061

D. 0.092

Answer: A

Solution:

Relative lowering of vapour pressure

$$= \frac{\Delta P}{P_1^0} = \frac{P_1^0 - P_1}{P_1^0}$$
$$= \frac{640 - 590}{640} = 0.078$$

Question99

Calculate the depression in freezing point of solution when $4 \mathrm{~g}$ nonvolatile solute of molar mass $126 \mathrm{~g} \mathrm{~mol}^{-1}$ dissolved in $80 \mathrm{~mL}$ water [Cryoscopic constant of water $= 1.86 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$]

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Options:

A. $0.55 \mathrm{~K}$

B. $0.74 \mathrm{~K}$

C. $0.86 \mathrm{~K}$

D. $0.90 \mathrm{~K}$

Answer: B

Solution:

To calculate the depression in freezing point, we use the formula derived from the concept of colligative properties:

$\Delta T_f = K_f \cdot \frac{w_2}{M_2} \cdot \frac{1}{w_1}$

Here:

- ΔT_f is the depression in freezing point.
- K_f is the cryoscopic constant of the solvent (water), which is $1.86 \text{ K kg mol}^{-1}$.
- w_2 is the mass of the solute, which is 4 g.
- M_2 is the molar mass of the solute, which is 126 g mol^{-1}.
- w_1 is the mass of the solvent, which we need to calculate from the given volume of water 80 mL.

First, we convert the volume of water to mass:

Since the density of water is approximately 1 g/mL,

$w_1 = 80 \text{ mL} \times 1 \text{ g/mL} = 80 \text{ g} = 0.08 \text{ kg}$

Now, substituting the values into the formula:

$\Delta T_f = 1.86 \cdot \frac{4}{126} \cdot \frac{1}{0.08}$

Simplify the expression step-by-step:

First calculate the fraction:

$\frac{4}{126} = 0.03175$

Then:

$\Delta T_f = 1.86 \cdot 0.03175 \cdot 12.5$

Next, we multiply 0.03175 by 12.5:

$0.03175 \cdot 12.5 = 0.396875$

Finally:

$\Delta T_f = 1.86 \cdot 0.396875 = 0.7381875 \approx 0.74 \text{ K}$

Therefore, the correct answer is:

Option B: 0.74 K

Question100

Which among the following is NOT colligative property?

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Options:

A. Vapour pressure lowering

B. Boiling point

C. Freezing point depression

D. Osmotic pressure

Answer: B

Solution:

Colligative properties are those properties of solutions that depend on the number of dissolved particles in solution, but not on their identity. Among the options given, all except Option B (Boiling point) refer directly to colligative properties.

Before clarifying why boiling point alone isn't the correct answer, let's define each term:

Vapour Pressure Lowering (Option A): A colligative property that refers to the decrease in the vapor pressure of a solvent when a non-volatile solute is dissolved in it compared to the vapor pressure of the pure solvent.

Freezing Point Depression (Option C): Another colligative property which states that the freezing point of a solvent will be lower when a solute is dissolved in it compared to the freezing point of the pure solvent.

Osmotic Pressure (Option D): Also a colligative property, osmotic pressure is the pressure required to stop the flow of solvent molecules through a semipermeable membrane from a region of low solute concentration to a region of high solute concentration.

Question101

What is the molality of solution of a non-volatile solute having boiling point elevation $7.15 \mathrm{~K}$ and molal elevation constant $2.75 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$?

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Options:

A. 3.2 m

B. 2.0 m

C. 2.6 m

D. 3.8 m

Answer: C

Solution:

$\Delta T_b = 7.15$ K

$K_b = 2.75$ K kg mol^{-1}

$m = ?$

Using, $\Delta T_b = K_b m$

$\Rightarrow m = \frac{\Delta T_b}{K_b} = \frac{7.15}{2.75} = 2.6 \mathrm{~m} \Rightarrow m = 2.6$ molal

----

Question102

Which among the following salts exhibits inverse relation between it's solubility and temperature?

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Options:

A. NaBr

B. $NaNO_3$

C. KNO_3

D. Na_2SO_4

Answer: D

Solution:

The dissolution process for Na_2SO_4 is exothemic while for $NaBr, NaNO_3$ and KNO_3 is endothermic in nature. Therefore, solubility decrease with increase in temperature for Na_2SO_4 salt. Thus, it exhibits inverse relation between its solubility and temperature.

Question103

If 0.01 m aqueous solution of an electrolyte freezes at -0.056 K. Calculate van't Hoff factor for an electrolyte (cryoscopic constant of water $= 1.86$ K kg mol^{-1})

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Options:

A. 1.30

B. 2.33

C. 3.00

D. 4.11

Answer: C

Solution:

$$\begin{aligned} \text{Given, molality } &= 0.01 \text{ m} \\ \Delta T_f &= -0.056 \text{ K} \\ K_f &= 1.86 \text{ K kg mol}^{-1} \\ i &= ? \end{aligned}$$

Using formula,

$$\Delta T_f = i K_f m$$

$$\Rightarrow \quad i = \frac{0.056}{1.86 \times 0.01} = 3$$

$$\therefore \quad i = 3$$

Question104

Calculate van't Hoff factor of K_2SO_4 if 0.1 m aqueous solution of K_2SO_4 freezes at $-0.43°C$ and cryoscopic constant of water is $1.86 \text{ K kg mol}^{-1}$.

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Options:

A. 2.3

B. 2.7

C. 3.1

D. 3.5

Answer: A

Solution:

$$\begin{aligned} & \Delta T_f = iK_f m \\ \therefore \quad & 0.43 = i \times 1.86 \times 0.1 \\ \therefore \quad & i = \frac{0.43}{1.86 \times 0.1} = 2.3 \end{aligned}$$

[Note: In the question, the freezing point of aqueous solution is changed from -0.43 K to -0.43°C to apply appropriate textual concepts.]

Question105

A solution of 5.6 g non-volatile solute in 50 g solvent has elevation in boiling point 1.75 K. What is the molar mass of solute $\left(\mathrm{K_b} = 3\ \mathrm{K\ kg\ mol^{-1}}\right)$?

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Options:

A. $192\ \mathrm{g\ mol^{-1}}$

B. $200\ \mathrm{g\ mol^{-1}}$

C. $184\ \mathrm{g\ mol^{-1}}$

D. $176\ \mathrm{g\ mol^{-1}}$

Answer: A

Solution:

$$\begin{aligned} \mathrm{M_2} &= \frac{1000\ \mathrm{K_b W_2}}{\Delta \mathrm{T_b W_1}} \\ &= \frac{1000 \times 3 \times 5.6}{1.75 \times 50} = 192\ \mathrm{g\ mol^{-1}} \end{aligned}$$

Question106

What is the molal elevation constant if one gram mole of a nonvolatile solute is dissolved in $1 \mathrm{~kg}$ of ethyl acetate? $\left(\Delta \mathrm{T}_{\mathrm{b}}=x \mathrm{~K}\right)$

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Options:

A. $x \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

B. $\frac{x}{2} \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

C. $2 x \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

D. $3 x \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

Answer: A

Solution:

$$
\begin{aligned}
& \mathrm{M}_2=\frac{1000 \mathrm{~K}_{\mathrm{b}} \mathrm{W}_2}{\Delta \mathrm{T}_{\mathrm{b}} \mathrm{W}_1} \\
\therefore \quad & \mathrm{K}_{\mathrm{b}}=\frac{\mathrm{M}_2 \Delta \mathrm{T}_{\mathrm{b}} \mathrm{W}_1}{1000 \mathrm{~W}_2} \\
& \text { Now, } \mathrm{W}_2=\text { one gram mole }=\mathrm{M}_2 \mathrm{~g} \\
& \mathrm{W}_1=1 \mathrm{~kg}=1000 \mathrm{~g} \\
\therefore \quad & \mathrm{K}_{\mathrm{b}}=\frac{\mathrm{M}_2 \times x \times 1000}{1000 \times \mathrm{M}_2}=x \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}
\end{aligned}
$$

One gram mole solute dissolved in $1 \mathrm{~kg}$ solvent $=1$ molal solution

When concentration of solution is 1 molal, elevation in boiling point $\left(\Delta T_b\right)$ is equal to molal elevation constant $\left(\mathrm{K}_{\mathrm{b}}\right)$.

Therefore, $\mathrm{K}_{\mathrm{b}}=x \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

Question107

If $0.15 \mathrm{~m}$ aqueous solution of KCI freezes at $-0.511^{\circ} \mathrm{C}$, calculate van't Hoff factor of KCI (cryoscopic constant of water is

$1.86 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$)

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Options:

A. 1.45

B. 1.26

C. 1.82

D. 3.00

Answer: C

Solution:

$$\begin{aligned} & \Delta \mathrm{T}_f=\mathrm{T}_{\mathrm{f}}^0-\mathrm{T}_{\mathrm{f}}=0-\left(-0.51^{\circ} \mathrm{C}\right)=0.51^{\circ} \mathrm{C}=0.51 \mathrm{~K} \\ \therefore \quad & \Delta \mathrm{T}_{\mathrm{f}}=\mathrm{iK}_{\mathrm{f}} \mathrm{m} \\ \therefore \quad & \mathrm{i}=\frac{\Delta \mathrm{T}}{\mathrm{K}_{\mathrm{f}} \mathrm{m}}=\frac{0.51 \mathrm{~K}}{1.86 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1} \times 0.15 \mathrm{~mol} \mathrm{~kg}^{-1}}=1.82\end{aligned}$$

Question108

What is the solubility of gas in water at $25^{\circ} \mathrm{C}$ if partial pressure is 0.346 bar [Henry's law constant is $\left[0.159 \mathrm{~mol} \mathrm{dm}^{-3} \mathrm{~bar}^{-1}\right]$?

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Options:

A. $0.055 \mathrm{~mol} \mathrm{dm}^{-3}$

B. $0.028 \mathrm{~mol} \mathrm{dm}^{-3}$

C. $0.083 \mathrm{~mol} \mathrm{dm}^{-3}$

D. $0.11 \mathrm{~mol} \mathrm{dm}^{-3}$

Answer: A

Solution:

$$S = K_H P = 0.159 \text{ mol dm}^{-3} \text{ bar}^{-1} \times 0.346 \text{ bar}$$
$$= 0.055 \text{ mol dm}^{-3}$$

Question109

The partial vapour pressure of any volatile component of a solution is equal to the vapour pressure of the pure component multiplied by its mole fraction in the solution is called

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Options:

A. Dalton's law

B. Avogadro's law

C. Raoult's law

D. Henry's law

Answer: C

Solution:

The statement that "The partial vapor pressure of any volatile component of a solution is equal to the vapor pressure of the pure component multiplied by its mole fraction in the solution" is known as Raoult's law. Therefore, the correct answer is:

Option CRaoult's law

Raoult's law can be mathematically expressed as:

$$P_A = P_A^* \cdot X_A$$

Where:

- P_A = partial vapor pressure of component A in the solution
- P_A^* = vapor pressure of the pure component A
- X_A = mole fraction of component A in the solution

This law applies to ideal solutions, where the interactions between molecules of different components are nearly the same as the interactions between molecules of the same components. In such mixtures, the presence of other components does not affect the vapor pressure of any given component significantly, other than through the mole fraction. Actual solutions may deviate from Raoult's law, especially at higher concentrations or with components that interact strongly with each other.

Question110

Which among the following solutions has minimum boiling point elevation?

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Options:

A. 0.1 m NaCl

B. 0.2 m KNO_3

C. 0.1 m Na_2SO_4

D. 0.05 m $CaCl_2$

Answer: D

Solution:

$NaCl \longrightarrow Na^+ + Cl^-$

$\text{Total ions} = 0.1 + 0.1 = 0.2 \text{ ions}$

$KNO_3 \longrightarrow K^+ + NO_3^-$

$\text{Total ions} = 0.2 + 0.2 = 0.4 \text{ ions}$

$Na_2SO_4 \longrightarrow 2Na^+ + SO_4^{2-}$

$\text{Total ions} = 0.2 + 0.1 = 0.3 \text{ ions}$

$CaCl_2 \longrightarrow Ca^{2+} + 2Cl^-$

$\text{Total ions} = 0.05 + 0.1 = 0.15 \text{ ions}$

0.05 $mCaCl_2$ solution has minimum ions in solution, so it shows minimum boiling point elevation.

Question111

Calculate osmotic pressure of solution of 0.025 mole glucose in $100 \mathrm{~mL}$ **water at** $300 \mathrm{~K}$**.** $\left[\mathrm{R}=0.082 \mathrm{~atm dm mol}^{-1} \mathrm{~K}^{-1}\right]$

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Options:

A. 1.54 atm

B. 2.05 atm

C. 6.15 atm

D. 3.08 atm

Answer: C

Solution:

$$\begin{aligned}
\pi & =\mathrm{MRT}=\frac{\mathrm{n}_2 \mathrm{RT}}{\mathrm{V}} \\
& =\frac{0.025 \mathrm{~mol} \times 0.082 \mathrm{~dm}^3 \mathrm{~atm} \mathrm{~mol}^{-1} \mathrm{~K}^{-1} \times 300 \mathrm{~K}}{0.1 \mathrm{dm}^3} \\
& =6.15 \mathrm{~atm}
\end{aligned}$$

Question112

Calculate molality of solution of a nonvolatile solute having boiling point elevation $1.89 \mathrm{~K}$ **if boiling point elevation constant of solvent is** $3.15 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$**.**

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Options:

A. 0.4 m

B. 0.8 m

C. 0.6 m

D. 0.3 m

Answer: C

Solution:

$$\begin{aligned} & \Delta \mathrm{T_b} = \mathrm{K_b} \times \mathrm{m} \\ & 1.89 = 3.15 \times \mathrm{m} \\ \therefore \quad & \mathrm{m} = \frac{1.89}{3.15} = 0.6 \mathrm{~mol~kg}^{-1} \end{aligned}$$

Question113

A solution of nonvolatile solute is obtained by dissolving $1.5 \mathrm{~g}$ in $30 \mathrm{~g}$ solvent has boiling point elevation $0.65 \mathrm{~K}$. Calculate the molal elevation constant if molar mass of solute is $150 \mathrm{~g~mol}^{-1}$.

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Options:

A. $1.95 \mathrm{~K~kg~mol}^{-1}$

B. $2.23 \mathrm{~K~kg~mol}^{-1}$

C. $1.52 \mathrm{~K~kg~mol}^{-1}$

D. $2.72 \mathrm{~K~kg~mol}^{-1}$

Answer: A

Solution:

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

$$K_b = \frac{M_2 \times \Delta T_b \times W_1}{1000 \times W_2} = \frac{150 \times 0.65 \times 30}{1000 \times 1.5}$$

$$= 1.95\ K\ kg\ mol^{-1}$$

Question114

Identify the FALSE statement about ideal solution from following.

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Options:

A. Ideal solutions obey Raoult's law over entire range of concentration.

B. No heat is evolved or absorbed when two components forming an ideal solution are mixed.

C. Volume of ideal solution is same as sum of volumes of two components taken for mixing.

D. The vapour pressure of ideal solution is either higher or lower than vapour pressure of pure components.

Answer: D

Solution:

The vapour pressure of ideal solution always lies between vapour pressures of pure components.

Question115

Calculate osmotic pressure of $0.2\ M$ aqueous KCl solution at $0^{\circ}C$ if van't Hoff factor for KCl is 1.83. $\left[R = 0.082\ dm^3\ atm\ mol^{-1}\ K^{-1}\right]$

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Options:

A. 8.2 atm

B. 9.4 atm

C. 10.6 atm

D. 6.5 atm

Answer: A

Solution:

$$\begin{aligned} \pi &= \text{iMRT} \\ &= 1.83 \times 0.2 \times 0.082 \times 273 \\ &= 8.2 \text{ atm} \end{aligned}$$

Question116

If K_b denote molal elevation constant of water, then boiling point of an aqueous solution containing 36 g glucose (molar mass = 180) per dm^3 is:

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Options:

A. $(100 + K_b)^\circ C$

B. $(100 + 2\ K_b)^\circ C$

C. $\left(100 + \frac{K_b}{10}\right)^\circ C$

D. $\left(100 + \frac{2\ K_b}{10}\right)^\circ C$

Answer: D

Solution:

The aqueous solution contains 36 g glucose per dm^3, so mass of solute W_2 is 36 g.

Assuming that the density of solution is 1 g/dm^3, the mass of solvent (water) is 1000 g.

$$\Delta T_b = \frac{1000\ K_b W_2}{M_2\ W_1}$$
$$\Delta T_b = \frac{1000\ g\ kg^{-1} \times K_b \times 36\ g}{180\ g \times 1000\ g}$$
$$\Delta T_b = \frac{2\ K_b}{10}$$
$$\Delta T_b = T_b - T_b^\circ$$
$$T_b = T_b^\circ + \Delta T_b$$
$$\therefore\quad T_b = \left(100 + \frac{2\ K_b}{10}\right)^\circ C$$

Question117

What is vapour pressure of a solution containing 1 mol of a nonvolatile solute in 36 g of water $\left(P_1^0 = 32\ \text{mm Hg}\right)$?

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Options:

A. 8.14 mm Hg

B. 12.31 mm Hg

C. 16.08 mm Hg

D. 21.44 mm Hg

Answer: D

Solution:

$$n_2 = 1\ \text{mol}$$
$$n_1 = \frac{36}{18} = 2\ \text{mol}$$

Relative lowering of vapour pressure

$$= \frac{P_1^0 - P_1}{P_1^0} = x_2 = \frac{n_2}{n_1 + n_2}$$

$\therefore \quad \frac{32 \text{ mmHg} - P_1}{32 \text{ mmHg}} = \frac{1}{3}$

$96 \text{ mm Hg} - 3P_1 = 32 \text{ mm Hg}$

$64 \text{ mm Hg} = 3P_1$

$\therefore \quad P_1 = 21.33 \text{ mm Hg} \approx 21.44 \text{ mm Hg}$

Question118

Which among following salts shows decrease in solubility with increase in temperature?

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Options:

A. Na_2SO_4

B. KNO_3

C. $NaNO_3$

D. KBr

Answer: A

Solution:

Dissolution of Na_2SO_4 in water is an exothermic process. When a substance dissolves in water by an exothermic process, its solubility decreases with an increase in temperature. Hence, solubility of Na_2SO_4 in water decreases with increase in temperature.

Question119

The solubility of a gas in a liquid is directly proportional to the pressure of the gas over the solution. Identify the law for this statement.

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Options:

A. Henry's law

B. Raoult's law

C. Dalton's law

D. Avogadro's law

Answer: A

Solution:

The correct law is:

A) Henry's law

Explanation:
Henry's law states that *the solubility of a gas in a liquid is directly proportional to the pressure of that gas above the liquid* , at constant temperature.

So the correct answer is **Henry's law**

Question120

What is the solubility of a gas in water at $25^{\circ}\mathrm{C}$ if partial pressure is $0.18 \mathrm{~atm}$?

$$\left(\mathrm{K}_{\mathrm{H}}=0.16 \mathrm{~moldm}^{-3} \mathrm{~atm}^{-1}\right)$$

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Options:

A. $0.029 \mathrm{~mol~dm}^{-3}$

B. $0.022 \mathrm{~mol~dm}^{-3}$

C. $0.032 \mathrm{~mol~dm}^{-3}$

D. $0.038 \mathrm{~mol~dm}^{-3}$

Answer: A

Solution:

$$\begin{aligned} S &= K_H \times P \\ &= 0.16 \times 0.18 \\ &= 0.0288 \text{ mol dm}^{-3} \end{aligned}$$

Question121

What is osmotic pressure of solution of $1.7 \text{ g } CaCl_2$ in 1.25 dm^3 water at 300 K if van't Hoff factor and molar mass of $CaCl_2$, are 2.47 and 111 g mol^{-1} respectively?

$$\left[R = 0.082 \text{ dm}^3 \text{ atm mol}^{-1} \text{ K}^{-1}\right]$$

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Options:

A. 0.625 atm

B. 0.744 atm

C. 0.827 atm

D. 0.936 atm

Answer: B

Solution:

$\pi = \text{iMRT} = \frac{i \times W_2 RT}{M_2 V}$

$\pi = \frac{2.47 \times 1.7 \text{ g} \times 0.082 \text{ dm}^3 \text{ atm mol}^{-1} \text{ K}^{-1} \times 300 \text{ K}}{111 \text{ g mol}^{-1} \times 1.25 \text{ dm}^3} = 0.744 \text{ atm}$

Question122

What type of solution is the ethyl alcohol in water?

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Options:

A. Liquid in solid

B. Solid in liquid

C. Liquid in liquid

D. Gas in liquid

Answer: C

Solution:

Ethyl alcohol (ethanol) dissolved in water forms a **homogeneous mixture**, which is a type of solution where the solute and the solvent are both in the same phase. In this case, both ethanol and water are **liquid** at room temperature. Therefore, when ethanol is mixed with water, you get a liquid dissolved in another liquid. This type of solution is commonly known as a **liquid in liquid** solution.

Based on this information, the correct option is:

Option C - Liquid in liquid

Question123

Find the depression in freezing point of solution when 3.2 gram non volatile solute with molar mass $128 \text{ gram mol}^{-1}$ is dissolved in 80 gram solvent if cryoscopic constant of solvent is $4.8 \text{ K kg mol}^{-1}$.

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Options:

A. 3.0 K

B. 1.5 K

C. 2.0 K

D. 2.5 K

Answer: B

Solution:

$$
\begin{aligned}
\Delta T_f &= \frac{1000\ K_f W_2}{M_2\ W_1} \\
&= \frac{1000\ g\ kg^{-1} \times 4.8\ K\ kg\ mol^{-1} \times 3.2\ g}{128\ g\ mol^{-1} \times 80\ g} \\
&= 1.5\ K
\end{aligned}
$$

Question124

Which of the following solutions exhibits lowest value of boiling point elevation assuming complete dissociation?

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Options:

A. 0.1 m $AlCl_3$

B. 0.01 m $MgCl_2$

C. 1 m KCl

D. 0.5 m NaCl

Answer: B

Solution:

	Solution	Moles of particles in 1 kg solution
(A)	0.1 m $AlCl_3$	0.4
(B)	0.01 m $MgCl_2$	0.03

	Solution	Moles of particles in 1 kg solution
(C)	1 m KCl	2
(D)	0.5 m NaCl	1

0.01 m $\mathrm{MgCl_2}$ solution has minimum number of particles in solution, so it shows the lowest value of boiling point elevation.

Question125

A solution of $8\ \mathrm{g}$ of certain organic compound in $2\ \mathrm{dm}^3$ water develops osmotic pressure $0.6\ \mathrm{atm}$ at $300\ \mathrm{K}$. Calculate the molar mass of compound. [$\mathrm{R} = 0.082$ atm $\mathrm{dm}^3\ \mathrm{K}^{-1}\ \mathrm{mol}^{-1}$]

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Options:

A. $148\ \mathrm{g\ mol}^{-1}$

B. $164\ \mathrm{g\ mol}^{-1}$

C. $172\ \mathrm{g\ mol}^{-1}$

D. $180\ \mathrm{g\ mol}^{-1}$

Answer: B

Solution:

$$\begin{aligned} \pi &= \frac{\mathrm{W_2RT}}{\mathrm{M_2\ V}} \\ \therefore \quad \mathrm{M_2} &= \frac{\mathrm{W_2RT}}{\pi \mathrm{V}} \\ \therefore \quad \mathrm{M_2} &= \frac{8\ \mathrm{g} \times 0.082\mathrm{dm}^3\ \mathrm{atm\ K}^{-1}\ \mathrm{mol}^{-1} \times 300\ \mathrm{K}}{0.6\ \mathrm{atm} \times 2\mathrm{dm}^3} \\ &= 164\ \mathrm{g\ mol}^{-1} \end{aligned}$$

Question126

Which from the following compound solutions in water of equal concentration has electrical conductivity nearly same as distilled water?

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Options:

A. Urea

B. Sodium chloride

C. Sodium hydroxide

D. Acetic acid

Answer: A

Solution:

Urea is nonelectrolyte and hence, it has electrical conductivity nearly same as distilled water. Sodium chloride and sodium hydroxide are strong electrolytes while acetic acid is weak electrolyte.

Question127

A solution of nonvolatile solute is obtained by dissolving $1 \mathrm{~g}$ in $100 \mathrm{~g}$ solvent, decreases its freezing point by $0.3 \mathrm{~K}$. Calculate cryoscopic constant of solvent if molar mass of solute is $60 \mathrm{~g} \mathrm{~mol}^{-1}$.

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Options:

A. $1.0 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

B. $1.4 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

C. $2.4 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

D. $1.8 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

Answer: D

Solution:

$$
\begin{aligned}
& \Delta \mathrm{T}_{\mathrm{f}}=\frac{1000 \mathrm{~K}_{\mathrm{f}} \mathrm{W}_2}{\mathrm{M}_2 \mathrm{~W}_1} \\
\therefore \quad & \mathrm{K}_{\mathrm{f}}=\frac{\Delta \mathrm{T}_{\mathrm{f}} \mathrm{M}_2 \mathrm{~W}_1}{1000 \mathrm{~W}_2} \\
& =\frac{0.3 \mathrm{~K} \times 60 \mathrm{~g} \mathrm{~mol}^{-1} \times 100 \mathrm{~g}}{1000 \times 1 \mathrm{~g}} \\
& =1.8 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}
\end{aligned}
$$

Question128

Which among the following gases exhibits very low solubility in water at room temperature?

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Options:

A. O_2

B. CO_2

C. NH_3

D. HCl

Answer: A

Solution:

Option A, O_2 (Oxygen), is the gas that exhibits very low solubility in water at room temperature among the options provided.

Oxygen is a non-polar molecule and has a very low solubility in water, which is a polar solvent. This low solubility is due to the lack of strong interactions between the oxygen molecules and the water molecules.

Regarding the other options :

- CO_2 (Carbon Dioxide) is more soluble than oxygen due to its ability to react with water to form carbonic acid.
- NH_3 (Ammonia) is highly soluble in water and forms ammonium hydroxide.
- HCl (Hydrochloric Acid) is extremely soluble in water and dissociates completely to form hydrochloric acid.

Question129

$0.2\ \mathrm{M}$ aqueous solution of glucose has osmotic pressure 4.9 atm at $300\ \mathrm{K}$. What is the concentration of glucose if it has osmotic pressure $1.5\ \mathrm{atm}$ at same temperature?

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Options:

A. 0.03 M

B. 0.04 M

C. 0.05 M

D. 0.06 M

Answer: D

Solution:

$\pi=\mathrm{M}\times\mathrm{R}\times\mathrm{T}$

$\therefore \quad \mathrm{M}=\frac{\pi}{\mathrm{RT}}=\frac{1.5}{0.082\times300}=0.06\ \mathrm{M}$

Question130

A solution of nonvolatile solute is obtained by dissolving $3.5\ \mathrm{g}$ in $100\ \mathrm{g}$ solvent has boiling point elevation $0.35\ \mathrm{K}$. Calculate the molar mass of solute.

(Molal elevation constant $= 2.5 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$)

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Options:

A. $270 \mathrm{~g} \mathrm{~mol}^{-1}$

B. $260 \mathrm{~g} \mathrm{~mol}^{-1}$

C. $250 \mathrm{~g} \mathrm{~mol}^{-1}$

D. $240 \mathrm{~g} \mathrm{~mol}^{-1}$

Answer: C

Solution:

$\mathrm{M}_2=\frac{\mathrm{K}_{\mathrm{b}} \times \mathrm{W}_2 \times 1000}{\Delta \mathrm{T}_{\mathrm{b}} \times \mathrm{W}_1}=\frac{2.5 \times 3.5 \times 1000}{0.35 \times 100}=250 \mathrm{~g} \mathrm{~mol}^{-1}$

Question131

Calculate the solubility of a gas in water at $0.8 \mathrm{~atm}$ and $25^{\circ} \mathrm{C}$.

[Henry's law constant is $6.85 \times 10^{-4} \mathrm{~mol} \mathrm{~dm}^{-3} \mathrm{~atm}^{-1}$]

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Options:

A. $2.74 \times 10^{-4} \mathrm{~mol} \mathrm{~dm}^{-3}$

B. $3.94 \times 10^{-4} \mathrm{~mol} \mathrm{~dm}^{-3}$

C. $6.85 \times 10^{-4} \mathrm{~mol} \mathrm{~dm}^{-3}$

D. $5.48 \times 10^{-4} \mathrm{~mol} \mathrm{~dm}^{-3}$

Answer: D

Solution:

The solubility of a gas in a liquid according to Henry's law can be determined using the formula:

$S = k_H \cdot P$

Where:

- S is the solubility of the gas in the liquid (in $\mathrm{mol\ dm^{-3}}$).
- k_H is Henry's law constant (in $\mathrm{mol\ dm^{-3}\ atm^{-1}}$).
- P is the partial pressure of the gas (in atm).

Given the Henry's law constant (k_H) is $6.85 \times 10^{-4}\ \mathrm{mol\ dm^{-3}\ atm^{-1}}$ and the partial pressure (P) of the gas is 0.8 atm, we can calculate the solubility (S) of the gas in water at the given conditions:

$S = (6.85 \times 10^{-4}\ \mathrm{mol\ dm^{-3}\ atm^{-1}}) \cdot (0.8\ \mathrm{atm})$

$S = 5.48 \times 10^{-4}\ \mathrm{mol\ dm^{-3}}$

Therefore, the solubility of the gas in water at 0.8 atm and 25°C is $5.48 \times 10^{-4}\ \mathrm{mol\ dm^{-3}}$. The correct option is:

Option D

$5.48 \times 10^{-4}\ \mathrm{mol\ dm^{-3}}$

Question132

Identify the concentration of the solution from following so that values of ,$\Delta \mathrm{T_f}$ and $\mathrm{K_f}$ are same.

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Options:

A. 1 m

B. 1 M

C. 1 N

D. $\frac{\mathrm{N}}{10}$

Answer: A

Solution:

To identify the correct option, we first need to understand the relationship between the freezing point depression (ΔT_f) and the molal freezing point depression constant (K_f) of the solvent. This relationship is given by the colligative property equation for freezing point depression:

$\Delta T_f = K_f \times m$

where:

- ΔT_f is the freezing point depression,
- K_f is the cryoscopic constant (also known as the molal freezing point depression constant), and
- m is the molality of the solution.

The key to solving this problem is noting that the values for ΔT_f and K_f need to be the same, which means:

$\Delta T_f = K_f \Rightarrow K_f = m$

Thus, the molality of the solution needs to be 1 mol/kg since the only situation where the equality holds is if the molality m is equal to 1.

Now, let's analyze the options:

- Option A: 1 m - This represents a concentration of 1 molal, which means 1 mole of solute per 1 kilogram of solvent.
- Option B: 1 M - This represents a concentration of 1 molar, which is 1 mole of solute per 1 liter of solution. Molarity and molality are not the same and the density of the solution would matter to convert between the two.
- Option C: 1 N - This represents a 1 normal solution, which is related to the equivalent concept of moles of reactive species. Normality can vary based on the equivalent factor of the solute, and is not purely a measure of the number of moles of solute per liter of solution.
- Option D: $\frac{N}{10}$ - This represents one-tenth the normality of the solution, which would be equivalent to a 0.1 N solution.

From the analysis above, we can conclude that the correct option is A, as it provides a molality of 1 which is necessary for ΔT_f to equal K_f when they are numerically the same:

$\Delta T_f = K_f = 1\ \mathrm{m}$

Question133

Calculate the amount of solute dissolved in 160 gram solvent that boils at $85^\circ\mathrm{C}$, the molar mass of solute is $120\ \mathrm{g\ mol^{-1}}$.($K_b$ for solvent $= 2.7^\circ\mathrm{C\ kg\ mol^{-1}}$ and boiling point of solvent $= 76^\circ\mathrm{C}$)

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Options:

A. 42 gram

B. 60 gram

C. 64 gram

D. 50 gram

Answer: C

Solution:

$$\Delta T_b = K_b \cdot m$$

$$9 = 2.7\left(\frac{m}{0.12} \times \frac{1}{0.16}\right)$$

$$m = \frac{9 \times 0.12 \times 0.16}{2.7} = 0.064 \text{ kg} = 64 \text{ g}$$

Question134

The solution containing 3 g urea (molar mass 60) per dm^3 of water and another solution containing 4.5 g of solute A per dm^3 boils at same temperature, then what is molar mass of A ?

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Options:

A. 54 g mol^{-1}

B. 180 g mol^{-1}

C. 120 g mol^{-1}

D. 90 g mol^{-1}

Answer: D

Solution:

To find the molar mass of solute A, we need to use the concept of boiling point elevation and the fact that both solutions boil at the same temperature. This implies that their boiling point elevations are the same, and thus their molalities are equal since the boiling point elevation constant (Kb) and the solvent are the same for both solutions.

The formula for molality (m) is given by:

$$m = \frac{\text{mass of solute (g)}}{\text{molar mass of solute (g/mol)} \times \text{mass of solvent (kg)}}$$

Since the solutions are dilute and the solvent is water, we can assume 1 dm^3 of water (which is 1 kg). Now, we also know that the boiling point elevation per unit molality (Kb) for water is a constant and will not affect our calculations since both solutions have the same change in boiling point.

For the urea solution:

mass of urea $= 3$ g

molar mass of urea $= 60 \text{ g mol}^{-1}$

So, the molality for urea solution is:

$$m_{\text{urea}} = \frac{3 \text{ g}}{60 \text{ g mol}^{-1} \times 1 \text{ kg}} = \frac{3}{60} = 0.05 \text{ mol kg}^{-1}$$

For the unknown solute A solution:

mass of solute A $= 4.5$ g

Let the molar mass of solute A be M_{A} .

So, the molality for solution A is:

$$m_{\text{A}} = \frac{4.5 \text{ g}}{M_{\text{A}} \times 1 \text{ kg}} = \frac{4.5}{M_{\text{A}}} \text{ mol kg}^{-1}$$

Since the boiling point elevations of both solutions are the same, their molalities are equal:

$$0.05 = \frac{4.5}{M_{\text{A}}}$$

Solving for M_{A} :

$$M_{\text{A}} = \frac{4.5}{0.05} = 90 \text{ g mol}^{-1}$$

Therefore, the molar mass of solute A is:

Option D: 90 g mol^{-1}

Question135

Vapour pressure of solution and of pure solvent are P_1 and $P_1{}^0$ respectively. If $\frac{P_1}{P_1^0}$ is 0.15, find the mole fraction of solute.

Options:

A. 0.66

B. 0.85

C. 0.15

D. 0.33

Answer: B

Solution:

We can solve this problem using Raoult's Law. According to Raoult's Law, the vapor pressure of a solution is given by:

$\mathrm{P}_1 = \chi_{\text{solvent}} \cdot \mathrm{P}_1^0$

Where P_1 is the vapor pressure of the solution, P_1^0 is the vapor pressure of the pure solvent, and χ_{solvent} is the mole fraction of the solvent.

Given that $\frac{P_1}{P_1^0} = 0.15$, we can write:

$\chi_{\text{solvent}} = 0.15$

Since the mole fractions of the solute and solvent must add up to 1, we have:

$\chi_{\text{solute}} + \chi_{\text{solvent}} = 1$

Substituting the given value of χ_{solvent}:

$\chi_{\text{solute}} + 0.15 = 1$

Solving for χ_{solute}, we get:

$\chi_{\text{solute}} = 1 - 0.15$

$\chi_{\text{solute}} = 0.85$

So, the mole fraction of the solute is 0.85. Therefore, the correct answer is Option B: 0.85.

Question136

According to Raoult's law mole fraction of solute in solution in given by formula

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Options:

A. $\frac{\Delta \mathrm{P}}{\mathrm{P}_1^0}$

B. $\frac{\mathrm{P}_1^0}{\mathrm{P}_1}$

C. $\frac{\mathrm{P}_1^0}{\Delta \mathrm{P}}$

D. $\frac{\mathrm{P}_1}{\mathrm{P}_1^0}$

Answer: A

Solution:

Raoult's law states that the partial vapor pressure of each volatile component in a solution is directly proportional to its mole fraction. For a solution, Raoult's law for the solvent can be expressed as:

$P_1 = P_1^0 \cdot x_1$

where:

- P_1 is the vapor pressure of the solvent in the solution.
- P_1^0 is the vapor pressure of the pure solvent.
- x_1 is the mole fraction of the solvent in the solution.

The mole fraction of the solute x_2 can be derived from the lowering of vapor pressure, which is expressed as:

$\Delta P = P_1^0 - P_1$

Here, ΔP is the decrease in vapor pressure of the solvent when the solute is added. To find the mole fraction of the solute x_2, we use the relation:

$x_2 = \frac{\Delta P}{P_1^0}$

Therefore, the correct answer is:

Option A

$\frac{\Delta \mathrm{P}}{\mathrm{P}_1^0}$

Question137

What is vapour pressure of solution containing 0.1 mole solute dissolved in 1.8×10^{-2} kg H_20 ? ($P_1^0 = 24$ mm Hg)

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Options:

A. 12.40 mm Hg

B. 18.12 mm Hg

C. 15.72 mm Hg

D. 21.84 mm Hg

Answer: D

Solution:

$n_2 = 0.1$ mol, $P_1^0 = 24$ mmHg
$W_1 = 1.8 \times 10^{-2}$ kg, $H_2O = 18$ g, $M_1 = 18$ g mol^{-1}

$$\therefore n_1 = \frac{W_1}{M_1} = \frac{18 \text{ g}}{18 \text{ g mol}^{-1}} = 1 \text{ mol}$$

Now, $\frac{P_1^0 - P_1}{P_1^0} = \frac{n_2}{n_1 + n_2}$

$$\therefore \quad \frac{24\text{mmHg} - P_1}{24\text{mmHg}} = \frac{0.1}{1 + 0.1} = \frac{0.1}{1.1} = 0.09$$
$$\therefore \quad 24 \text{ mmHg} - P_1 = 2.16 \text{ mmHg}$$
$$\therefore \quad P_1 = 24 \text{ mmHg} - 2.16 \text{ mmHg} \quad = 21.84 \text{ mmHg}$$

Question138

What is vapour pressure of a solution containing 1 mol of a non-volatile solute in 36 g of water? ($P_1^0 = 400$ mmHg)

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Options:

A. 334 mm Hg

B. 267 mm Hg

C. 240 mm Hg

D. 284 mm Hg

Answer: B

Solution:

$\mathrm{P_1^0 = 400\ mmHg, n_2 = 1\ mol, n_1 = \frac{36}{18} = 2\ mol}$

$$\frac{P_1^0 - P_1}{P_1^0} = \frac{n_2}{n_1 + n_2}$$
$$\therefore \frac{400 - P_1}{400} = \frac{1}{3}$$
$$\therefore 400 - P_1 = 133.33$$
$$\therefore 400 - 133.33 = P_1$$
$$\therefore P_1 = 266.67 \text{ mm Hg}$$

Question139

Which of the following solutions behaves nearly as an ideal solution?

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Options:

A. Benzene + toluene

B. Phenol + aniline

C. Chloroform + acetone

D. Ethanol + acetone

Answer: A

Solution:

The solution that behaves nearly as an **ideal solution** is:

A) Benzene + toluene

Reason:
Ideal solutions are formed when the intermolecular interactions between unlike molecules are nearly the same as those between like molecules. Benzene and toluene are chemically similar (both are non-polar aromatic hydrocarbons) and have similar molecular sizes, so they obey **Raoult's law** closely.

Why the others are non-ideal:

- **Phenol + aniline** → strong hydrogen bonding
- **Chloroform + acetone** → specific intermolecular interactions
- **Ethanol + acetone** → hydrogen bonding differences

Correct answer: Benzene + toluene

Question140

The solution containing $6 \mathrm{~g}$ urea (molar mass 60) per dm^3 of water and another solution containing $9 \mathrm{~g}$ of solute A per dm^3 water freezes at same temperature. What is molar mass of A ?

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Options:

A. 90

B. 180

C. 54

D. 120

Answer: A

Solution:

$$\frac{M_{2\,(\text{urea})}}{M_{2(A)}} = \frac{W_{2\,(\text{urea})}}{W_{2(A)}} \quad [\because \text{ solvent is same }]$$

$$\therefore \quad \frac{60}{M_{2(A)}} = \frac{6}{9}$$

$$\therefore \quad M_{2(A)} = \frac{60\times 9}{6} = 90$$

Question141

What is vapour pressure of solution containing 1.8 g glucose in 16.2 g water?

($P_1^0 = 24$ mmHg and Molar mass of glucose $= 180 \text{ g mol}^{-1}$)

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Options:

A. 18.1 mm Hg

B. 15.7 mm Hg

C. 12.4 mm Hg

D. 23.8 mm Hg

Answer: D

Solution:

$W_2 = 1.8 \text{ g}, M_2 = 180 \text{ g mol}^{-1},\ W_1 = 16.2 \text{ g},$
$M_1 = 18 \text{ g mol}^{-1}, P_1^0 = 24 \text{ mmHg}, P_1 = \ ?$

$$\frac{P_1^0 - P_1}{P_1^0} = \frac{W_2 M_1}{M_2\, W_1}$$

$$\therefore \frac{24 \text{ mmHg} - P_1}{24 \text{ mmHg}} = \frac{1.8 \text{ g} \times 18 \text{ g mol}^{-1}}{180 \text{ g mol}^{-1} \times 16.2 \text{ g}}$$

$$\therefore 24 \text{ mmHg} - P_1 = 24 \text{ mmHg}(0.0111)$$

$$\therefore \quad P_1 = 24 \text{ mmHg} - 0.2664 \text{ mmHg}$$

$$\therefore \quad P_1 = 23.73 \text{ mmHg}$$

Question142

What is the freezing point of 1 molal aqueous solution of a non volatile solute?

$(\mathrm{K_f} = 1.86\ \mathrm{K\ kg\ mol^{-1}})$ $(\mathrm{T_f^0}$ **for water** $= 0^\circ\mathrm{C})$

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Options:

A. $-0.93^\circ\mathrm{C}$

B. $-2.43^\circ\mathrm{C}$

C. $-3.72^\circ\mathrm{C}$

D. $-1.86^\circ\mathrm{C}$

Answer: D

Solution:

To calculate the freezing point of a 1 molal aqueous solution of a non-volatile solute, we need to employ the freezing point depression formula:

$\Delta T_f = K_f \cdot m$

where:

- ΔT_f is the freezing point depression.
- K_f is the cryoscopic constant (freezing point depression constant) for the solvent.
- m is the molality of the solution.

Given:

- $K_f = 1.86\,\mathrm{K\,kg\,mol^{-1}}$
- $m = 1\,\mathrm{mol\,kg^{-1}}$ (since the solution is 1 molal)
- T_f^0 (the freezing point of pure water) $= 0^\circ\mathrm{C}$

First, we calculate the freezing point depression:

$\Delta T_f = 1.86\,\mathrm{K\,kg\,mol^{-1}} \cdot 1\,\mathrm{mol\,kg^{-1}} = 1.86\,\mathrm{K}$

This means the freezing point of the solution is lowered by $1.86\,\mathrm{K}$ (or $1.86\,^\circ\mathrm{C}$) compared to the pure solvent.

Therefore, the new freezing point is:

$T_f = T_f^0 - \Delta T_f = 0^\circ\mathrm{C} - 1.86^\circ\mathrm{C} = -1.86^\circ\mathrm{C}$

Hence, the correct answer is:

Option D: $-1.86°C$

Question143

Air is an example of a solution of

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Options:

A. gas in solid

B. liquid in gas

C. gas in liquid

D. gas in gas

Answer: D

Solution:

The correct answer is **Option D: gas in gas**. Here's why:

A solution is a homogeneous mixture where one substance (the solute) is dissolved evenly into another substance (the solvent).

Air is a mixture of gases, primarily nitrogen (N_2) and oxygen (O_2), along with smaller amounts of other gases like carbon dioxide (CO_2) and argon (Ar). Since all components are in the gaseous state, air is classified as a solution of **gas in gas**.

Let's look at why the other options are incorrect:

Option A: gas in solid This describes a situation like hydrogen gas dissolved in a metal, which is not the case with air.

Option B: liquid in gas This describes a situation like water vapor in air, which is a component of air but not the primary definition of air itself.

Option C: gas in liquid This describes a situation like carbon dioxide dissolved in water, which is a different mixture than air.

Question144

5 g sucrose (molar mass = 342) is dissolved in 100 g of solvent, decreases the freezing point by 2.15 K. What is cryoscopic constant of solvent?

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Options:

A. $14.7 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

B. $2.15 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

C. $4.30 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

D. $7.35 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

Answer: A

Solution:

$$
\begin{aligned}
& \mathrm{W}_2=5 \mathrm{~g}, \mathrm{M}_2=342 \mathrm{~g} \mathrm{~mol}^{-1}, \mathrm{~W}_1=100 \mathrm{~g}, \Delta \mathrm{T}_{\mathrm{f}}=2.15 \mathrm{~K}, \mathrm{~K}_{\mathrm{f}}=\ ? \\
& \Delta \mathrm{T}_{\mathrm{f}}=\mathrm{K}_{\mathrm{f}} \frac{1000 \mathrm{~W}_2}{\mathrm{M}_2 \mathrm{~W}_1} \\
& \therefore \mathrm{K}_{\mathrm{f}}=\frac{\Delta \mathrm{T}_{\mathrm{f}} \mathrm{M}_2 \mathrm{~W}_1}{1000 \mathrm{~W}_2} \\
& \therefore \mathrm{K}_{\mathrm{f}}=\frac{2.15 \mathrm{~K} \times 342 \mathrm{~g} \mathrm{~mol}^{-1} \times 100 \mathrm{~g}}{1000 \mathrm{~g} \mathrm{~kg}^{-1} \times 5 \mathrm{~g}} \\
& \therefore \mathrm{K}_{\mathrm{f}}=14.7 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}
\end{aligned}
$$

Question145

What is Henry's law constant if solubility of a gas in water at $298 \mathrm{~K}$ and 1 bar pressure is $7 \times 10^{-4} \mathrm{~mol} \mathrm{~L}^{-1}$?

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Options:

A. 2.0×10^{-5} mol L^{-1} bar^{-1}

B. 7.0×10^{-4} mol $L^{-1} bar^{-1}$

C. 3.5×10^{-3} mol L^{-1} bar^{-1}

D. 3.1×10^{-5} mol L^{-1} bar^{-1}

Answer: B

Solution:

$$\begin{aligned} &\mathrm{P} = 1 \text{ bar}, \mathrm{S} = 7 \times 10^{-4} \text{ mol L}^{-1} \\ &\qquad \mathrm{K_H} = \frac{\mathrm{S}}{\mathrm{P}} = \frac{7 \times 10^{-4} \text{ mol L}^{-1}}{1 \text{ bar}^{-1}} \\ &\qquad\quad = 7 \times 10^{-4} \text{ mol L}^{-1} \text{bar}^{-1} \end{aligned}$$

Question146

Which of the following formulae is used to obtain depression in freezing point?

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Options:

A. $\Delta T_f = \frac{T_f^{\circ}}{T_f}$

B. $\Delta T_f = T_f^{\circ} - T_f$

C. $\Delta \mathrm{T_f} - \mathrm{T_f} = \mathrm{T_f^{\circ}}$

D. $\Delta T_f = \frac{T_f}{T_f^{\circ}}$

Answer: B

Solution:

The correct formula to obtain the depression in freezing point is encapsulated in Option B, which is:

$\Delta T_f = T_f^\circ - T_f$

This equation accurately represents the depression in freezing point, where:

- ΔT_f is the depression in freezing point.
- T_f° is the normal freezing point of the pure solvent.
- T_f is the freezing point of the solution.

The concept underlying this formula arises from colligative properties, which are properties that depend on the number of solute particles in a solution and not on the nature of the solute itself. The depression in freezing point is a colligative property that refers to the lowering of the freezing point of a solvent when a solute is dissolved in it. Therefore, the difference between the normal freezing point of the solvent (T_f°) and the freezing point of the solution (T_f) gives the depression in freezing point (ΔT_f).

Options A, C, and D present incorrect formulae which do not correctly represent the concept of depression in freezing point in terms of the existing conventions and the fundamental principles of physical chemistry.

Question147

The vapour pressure of a solvent decreases by 2.5 mm Hg by adding a solute. What is the mole fraction of solute? (Vapour pressure of pure solvent is 250 mm Hg)

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Options:

A. 0.88

B. 0.01

C. 0.1

D. 0.99

Answer: B

Solution:

To solve this problem, we need to use Raoult's Law, which states that the vapor pressure of a solution is directly proportional to the mole fraction of the solvent. Mathematically, Raoult's Law is expressed as:

$P_{\text{solution}} = x_{\text{solvent}} \cdot P_{\text{solvent}}^{\text{pure}}$

Where:

P_{solution} = vapor pressure of the solution

x_{solvent} = mole fraction of the solvent

$P_{\text{solvent}}^{\text{pure}}$ = vapor pressure of the pure solvent

Given that the vapor pressure of the pure solvent, $P_{\text{solvent}}^{\text{pure}}$, is 250 mm Hg, and the vapor pressure of the solution, P_{solution}, decreases by 2.5 mm Hg, we can write:

$P_{\text{solution}} = P_{\text{solvent}}^{\text{pure}} - \Delta P$

Where:

ΔP = decrease in vapor pressure = 2.5 mm Hg

So:

$P_{\text{solution}} = 250 \text{ mm Hg} - 2.5 \text{ mm Hg} = 247.5 \text{ mm Hg}$

Using Raoult's Law:

$247.5 = x_{\text{solvent}} \cdot 250$

Solve for the mole fraction of the solvent:

$x_{\text{solvent}} = \frac{247.5}{250} = 0.99$

Now, knowing that the sum of the mole fractions of the solute and the solvent must be equal to 1, we can write:

$x_{\text{solute}} = 1 - x_{\text{solvent}}$

So:

$x_{\text{solute}} = 1 - 0.99 = 0.01$

Therefore, the mole fraction of the solute is 0.01, which corresponds to **option B**.

Question148

Molal depression constant for a liquid is $2.77^{\circ}\text{C kg mol}^{-1}$, in Kelvin scale it's value is

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Options:

A. $275.77 \text{ K kg mol}^{-1}$

B. 271.77 $K\ kg\ mol^{-1}$

C. 2.77 $K\ kg\ mol^{-1}$

D. 27.7 $K\ kg\ mol^{-1}$

Answer: C

Solution:

The molal depression constant, also known as the cryoscopic constant, is a property of a solvent which describes the freezing point depression (lowering) per molal concentration of solute particles. This constant is usually given in degrees Celsius per kilogram per mole.

In this particular problem, you are given the molal depression constant as $2.77^{\circ}C\ kg\ mol^{-1}$. The question asks for its value in the Kelvin scale.

Since the Kelvin and Celsius scales have the same size of division (i.e., the difference of 1 Kelvin is equal to the difference of 1 degree Celsius), converting a temperature change or difference from degrees Celsius to Kelvin simply retains its numerical value. The only modification needed is in the symbol notation (from $^{\circ}C$ to K).

Therefore, the value of the molal depression constant in the Kelvin scale is the same numerical value as in the Celsius scale:

So, the molal depression constant in the Kelvin scale is:

$2.77\ K\ kg\ mol^{-1}$

Hence, the correct answer is:

Option C

$2.77\ K\ kg\ mol^{-1}$

Question149

If 6 g of solute dissolved in 100 g of water lowers the freezing point by 0.93 K. What is molar mass of solute? ($K_f = 1.86\ K\ kg\ mol^{-1}$)

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Options:

A. 120 $g\ mol^{-1}$

B. 60 g mol^{-1}

C. 90 g mol^{-1}

D. 180 g mol^{-1}

Answer: A

Solution:

To determine the molar mass of the solute, we can use the formula for freezing point depression:

$\Delta T_{\mathrm{f}} = K_{\mathrm{f}} \cdot m$

Where:

- ΔT_{f} is the freezing point depression (0.93 K in this case)
- K_{f} is the cryoscopic constant (1.86 K kg mol^{-1})
- m is the molality of the solution

The molality (m) is given by:

$m = \frac{n}{kg \text{ of solvent}}$

Where n is the number of moles of solute. The number of moles of solute can be calculated using the mass of the solute and its molar mass (M):

$n = \frac{mass\ of\ solute}{molar\ mass\ (M)} = \frac{6\ g}{M}$

So, the molality (m) will be:

$m = \frac{\frac{6\ g}{M}}{0.1\ kg} = \frac{60}{M}$

We can now substitute this back into the formula for freezing point depression:

$\Delta T_{\mathrm{f}} = K_{\mathrm{f}} \cdot \frac{60}{M}$

Given that the freezing point depression ΔT_{f} is 0.93 K, we have:

$0.93 = 1.86 \cdot \frac{60}{M}$

Solving for M, the molar mass of the solute:

$M = 1.86 \cdot \frac{60}{0.93}$

$M = 120\ g\ mol^{-1}$

Therefore, the molar mass of the solute is **120 g mol**$^{-1}$.

The correct answer is Option A: 120 g mol^{-1}.

Question150

What is vapour pressure of a solution when $2\ \mathrm{mol}$ of a non-volatile solute are dissolved in $20\ \mathrm{mol}$ of water? $\left(P_1^0 = 32\ \mathrm{mmHg}\right)$

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Options:

A. 29.1 mm Hg

B. 12 mm Hg

C. 6 mm Hg

D. 9 mm Hg

Answer: A

Solution:

$$P_1^0 = 32\ \mathrm{mm\ Hg}, n_1 = 20\ \mathrm{mol}, n_2 = 2\ \mathrm{mol}$$

$$\therefore x_1 = \frac{n_1}{n_1 + n_2} = \frac{20}{20+2} = 0.909$$

$$\begin{aligned} P_1 &= P_1^0 x_1 = 32\ \mathrm{mm\ Hg} \times 0.909 \\ &= 29.1\ \mathrm{mm\ Hg} \end{aligned}$$

Question151

In which of the following salts, the solubility increases appreciably with increase in temperature?

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Options:

A. KBr

B. NaBr

C. NaCl

D. KCl

Answer: A

Solution:

Solubilities of NaBr, NaCl and KCl changes slightly with temperature.

Question152

What is the boiling point of 0.5 molal aqueous solution of sucrose if 0.1 molal aqueous solution of glucose boils at 100.16°C?

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Options:

A. 100.32°C

B. 100.80°C

C. 100.16°C

D. 100.62°C

Answer: B

Solution:

For glucose solution, $\mathrm{m} = 0.1\ \mathrm{m},\ \mathrm{T_b} = 100.16^\circ\mathrm{C}$

$$\begin{aligned} \therefore \Delta \mathrm{T_b} &= \mathrm{T_b} - \mathrm{T_b}^\circ \\ &= 100.16 - 100 = 0.16^\circ\mathrm{C} \end{aligned}$$

$$\mathrm{K_b} = \frac{\Delta \mathrm{T_b}}{\mathrm{m}} = \frac{0.16^\circ\mathrm{C}}{0.1\ \mathrm{m}} = 1.6^\circ\mathrm{C/m}$$

For sucrose solution, $\mathrm{m} = 0.5\ \mathrm{m},\ \mathrm{K_b} = 1.6^\circ\mathrm{C/m}$

$$\begin{aligned} \Delta \mathrm{T_b} &= \mathrm{K_b} \times \mathrm{m} \\ &= 1.6^\circ\mathrm{C/m} \times 0.5\ \mathrm{m} \quad = 0.80^\circ\mathrm{C} \\ \Delta \mathrm{T_b} &= \mathrm{T_b} - \mathrm{T_b^\circ} \\ \mathrm{T_b} &= \Delta \mathrm{T_b} + \mathrm{T_b^\circ} \\ &= 0.80 + 100 \quad = 100.80^\circ\mathrm{C} \end{aligned}$$

Question153

Which of the following solutions shows positive deviation from Raoult's law?

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Options:

A. Ethanol + Acetone

B. Chloroform + Acetone

C. Benzene + Toluene

D. Phenol + Aniline

Answer: A

Solution:

Ethanol + Acetone show positive deviation from Raoult's law while other solution show negative deviations.

Question154

What is vapour pressure of a solution containing 0.1 mol of non-volatile solute dissolved in 16.2 g of water ? ($P_1^0 = 32$ mm Hg)

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Options:

A. 21.6 mm Hg

B. 28.8 mm Hg

C. 15.7 mm Hg

D. 18.1 mm Hg

Answer: B

Solution:

$$
\begin{aligned}
& n_2 = 0.1 \text{ mol}, n_1 = \frac{16.2}{18} = 0.9 \text{ mol} \\
& x_1 = \frac{n_1}{n_1 + n_2} = \frac{0.9}{0.1 + 0.9} = 0.9 \\
& P_1 = P_1^0 x_1 = 32 \text{ mmHg} \times 0.9 \\
& \qquad\qquad = 28.8 \text{ mm Hg}
\end{aligned}
$$

Question155

What is boiling point of a decimolal aqueous solution of glucose if molal elevation constant for water is 0.52°C kg mol^{-1} ?

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Options:

A. 101.52°C

B. 99.95°C

C. 99.48°C

D. 100.052°C

Answer: D

Solution:

$$
\begin{aligned}
& m = 0.1 \text{ m}, \ K_b = 0.52^{\circ}\text{C kg mol}^{-1} \\
& \Delta T_b = K_b \times m = 0.52^{\circ}\text{C kg mol}^{-1} \times 0.1 \text{ mol kg}^{-1} = 0.052^{\circ}\text{C} \\
& \Delta T_b = T_b - T_b^0 \quad \therefore T_b = \Delta T_b + T_b^0 \\
& \therefore T_b = 0.052 + 100 = 100.052^{\circ}\text{C}
\end{aligned}
$$

Question156

What is cryoscopic constant of water if $5 \mathrm{~g}$ of glucose in $100 \mathrm{~g}$ of water has depression in freezing point $2.15 \mathrm{~K}$? (Molar mass of glucose $=180$)

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Options:

A. $7.74 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

B. $0.52 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

C. $1.32 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

D. $3.86 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

Answer: A

Solution:

$$
\begin{aligned}
& \mathrm{W}_2=5 \mathrm{~g}, \quad \mathrm{~W}_1=100 \mathrm{~g} \\
& \mathrm{M}_2=180 \mathrm{~g}, \quad \Delta \mathrm{T}_{\mathrm{f}}=2.15 \mathrm{~K} \\
& \mathrm{~K}_{\mathrm{f}}=? \\
& \Delta \mathrm{T}_{\mathrm{f}}=\mathrm{K}_{\mathrm{f}} \frac{1000 \mathrm{~W}_2}{\mathrm{M}_2 \mathrm{~W}_1} \\
& \therefore \mathrm{K}_{\mathrm{f}}=\frac{\Delta \mathrm{T}_{\mathrm{f}} \mathrm{M}_2 \mathrm{~W}_1}{1000 \mathrm{~W}_2}=\frac{2.15 \times 180 \times 100}{1000 \times 5} \\
& \quad=7.74 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}
\end{aligned}
$$

Question157

Which of the following statements is correct for boiling point of a liquid?

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Options:

A. Temperature at which a liquid boils at any pressure.

B. Temperature at which solid is in equilibrium with its liquid.

C. Temperature at which vapour pressure equals the applied pressure.

D. Temperature at which applied pressure is greater than vapour pressure of liquid.

Answer: C

Solution:

The correct statement for the **boiling point of a liquid** is:

C) Temperature at which vapour pressure equals the applied pressure.

Explanation:
A liquid boils when its **vapour pressure becomes equal to the external (applied) pressure** . At this point, bubbles of vapour can form throughout the liquid and rise to the surface.

Correct answer: C

Question158

Henry's law constant for $\mathrm{CH_3Br}$ is $0.16 \mathrm{~mol~L}^{-1} \mathrm{bar}^{-1}$ at $298 \mathrm{~K}$. What pressure is required to have solubility of $0.08 \mathrm{~mol~L}^{-1}$?

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Options:

A. 0.24 bar

B. 1.6 bar

C. 0.5 bar

D. 4.0 bar

Answer: C

Solution:

$\mathrm{K_H}=0.16 \mathrm{~mol~L}^{-1} \mathrm{bar}^{-1}, \mathrm{~S}=0.08 \mathrm{~mol~L}^{-1}$

According to Henry's law,

$S = K_H P$

$$\therefore P = \frac{S}{K_H} = \frac{0.08 \text{mol} L^{-1}}{0.16 \text{ mol } L^{-1} \text{bar}^{-1}} = 0.5 \text{ bar}$$

Question159

Which of the following solutions does not flow in either direction, when separated by semipermeable membrane? (Molar mass : glucose = 180, urea = 60)

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Options:

A. 18 g urea dm^{-3} and 18 g glucose dm^{-3}

B. 6 g urea dm^{-3} and 36 g glucose dm^{-3}

C. 6 g urea dm^{-3} and 24 g glucose dm^{-3}

D. 12 g urea dm^{-3} and 36 g glucose dm^{-3}

Answer: D

Solution:

$n_{urea} = \frac{12}{60} = 0.2$ and $n_{glucose} = \frac{36}{180} = 0.2$

Now, $n_{urea} = n_{glucose}$

$\therefore \pi_{urea} = \pi_{glucose}$

Hence, if these solutions are separated by a semipermeable membrane, there is no flow of solvent in either direction.

Question160

A solution of 6 g of solute in 100 g of water boils at 100.52°C. The molal elevation constant of water is 0.52 K kg mol^{-1}. What is molar mass of solute?

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Options:

A. 60 g mol^{-1}

B. 120 g mol^{-1}

C. 90 g mol^{-1}

D. 180 g mol^{-1}

Answer: A

Solution:

$W_2 = 6$ g, $W_1 = 100$ g, $K_b = 0.52$ K kg mol^{-1}, $T_b = 100.52°C$

$\therefore \Delta T_b = T_b - T_b^o = (100.52 + 273) - (100 + 273) = 0.52\ K$

Now,

$$M_2 = \frac{1000 \cdot K_b \cdot W_2}{\Delta T_b \cdot W_1} = \frac{1000\ g\ kg^{-1} \times 0.52\ K\ kg\ mol^{-1} \times 6\ g}{0.52\ K \times 100\ g}$$

$\therefore \quad M_2 = 60\ g\ mol^{-1}$

Question161

If vapour pressure of pure solvent and solution are 240 and 216 mm Hg respectively then mole fraction of solvent in solution is

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Options:

A. 0.9

B. 0.1

C. 0.6

D. 0.4

Answer: A

Solution:

$$P_1 = 216 \text{ mm Hg}, \mathrm{P}_1^0 = 240 \text{ mmHg}$$
$$\Delta P = P_1^0 - P_1 = 240 - 216 = 24 \text{ mm Hg}$$

Now, $\frac{\Delta \mathrm{P}}{\mathrm{P}_1^0} = \mathrm{x}_2$

$\therefore \quad x_2 = \frac{24 \text{ mm Hg}}{240 \text{ mm Hg}} = 0.1$

Hence, mole fraction of solvent, $\mathrm{x}_1 = 1 - \mathrm{x}_2 = 1 - 0.1 = 0.9$

Question162

What is the unit of viscosity?

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Options:

A. Nm^{-1}

B. Nsm^{-2}

C. kgs^{-1}

D. kgs^2

Answer: B

Solution:

Unit of viscosity is Nsm^{-2}

$$\text{Viscosity } (\mu) = \frac{\text{Force } \times \text{ Time}}{(\text{ Length })^2} = \frac{\text{Newton } \times \text{ Second}}{(\text{ Metre })^2}$$
$$= \mathrm{Nsm}^{-2}$$

Question163

The vapour pressure of solvent decreases by 10 mm Hg , if mole fraction of non-volatile solute is 0.2 . Calculate vapour pressure of solvent.

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Options:

A. 70 mm of Hg

B. 40 mm of Hg

C. 50 mm of Hg

D. 60 mm of Hg

Answer: C

Solution:

Given,

Lowering of vapour pressure

$\left(p_A^{\circ}-p_A\right)=10 \text{ mm Hg}$

Mole fraction of non-volatile solution $\left(\chi_A\right)=0.2$

Mole fraction of solution,

$$\begin{aligned} & \chi_A=\frac{p_A^{\circ}-p_A}{p_A} \\ & 0.2=\frac{100 \text{ mmHg}}{p_A} \\ \Rightarrow \quad & p_A=\frac{10}{0.2} \text{ mmHg}=50 \text{ mm of Hg} \end{aligned}$$

Question164

Which of the following $0-10 \mathrm{~m}$ aqueous solutions will have maximum ΔT_f value?

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Options:

A. $C_{12}H_{22}O_{11}$

B. KI

C. $Al_2(SO_4)_3$

D. $NH_2 - CO - NH_2$

Answer: C

Solution:

On the give 0.10 m aqueous solution, $Al_2(SO_4)_3$ will have maximum ΔT_1 value.The expression for the freezing point of the solution is

$\Delta T_f = iK_f \cdot m$

van't Hoff factor (ℓ), is highest for $Al_2(SO_4)_3$, which is 5 . Hence, it will give maximum ΔT_f value.

Question165

If boiling point of urea solution is $100.18^\circ C$ and K_b for water is $0.512\ K\ kg\ mol^{-1}$, molality of solution is (Boiling point of water $= 100^\circ C$)

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Options:

A. $0.6\ mol\ kg^{-1}$

B. $0.25\ mol\ kg^{-1}$

C. $0.35\ mol\ kg^{-1}$

D. $0.45\ mol\ kg^{-1}$

Answer: C

Solution:

Given,

Boiling point of solution $= 100.18^{\circ}\mathrm{C} \Rightarrow 373.18\ \mathrm{K}$

$K_b = 0.512\ \mathrm{K\ kg\ mol^{-1}}$

Boiling point of water $= 100^{\circ}\mathrm{C} \Rightarrow 373\ \mathrm{K}$

$\Delta T_{\text{Bolling}} \quad = 373.18\ \mathrm{K} - 373\ \mathrm{K} = 0.18\ \mathrm{K}$

We know that,

$$\begin{aligned} \Delta T &= K_b \times \text{ molarity} \\ \Rightarrow \quad \text{Molarity } &= \frac{\Delta T}{K_b} = \frac{0.18\ \mathrm{K}}{0.512\ \mathrm{mol^{-1}\ kg\ K}} \\ &= 0.35\ \mathrm{mol\ kg^{-1}} \end{aligned}$$

Question166

What is osmotic pressure of a semi molar solution at $27^{\circ}\mathrm{C}$? $(R = 0.082)$

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Options:

A. 20.5 atm

B. 4.96 atm

C. 2.46 atm

D. 12.3 atm

Answer: D

Solution:

Given,

Concentration (C) = Semi-molar solution that means concentration $\frac{1}{2}m$ of solution is half.

$T = 27^{\circ}\mathrm{C} = 300\ \mathrm{K}$
$R = 0.082$

We know that, osmotic pressure $(\pi) = \mathrm{CRT}$

$\pi = \frac{1}{2} \times 0.082 \times 300 = 12.3\ \mathrm{atm}$

Question167

Solutions A, B, C and D are respectively 0.2 M urea, $0.10\ \mathrm{M\ NaCl}, 0.05\ \mathrm{M\ BaCl_2}$ and 0.05 M $\mathrm{AlCl_3}$. All solutions are isotonic with each other except

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Options:

A. B

B. A

C. D

D. C

Answer: D

Solution:

Isotonic solution have the same osmotic pressure.

The osmotic pressure of 0.2 M urea (A).

$$\begin{aligned} \pi &= iMRT \\ &= 1 \times 0.2RT \\ &= 0.2RT \end{aligned}$$

The osmotic pressure of $0.10 \mathrm{M NaCl}(B)$.

$$\begin{aligned} \pi &= iMRT \\ &= 2 \times 0.1RT \\ &= 0.2RT \end{aligned}$$

The osmotic pressure of $0.05 \mathrm{M BaCl_2}(C)$.

$$\begin{aligned}\pi &= IMRT \\ &= 3 \times 0.05 RT \\ &= 0.15 RT\end{aligned}$$

The osmotic pressure of $0.05 \mathrm{AlCl}_3 (D)$.

$$\begin{aligned}\pi &= iMRT \\ &= 4 \times 0.05 RT \\ &= 0.2 RT\end{aligned}$$

Hence, all solution are isotonic with each other except solution C.

Question168

A solution has an osmotic pressure of 'x' kPa at $300 \mathrm{~K}$ having 1 mole of solute in $10.5 \mathrm{~m}^3$ of solution. If it's osmotic pressure is reduced to $\left(\frac{1}{10}\right)$th of it's initial value, what is the new volume of solution?

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Options:

A. $11.0 \mathrm{~m}^3$

B. $105 \mathrm{~m}^3$

C. $30 \mathrm{~m}^3$

D. $110 \mathrm{~m}^3$

Answer: B

Solution:

The given values are

$\pi_1 =' x' \mathrm{kPa}, \pi_2 = \frac{x}{10} \mathrm{kPa}, T_1 = 300 \mathrm{~K}, n_1 = 1$ mole

Using formula,

$$\begin{aligned} \pi_1 V &= x_1 RT \\ x \times 10.5 &= 1 \times 8.314 \times 300 \\ \Rightarrow \quad x &= \frac{8.314 \times 300}{10.5}\end{aligned}$$

If osmotic pressure is reduced $\frac{1}{10}$ th then, value $\pi_2 = \frac{x}{10}$ kPa

$$\begin{aligned} \pi_2 V &= nRT \\ \frac{x}{10} \times V &= n \times R \times T \\ \frac{8.314 \times 300}{10.5 \times 10} \times V &= 1 \times 8.314 \times 300 \\ \Rightarrow \quad V &= 10.5 \times 10 = 105 \mathrm{~m}^3 \end{aligned}$$

Question169

If a centimolal aqueous solution of $K_3[Fe(CN)_6]$ has degree of dissociation 0.78. What is the value of van't Hoff factor?

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Options:

A. 1.2

B. 4.0

C. 3.34

D. 2.5

Answer: C

Solution:

Given,

Degree of dissociation $(\alpha) = 0.78$

$K_3\left[Fe(CN)_6\right] \longrightarrow 3 K^{+} + \left[Fe(CN)_6\right]^{3-}$

That means, total moles formed after dissociation $(n) = 4$

We know that,

$$\begin{aligned} \text{van't Hoff factor (i)} &= \alpha n + (1 - \alpha) \\ &= 0.78 \times 4 + (1 - 0.78) \\ &= 3.122 + 0.22 = 3.34 \end{aligned}$$

Question170

The elevation in boiling point of 0.25 molal aqueous solution of a substance is $\left(K_o = 0.52 \text{ K kg mol}^{-1}\right)$

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Options:

A. 0.15 K

B. 0.50 K

C. 0.13 K

D. 2.08 K

Answer: C

Solution:

Given,

$m = 0.25$
$K_b = 0.52 \text{ K kg mol}^{-1}$

The elevation in boiling point (ΔT_b) of 0.25 molal aqueous solution of a substance is

$\Delta T_b = K_b \times m = 0.52 \times 0.25 = 0.13 \text{ K}$

Question171

Which among the following salts, solubility decreases with increase in temperature?

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Options:

A. Na_2SO_4

B. NaBr

C. NaCl

D. KCl

Answer: A

Solution:

For Na_2SO_4 salt, solubility decreases with increase in temperature because reaction of Na_2SO_4 with water is an exothermic reaction, i.e. $\Delta_{\text{sol}} \quad H < 0$. As a result solubility decreases. For NaBr, NaCl and KCl , the dissolution process is endothermic.

Question172

18 gram glucose (Molar mass $= 180$) is dissolved in 100 ml of water at 300 K . If $R = 3\ 0.0821\ \mathrm{L{-}atm\ mol^{-1}\ K^{-1}}$ what is the osmotic pressure of solution?

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Options:

A. 2.463 atm

B. 24.63 atm

C. 8.21 atm

D. 0.821 atm

Answer: B

Solution:

The various quantities known to us are as follows:

$R = 0.0821\ \mathrm{L - atm\ mol^{-1}\ K^{-1}}$
$\mathrm{w}_2 = 18$ gram

Molar mass $(M_2) = 180$

$T = 300$ K
$V = 100$ mL

To calculate the osmotic pressure of solution, we use the following formula,

$$\pi = \frac{W_2 RT}{M_2 V} = \frac{18 \times 0.0821 \times 300 \times 1000}{180 \times 100}$$
$$= 24.63 \text{ atm}$$

Question173

Relationship between van't Hoff's factor (i) and degree of dissociation (α) is

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Options:

A. $i = \frac{\alpha-1}{n'-1}$

B. $i = \frac{\alpha-1}{1-n'}$

C. $\alpha = \frac{1-i}{n'-1}$

D. $\alpha = \frac{i-1}{n'-1}$

Answer: D

Solution:

Relationship between vant Hoff factor (i) and degree of dissociation (α) is given by

$\alpha = \frac{1-i}{n'-1}$

where, n is the number of ions formed after dissociation.

The relationship can be obtained as follows:

For the reaction, $A \rightleftharpoons n'B$

Initially 1 mole 0

After dissociation $(1-\alpha)$ mole $r'\alpha$

Total number of moles present in the solution

$= (1-\alpha) + n'\alpha = 1 + (n'-1)\alpha$

van't Hoff factor, $i = 1 + (n'-1), \alpha > 1$ if $n' \geq 2$

$\therefore \quad \alpha = \frac{i-1}{n'-1}$

Question174

9 gram anhydrous oxalic acid (mol. $\mathrm{wt.} = 90$) was dissolved in 9.9 moles of water. If vapour pressure of pure water is p_1° the vapour pressure of solution is

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Options:

A. $0.99 p_1^\circ$

B. $0.1 p_1^\circ$

C. $0.99 p_1^\circ$

D. $1.1 p_1^\circ$

Answer: A

Solution:

The total vapour pressure of a solution in this case only depends on vapour pressure of water as anhydrous oxalic acid is a non-volatile compound.

$\therefore$ Vapour pressure of solution = vapour pressure of water (p_w)

According to Raoult's law

$p_w = x_w p_w^\circ$

Number of moles of oxalic acid $= \frac{9}{90} = 0.1$ moles

$\therefore \quad x_w = \frac{9.9}{9.9+0.1} = 0.99$
$\Rightarrow \quad p_s = p_w = 0.99 \times p_1^\circ$

Question175

Which of the following sets of solutions of urea (mol . mass 60 g mol^{-1}) and sucrose (mol . mass 342 g mol^{-1}) is isotonic?

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Options:

A. 9.1 gL^{-1} urea and 6.0 gL^{-1} sucrose

B. 3.0 gL^{-1} urea and 3.0 gL^{-1} sucrose

C. 6.0 gL^{-1} urea and 9.0 gL^{-1} sucrose

D. 3.0 gL^{-1} urea and 17.1 gL^{-1} sucrose

Answer: D

Solution:

Formula for osmotic pressure, $\pi = CRT$

Considering the set given in option (d), i.e. 3.0 gL^{-1} urea and 17.1 gL^{-1} sucrose.

Given, molecular mass of urea 60 g mol^{-1} and molecular mass of sucrose 342 g mol^{-1}.

For urea,

conc. $C = \frac{3}{60} = \frac{1}{20}$

Osmotic pressure $\pi_1 = \frac{1}{20} \times R \times T$

For sucrose conc. $C = \frac{17.1}{342} = \frac{1}{20}$

$\therefore$ Osmotic pressure, $\pi_2 = \frac{1}{20} RT$

Thus, set of solution of urea and sucrose given in option (d) is isotonic.

Question176

Which of the following sets of components form homogeneous mixture?

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Options:

A. Phenol + Water

B. Sugar + Benzene

C. Silver chloride + Water

D. Ethyl alcohol + Water

Answer: A

Solution:

Phenol + water and ethyl alcohol + water both forms homogeneous mixture but homogeneousily of phenol + water is more as compared to ethyl alcohol + water ($\because$ larger alkyl group).Thus, the correct option is ' a '.

Question177

Calculate van't Hoff-factor for 0.2 m aqueous solution of KCl which freezes at $-0.680^{\circ}\mathrm{C}$. $\left(K_f=1.86\ \mathrm{K\ kg\ mol}^{-1}\right)$

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Options:

A. 3.72

B. 1.83

C. 6.8

D. 1.86

Answer: B

Solution:

Given, Molality $= 0.2$ m

$$\begin{aligned} K_f &= 1.86 \text{ K kg mol} \\ \Delta T_f &= 0.680^{\circ}\mathrm{C} = 0.680 \text{ K} \end{aligned}$$

As, we know that

$$\begin{aligned} \Delta T_f &= i \times K_f \times m \\ 0.680 &= i \times 1.86 \times 0.2 \\ \frac{0.680}{1.86 \times 0.2} &= i \\ 1.83 &= i \end{aligned}$$

Question178

' K is Henry's constant and has the unit

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Options:

A. $\text{atm mol}^{-1}\text{dm}^3$

B. $\text{mol}^{-1}\text{dm}^3 \text{ atm}^{-1}$

C. atm mol dm^{-3}

D. $\text{mol dm}^{-3} \text{ atm}^{-1}$

Answer: D

Solution:

According to Henry's "the partial pressure of the gas in vapour phase (p) is proportional to the mole fraction of the gas (χ) in the solution". It is expressed as

$p \propto \chi$ or $p = K_H X$

where, $K_{\mathrm{H}} =$ Henry's law constant

Its unit is M |atm or $\text{moldm}^{-3} \text{ atm}^{-1}$.
