Liquid Solution

Question1

At $T(\mathrm{~K})$, the vapour pressure of water is $x \mathrm{kPa}$. What is the vapour pressure (in kPa) of 1 molal solution containing non-volatile solute?

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Options:

A.

$1.018 x$

B.

$0.8 x$

C.

$0.972 x$

D.

$0.982 x$

Answer: D

Solution:

Moles of water $n_{\text {water }} \quad=\frac{1000 \mathrm{~g}}{18 \mathrm{~g} / \mathrm{mol}}$

$=55.56 \mathrm{~mol}$

Mole of solute $=n_{\text {solute }} \quad=1 \mathrm{~mol}$

Mole fraction of water

$$\chi_{\text{water}} = \frac{n_{\text{water}}}{n_{\text{water}} + n_{\text{solute}}}$$

$$\chi_{\text{water}} = \frac{55.56}{55.56 + 1} \approx 0.9823$$

Vapour pressure of solution,

$$\begin{aligned} p_{\text{solution}} &= \chi_{\text{Water}} \times p_{\text{pure water}} \\ &= 0.982 \times xkP_a = 0.982x \end{aligned}$$

Question2

Elements X and Y form two non-volatile compounds (XY and XY_3). When 10 g of XY is dissolved in 50 g of ethanol, the depression in freezing point (ΔT_f) was 5.333 K . When 10 g of XY_3 is dissolved in 50 g of ethanol, the (was ΔT_f)2.2857 K. What are the atomic weights of X and Y respectively?

$$\left(K_f = 2 \text{ K kg mol}^{-1}\right)$$

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Options:

A.

$50u, 50u$

B.

$25u, 25u$

C.

$75u, 100u$

D.

$25u, 50u$

Answer: D

Solution:

Molality of XY and XY_3

$$m_{XY}=\frac{\Delta T_{\mathrm{PXY}}}{k_f}=\frac{5.333}{2}=2.6665\ \mathrm{mol/kg}$$

$$m_{XYy_3}=\frac{\Delta T_{f(XY_3)}}{k_f}=\frac{2.2857}{2}$$

$$=1.14285\ \mathrm{mol/kg}$$

Mass of solvent $=50\ \mathrm{g}=0.05\ \mathrm{kg}$

Moles of XY : $n_{XY}=m_{XY}\times$ mass of solvent

$=0.133325\ \mathrm{mol}$

Moles of XY_3 :

$n_{XY_3}=1.14285\times 0.05=0.05714\ \mathrm{mol}$

Molar mass of $XY=\frac{10}{0.133325}=75\ \mathrm{g/mol}$

Molar mass of $XY_3=\frac{10}{0.0571425}$

$=175\ \mathrm{g/mol}$

Let A_X be atomic weight of X and A_Y be atomic weight of Y.

$$A_X+A_Y=75 \quad ..(i)$$
$$A_X+3A_Y=175 \quad ..(ii)$$

Solving Eqs. (i) and (ii),

$A_Y=25\ \mathrm{g/mol}, A_Y=50\ \mathrm{g/mol}$

or 25 u and 50 u

Question3

What is the approximate molality of $10\%(w/w)$ aqueous glucose solution?

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

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Options:

A.

0.31 m

B.

0.62 m

C.

0.93 m

D.

1.24 m

Answer: B

Solution:

A $10\%(w/w)$ solution $=10$ g0f glucose in 100 g solution.

mass of solute $=10$ g, mass of solvent $=100-10=90$ g

$$\begin{aligned} \text{Moles of glucose } & =\frac{\text{Mass of glucose}}{\text{Molar mass}} \\ & =\frac{10}{180}=0.0556 \text{ mol} \\ \text{Molality } & =\frac{\text{No. of moles of solute}}{\text{Mass of water (kg)}} \\ & =\frac{0.0556}{0.090} \\ & =0.62 \text{ m} \end{aligned}$$

Question4

The van't Hoff factor for 0.5 m aqueous $\mathrm{CH_2FCOOH}$ solution is 1.075 . What is the experimentally observed ΔT_f (in K) for this solution?

($K_f=1.86 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$)

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Options:

A.

1.156

B.

1.075

C.

1.0

D.

0.95

Answer: C

Solution:

Given:

$i = 1.075, \quad m = 0.5\,\text{mol kg}^{-1}, \quad K_f = 1.86\,\text{K·kg·mol}^{-1}$

We are asked to find the **experimentally observed freezing point depression**, ΔT_f.

Formula

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

Substitute the values

$\Delta T_f = 1.075 \times 1.86 \times 0.5$

Perform the calculation

$1.86 \times 0.5 = 0.93$

$0.93 \times 1.075 = 0.99975 \approx 1.00$

✅ **Therefore,**

$\boxed{\Delta T_f = 1.0\,\text{K}}$

Correct Option:

Option C – 1.0

Question5

An aqueous solution containing 0.2 g of a non volatile solute ' A ' in 21.5 g of water freezes at 272.814 K . If the freezing point of water is 273.16 K , the molar mass (in gmol^{-1}) of solute A is

$$\left[\mathrm{K}_f\left(\mathrm{H}_2\mathrm{O}\right)=1.86\ \mathrm{K\ kg\ mol}^{-1}\right]$$

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Options:

A.

80

B.

75

C.

100

D.

50

Answer: D

Solution:

Step 1: Find the decrease in freezing point

The difference between the normal freezing point of water and the solution's freezing point is called the depression in freezing point. The calculation is:

$\Delta T_f=T_f^0-T_f=273.16-272.814$

$\Delta T_f=0.346\ \mathrm{K}$

Step 2: Use the freezing point depression formula

The formula is: $\Delta T_f=K_f\cdot m$ where m is molality.

$m=\frac{\Delta T_f}{K_f}=\frac{0.346}{1.86}=0.186\ \mathrm{mol/kg}$

Step 3: Calculate the moles of solute

Molality tells us moles of solute in 1 kg of solvent (water). Here, amount of water is $21.5\ \mathrm{g}=0.0215$ kg.

Moles of solute $= 0.186 \times 0.0215 = 0.00400$ mol

Step 4: Calculate the molar mass

Molar mass $= \frac{\text{mass of solute}}{\text{moles of solute}}$

Molar mass $= \frac{0.2}{0.00400} = 50 \text{ g/mol}$

Question6

At $T(\text{ K})$, the vapour pressure of x molal aqueous solution containing a non-volatile solute is $12.078 \mathrm{kPa}_{\mathrm{d}}$, The vapour pressure of pure water at $T(\text{ K})$ is 12.3 kPa . What is the value of x ?

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Options:

A.

10

B.

1.018

C.

0.1018

D.

0.018

Answer: B

Solution:

Mole fraction of solvent

$$\begin{aligned} \chi_{\text{solute}} &= \frac{p_0 - p_s}{p_0} \\ \Rightarrow \quad \chi_{\text{Solute}} &= \frac{12.3 - 12.078}{12.3} = 0.01805 \end{aligned}$$

Moles of water in 1000 g of water

$= \frac{1000}{18} = 55.56 \text{ mol}$

For dilute solution,

$\chi_{\text{Solute}} = \frac{n_{\text{solute}}}{n_{\text{solute}} + n_{\text{water}}} \approx \frac{n_{\text{solute}}}{n_{\text{water}}}$

Since molality x is moles of solute per 1 kg of solvent

$x = n_{\text{solute}} \Rightarrow x_{\text{solute}} = \frac{x}{55.56}$

$x = 0.01805 \times 55.56 \approx 1.018$

Question7

In aqueous glucose solution, the mole fraction of wate is 40 times to mole fraction of glucose. What is the weight percentage (w/w) of glucose in the solution?

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Options:

A.

40

B.

30

C.

20

D.

10

Answer: C

Solution:

Mole fraction of binary solution,

$$\begin{aligned} \chi_{\text{glucose}} &+ \chi_{\text{water}} &= 1 \\ \chi_{\text{glucose}} &+ 40 \times \chi_{\text{glucose}} &= 1 \\ \chi_{\text{glucose}} &= \frac{1}{41} \end{aligned}$$

Assume 1 mole of total solution Moles of glucose

$(n_{\text{glucose}}) = \chi_{\text{glucose}} \times 1 = \frac{1}{41}\ \mathrm{mol}$

Moles of water $(n_{\text{water}}) = \frac{40}{41}\ \mathrm{mol}$

Mass of glucose (M_{glucose})

$$\begin{aligned} &= n_{\text{glucose}} \times \text{Molar mass} \\ &= \frac{1}{41} \times 180 = \frac{180}{41} \end{aligned}$$

Mass of water $(m_{\text{water}}) = n_{\text{water}} \times$ Molar mass

$= \frac{40}{41} \times 18 = \frac{720}{41}$

$w/w\%$ of glucose

$$\frac{\text{Mass of glucose}}{\text{Total mass}} \times 100$$

$$= \text{Total mass} = \frac{180}{41} + \frac{720}{41} = \frac{900}{41}\ \mathrm{g}$$

Weight \% of glucose

$= \frac{\frac{180}{41}}{\frac{900}{41}} \times 100\% = 20\%$

Question8

Benzoic acid molecules undergo dimerisation in benzene. 2.44 g of benzoic acid when dissolved in 30 g of benzene caused depression in freezing point of 2 K . What is the percentage of association of it?

(Given $K_f\left(\mathrm{C}_6\mathrm{H}_6\right) = 5\ \mathrm{K\ kg\ mol}^{-1}$; molar mass of benzoic acid $= 122\ \mathrm{g\ mol}^{-1}$)

Options:

A.

80

B.

70

C.

60

D.

90

Answer: A

Solution:

$$\text{Moles of benzoic acid } = \frac{2.44}{122}$$
$$= 0.02 \text{ mol}$$

$$\text{Mass of benzene } = \frac{30}{1000 \text{ g/kg}} = 0.03 \text{ kg}$$

Molality, $m = \frac{0.02}{0.03} = 0.6667 \text{ mol/kg}$

Depression in freezing point,

$\Delta T_f = iK_f \cdot m$
$2 = i \times 5 \times 0.6667 = 0.6$

Using van't Hoff formula,

$0.6 = 1 - \alpha + \frac{\alpha}{2} \Rightarrow \alpha = 0.8$

Percentage of association

$= 0.8 \times 100 = 80\%$

Question9

A solution of urea in water has a boiling point or $100.18^{\circ}\mathrm{C}$. What is the freezing point of the same solution, if K_f and K_b of water are 1.86 and 0.52 K kg mol^{-1}.

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Options:

A.

$-0.34^{\circ}\mathrm{C}$

B.

$-022^{\circ}\mathrm{C}$

C.

$-0.64^{\circ}\mathrm{C}$

D.

$-0.32^{\circ}\mathrm{C}$

Answer: C

Solution:

Elevation in boiling point,

$\Delta T_b=100.18^{\circ}\mathrm{C}-100^{\circ}\mathrm{C}=0.18^{\circ}\mathrm{C}$

Formula for elevation in boiling point is

$$\begin{aligned} \Delta T_b &= K_b \cdot m \\ m &= \frac{0.18}{0.52}=0.346\ \mathrm{mol/kg} \end{aligned}$$

Depression in freezing point is,

$\Delta T_f=K_f \cdot m=1.86 \times 0.346=0.643\ \mathrm{K}$

Freezing point of sol

$0^{\circ}\mathrm{C}-0.643^{\circ}\mathrm{C}=-0.64^{\circ}\mathrm{C}$

Question10

At $T(\mathrm{~K})$, the vapour pressure of pure benzene and toluene are 75 and 22 mm Hg respectively. 23.4 g of benzene and 64.4 g of toluene are mixed to form an ideal solution. If the vapours are in equilibrium with the liquid mixture, the mole fraction of toluene in vapour phase (At.wt. of $\mathrm{C}=12, \mathrm{H}=1$)

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Options:

A.

0.406

B.

0.594

C.

0.539

D.

0.461

Answer: A

Solution:

Moles of benzene

$= n_B = \frac{23.4}{78} = 0.3$ mol

Moles of toluene $= n_{\text{rolu}} \quad = \frac{64.4}{92} = 0.7$ mol Moles fraction, $\chi_{\text{Tolu}} \quad = \frac{n_{\text{Tolu}}}{n_{\text{Total}}} = \frac{0.7}{1} = 0.7$;

$\chi_B = 1 - 0.7 = 0.3$

Partial pressure of toluene

$$\begin{aligned} & p_{\text{Tolu}} \quad = \chi_{\text{Total}} \quad \cdot p_{\text{Tolu}}^{\circ} \\ \Rightarrow \quad & 0.7 \times 22 = 15.4 \mathrm{mmHg} \end{aligned}$$

Partial pressure of benzene

$$\begin{aligned} & p_{\text{Ben}} \quad = \chi_B \cdot p_B \\ \Rightarrow \quad & 0.3 \times 75 = 225 \mathrm{mmHg} \end{aligned}$$

Total vapour, pressure

$= 22.5 + 15.4 = 37.9 \mathrm{mmHg}$

Mole fraction of toluene in vapour

$$Y_{\text{toluene}} = \frac{p_{\text{Toluene}}}{p_{\text{Total}}} = \frac{15.4}{37.9} = 0.406$$

Question11

1.95 g of non-volatile and non-electrolyte solute dissolved in 100 g of benzene lowered the freezing point of it by 0.64 K .

The molar mass of the solute (in gmol^{-1})

$$\left(K_f\left(\mathrm{C}_6 \mathrm{H}_6\right)=5.12 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}\right)$$

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Options:

A.

240

B.

156

C.

165

D.

265

Answer: B

Solution:

Depression in freezing point is given by

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

$$\Rightarrow \quad m = \frac{\Delta T_f}{K_f} = \frac{0.64}{5.12} = 0.125 \text{ mol/kg}$$

Mole of solute = molality × mass of solvent

$$= 0.125 \times \frac{100}{1000} = 0.0125 \text{ mol}$$

$$\text{Molar mass } = \frac{\text{Mass of solute}}{\text{Moles of solute}} = \frac{1.95}{0.0125} = 156 \text{ g/mol}$$

Question12

At $298 \text{ K}, 0.714$ moles of liquid A is dissolved in 5.555 moles of liquid B. The vapour pressure of the resultant solution is 475 torr. The vapour pressure of pure liquid A at the same temperature is 280.7 torr. What is the vapour pressure of pure liquid B in torr?

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Options:

A.

486

B.

550

C.

514

D.

500

Answer: D

Solution:

Given data

- Moles of $A = 0.714\,\text{mol}$
- Moles of $B = 5.555\,\text{mol}$
- $P_{\text{solution}} = 475\,\text{torr}$
- $P_A^0 = 280.7\,\text{torr}$
- $P_B^0 = ?$

Step 1. Mole fractions

$x_A = \frac{0.714}{0.714+5.555} = \frac{0.714}{6.269} = 0.1138$

$x_B = 1 - x_A = 0.8862$

Step 2. Raoult's law

For an ideal solution:

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

We know everything except P_B^0.

So:

$475 = (0.1138)(280.7) + (0.8862)P_B^0$

Step 3. Solve for P_B^0

Compute first term:

$0.1138 \times 280.7 = 31.93$

Thus

$475 = 31.93 + 0.8862P_B^0$

$0.8862P_B^0 = 475 - 31.93 = 443.07$

$P_B^0 = \frac{443.07}{0.8862} = 500.1$

✅ **Answer:** $P_B^0 = 500\,\text{torr}$

Correct Option: D (500 torr)

Question13

The mole fractions of glucose and water in aqueous glucose solution are 0.0244 and 0.9756 respectively. What is the weight percentage (w/w) of glucose in this solution?

Options:

A.

40

B.

25

C.

20

D.

10

Answer: C

Solution:

Let's assume we have 1 mole of the total solution.

Mole fraction of glucose is 0.0244. This means there are 0.0244 moles of glucose in the solution.

The mass of glucose is calculated by multiplying the number of moles by the molar mass of glucose ($180.16 \mathrm{~g} / \mathrm{mol}$):

$0.0244 \times 180.16=4.3959 \mathrm{~g}$

The mole fraction of water is 0.9756. This means there are 0.9756 moles of water in the solution.

The mass of water is found by multiplying the number of moles by the molar mass of water ($18.01 \mathrm{~g} / \mathrm{mol}$):

$0.9756 \times 18.01=17.570 \mathrm{~g}$

The total mass of the solution is just the mass of glucose plus the mass of water:

$4.3959+17.570=21.9659 \mathrm{~g}$

To find the weight percent (w / w) of glucose, divide the mass of glucose by the total mass and multiply by 100:

$\frac{4.3959}{21.9659} \times 100=20 \%$

Question14

At $T(\mathrm{~K})$, the vapour pressure of an aqueous solution of a non-volatile solute, whose mole fraction is 0.02 is found to be 34.65 mm Hg . What is the vapour pressure (in mm Hg) of pure water at the same temperature?

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Options:

A.

35.70

B.

35.36

C.

35.00

D.

34.30

Answer: B

Solution:

Step 1: Find the mole fraction of water.

The solution has a mole fraction of solute = 0.02.

Mole fraction of water = 1 minus mole fraction of solute.

So, $\chi_{\text{water}} = 1 - 0.02 = 0.98$

Step 2: Use Raoult's Law to relate vapour pressures.

Raoult's law says: Vapour pressure of solution = Vapour pressure of pure water × Mole fraction of water.

$p_{\text{solution}} = p_{\text{water}} \times \chi_{\text{water}}$

We know $p_{\text{solution}} = 34.65$ mm Hg and $\chi_{\text{water}} = 0.98$.

So, $34.65 = p_{\text{water}} \times 0.98$

Step 3: Solve for the vapour pressure of pure water.

Divide both sides by 0.98:

$p_{\text{water}} = \frac{34.65}{0.98} = 35.36 \text{ mm Hg}$

Question15

A centi molar solution of acetic acid is 50% dissociated at $27^{\circ}\mathrm{C}$. The osmotic pressure of the solution (in atm) is $\left(R = 0.083 \text{ L atm } \mathrm{K}^{-1} \text{ mol}^{-1}\right)$

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Options:

A.

0.37

B.

3.7

C.

0.037

D.

0.73

Answer: A

Solution:

Given, $T = 27^{\circ}\mathrm{C}$,

$$\begin{aligned} R &= 0.083 \text{ L at mK}^{-1} \text{ mol}^{-1} \\ \text{dissociation } \% &= 50\% \\ \pi &= ? \end{aligned}$$

The dissociation of acetic acid

$$\begin{aligned}&\mathrm{CH_3COOH} \rightleftharpoons \mathrm{CH_3COO^-} + \mathrm{H^+}\\ &\qquad\qquad n = 2\end{aligned}$$

Degree of dissociation $= 50\%$ or 0.5

$i = 1 + 0.5(2-1) = 1.5$

Osmotic pressure is given by

$$\begin{aligned}\pi &= iCRT\\ &= 1.5 \times 0.01 \times 0.023 \times 300 = 0.37 \text{ atm}\end{aligned}$$

Question16

At 300 K vapour pressure of a pure liquid. ' A ' is 70 mm Hg . It forms an ideal solution with another liquid ' B '. The mole fraction of B in the solution is 0.2 and total vapour pressure of solution is 84 mm Hg at same temperature. What is the vapour pressure (in mm) of pure liquid B at 300 K ?

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Options:

A.

140

B.

70

C.

280

D.

560

Answer: A

Solution:

Given data

- $P_A^0 = 70$ mm Hg
- Ideal solution → obeys **Raoult's law**
- $x_B = 0.2 \Rightarrow x_A = 1 - 0.2 = 0.8$
- Total vapour pressure of solution $P_{\text{total}} = 84$ mm Hg
- Temperature = 300 K (constant)

We are asked to find P_B^0 (vapour pressure of pure liquid B).

Step 1: Apply Raoult's Law

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

Substitute known values:

$84 = (0.8)(70) + (0.2)(P_B^0)$

Step 2: Simplify

$84 = 56 + 0.2 P_B^0$

$84 - 56 = 0.2 P_B^0$

$28 = 0.2 P_B^0$

Step 3: Solve for P_B^0

$P_B^0 = \frac{28}{0.2} = 140$ mm Hg

✅ **Answer:** $P_B^0 = 140$ mm Hg

Correct option: A (140)

Question17

248 g of ethylene glycol $(C_2H_6O_2)$ is added to 200 g of water to prepare antifreeze. What is the molality of resultant solution?

$(\mathrm{C} = 12\mathrm{u}; \mathrm{H} = 1\mathrm{u}; \mathrm{O} = 16\mathrm{u})$

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Options:

A.

5 m

B.

10 m

C.

20 m

D.

40 m

Answer: C

Solution:

Moles (n) of ethylene glycol,

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

$$\Rightarrow \quad n = \frac{248}{62} = 4 \text{ mol}$$

$$\text{Mass of water (in kg)} = 0.2 \text{ kg}$$

$$\text{Molality} = \frac{\text{Moles of solute}}{\text{Mass (in kg) of solvent}}$$

$$= \frac{4 \text{ mol}}{0.2} = 20 \text{ mol/kg}$$

Question18

A solution containing 7.5 g of urea (molar mass $= 60 \text{ g mol}^{-1}$) in 1 kg of water freezes at the same temperature as another solution containing 15 g of solute X, in the same amount of water. The molar mass of X $\left(\text{g mol}^{-1}\right)$ is

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Options:

A.

60

B.

180

C.

120

D.

240

Answer: C

Solution:

Moles of urea

$$= \frac{\text{Mass}}{\text{Molar mass}} = \frac{7.5}{60} = 0.125 \text{ mol}$$

$$\begin{aligned} \text{Molality of urea } &= \frac{\text{Moles of urea}}{\text{Mass of water in kg}} \\ &= \frac{0.125}{1} = 0.125 \text{ mol/kg} \end{aligned}$$

Since both solution freezes at the same temperature, their molalities are equal.

$$\begin{aligned} \text{Molality of } x = & \text{ Molality of urea} \\ & = 0.125 \text{ mol/kg} \end{aligned}$$

Moles of $x =$ Molality $\times$ mass of water

$= 0.125 \times 1$

$\Rightarrow \quad 0.125 \text{ mol}$

$$\begin{aligned} \text{Molar mass of } x &= \frac{\text{Mass}}{\text{Moles}} = \frac{15}{0.125} \\ &= 120 \text{ g/mol} \end{aligned}$$

Question19

At 300 K , the vapour prssures of A and B liquids are 500 and 400 mm Hg respectively. Equal moles of A and B are mixed to form an

ideal solution. The mole fraction of A and B in vapour state is respectively

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Options:

A. $0.5, 0.5$

B. $0.666, 0.333$

C. $0.444, 0.555$

D. $0.555, 0.444$

Answer: D

Solution:

Given the vapor pressures of liquids A and B at 300 K are $p_A^\circ = 500\,\text{mmHg}$ and $p_B^\circ = 400\,\text{mmHg}$, respectively, and knowing that equal moles of A and B are mixed to form an ideal solution, we proceed as follows:

Since equal moles of A and B are mixed:

$X_A = X_B = 0.5$

Using Raoult's law, we calculate the partial pressures:

For A:

$p_A = X_A \times p_A^\circ = 0.5 \times 500 = 250\,\text{mmHg}$

For B:

$p_B = X_B \times p_B^\circ = 0.5 \times 400 = 200\,\text{mmHg}$

Total pressure of the solution:

$\text{Total pressure} = p_A + p_B = 250 + 200 = 450\,\text{mmHg}$

Mole fraction of A in the vapor phase:

$Y_A = \frac{p_A}{\text{Total pressure}} = \frac{250}{450} = 0.555$

Mole fraction of B in the vapor phase:

$Y_B = \frac{p_B}{\text{Total pressure}} = \frac{200}{450} = 0.444$

Question20

Two statements are given below

Statements I : Liquids A and B form a non-ideal solution with positive deviation. The interactions between A and B are weaker than $A-A$ and $B-B$ interactions.

Statements II : For an ideal solution $\Delta_{\text{mix}} \quad H=0 ; \Delta_{\text{mix}} \quad V=0$

The correct answer is

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Options:

A. Both Statement I and Statement II are correct.

B. Both Statement I and Statement II are not correct.

C. Statement I is correct but Statement II is not correct.

D. Statement I is not correct and Statement II is correct.

Answer: A

Solution:

Both statements are correct.

Statement I: Positive deviation from Raoult's law means A–B interactions are weaker than A–A and B–B, so this is true.

Statement II: An ideal solution has no heat or volume change on mixing, i.e.

$\Delta_{\text{mix}} H=0, \quad \Delta_{\text{mix}} V=0$,

so this is also true.

Answer: Option A.

Question21

At 290 K , a vessel (I) contains equal moles of three liquids (A, B, C) . The boiling points of A, B and C are $350\ \mathrm{K}, 373\ \mathrm{K}$ and 308 K respectively. Vessel (I) is heated to 300 K and vapours were collected into vessel (II). Identify the correct statements. (Assume vessel (I) contains liquids and vapours and vessel (II)contains only vapours)

(I) Vessel - I is rich in liquid B

(II) Vessel - II is rich in vapour of C

(III) The vapour pressures of A, B, C in vessel (I) at 290 K follows the order $C > A > B$

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Options:

A. I, II, III

B. I, II only

C. I, III only

D. II, III only

Answer: A

Solution:

At 290 K, vessel (I) contains equal moles of liquids A, B, and C. Their boiling points are :

A: 350 K

B: 373 K

C: 308 K

Since a lower boiling point indicates a higher volatility (and hence a higher vapor pressure at a given temperature), at 290 K the order of vapor pressures will be:

$p_C > p_A > p_B$

This confirms statement (III).

Now, vessel (I) is heated to 300 K. At this higher temperature:

Liquid C (boiling point 308 K) becomes much more volatile.

Liquid A (boiling point 350 K) is moderately volatile.

Liquid B (boiling point 373 K) is the least volatile.

As a result, the more volatile liquids (C and A) evaporate preferentially. This means:

The remaining liquid in vessel (I) will be enriched in B because it evaporates very little. So, statement (I) "Vessel - I is rich in liquid B" is correct.

The vapours collected in vessel (II) will be richer in C, since C evaporates the most. So, statement (II) "Vessel - II is rich in vapour of C" is also correct.

In summary, all three statements are correct:

Statement (I): True.

Statement (II): True.

Statement (III): True.

Thus, the correct answer is Option A: I, II, III.

Question22

At $300\ \mathrm{K}$, $6\ \mathrm{g}$ of urea was dissolved in 500 mL of water. What is the osmotic pressure (in atm) of resultant solution?

$$\left(R = 0.082\ \mathrm{L}^{\mathrm{atm}\ \ \mathrm{K}-1}\ \mathrm{mol}^{-1}\right)$$

$$(\mathrm{C} = 12; \mathrm{N} = 14; \mathrm{O} = 16; \mathrm{H} = 1)$$

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Options:

A. 0.492

B. 2.46

C. 4.92

D. 49.2

Answer: C

Solution:

To calculate the osmotic pressure of a solution where 6 grams of urea is dissolved in 500 mL of water at 300 K, we'll use the following formula for osmotic pressure:

$\pi = \frac{n}{V} \times R \times T$

Given:

$R = 0.082\,\text{L atm K}^{-1}\text{mol}^{-1}$

Temperature, $T = 300\,\text{K}$

Volume of solution, $V = 500\,\text{mL} = 0.5\,\text{L}$

First, we need to determine the number of moles of urea (n). The molar mass of urea (NH_2CONH_2) is calculated as follows:

Molar mass of urea $= 1 \times 12 + 4 \times 1 + 2 \times 14 + 1 \times 16 = 60\,\text{g/mol}$

Therefore, the number of moles, n, is:

$n = \frac{\text{Mass}}{\text{Molar mass}} = \frac{6\text{ g}}{60\text{ g/mol}} = 0.1\,\text{mol}$

Substitute back into the osmotic pressure formula:

$\pi = \frac{n}{V} \times R \times T = \frac{0.1\text{ mol}}{0.5\text{ L}} \times 0.082\,\text{L atm K}^{-1}\text{mol}^{-1} \times 300\,\text{K}$

Calculate the values:

$\pi = \frac{0.1}{0.5} \times 0.082 \times 300 = 4.92\,\text{atm}$

Thus, the osmotic pressure of the solution is 4.92 atm.

Question23

In water, which of the following gases has the highest Henry's law constant at 293 K ?

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Options:

A. N_2

B. O_2

C. He

D. H_2

Answer: C

Solution:

In the context of Henry's law, the relationship between the Henry's constant (K_{H}) and solubility (S) can be understood as follows:

$K_{\mathrm{H}} \propto \frac{1}{S}$

This equation indicates that the Henry's constant is inversely proportional to the solubility of a gas in a liquid, given a specific external pressure and temperature. In simple terms, if a gas is less soluble in water, its Henry's constant will be higher.

For helium (He), it is known to have the lowest solubility among the gases considered here. Therefore, helium will have the highest Henry's constant at 293 K.

Question24

Benzoic acid undergoes dimerisation in benzene. x g of benzoic acid (molar mass $122 \mathrm{~g} \mathrm{~mol}^{-1}$) is dissolved in 49 g of benzene. The depression in freezing point is 1.12 K . If degree of association of acid is 88%. What is the value of x ? $\left(K_f\right.$ for benzene $\left.=4.9 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}\right)$

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Options:

A. 2.44

B. 1.22

C. 3.66

D. 4.88

Answer: A

Solution:

Given:

Weight of benzene (solvent), $W_A = 49\,\mathrm{g} = 49 \times 10^{-3}\,\mathrm{kg}$

Depression in freezing point, $\Delta T_f = 1.12\,\mathrm{K}$

Cryoscopic constant for benzene, $K_f = 4.9\,\mathrm{K\,kg/mol}$

Molar mass of benzoic acid, $M_B = 122\,\mathrm{g/mol}$

Degree of association, $\alpha = 0.88$

Formula for depression in freezing point:

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

Calculating the van't Hoff factor (i) for association:

For dimerization (association of 2 molecules into 1 dimer), the expression for i is:

$$i = 1 - \left(1 - \frac{1}{n}\right) \times \alpha$$

Substituting the values:

$$i = 1 - \left(1 - \frac{1}{2}\right) \times 0.88 = 0.56$$

Solving for W_B (weight of solute, x):

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

Substituting the known values:

$$W_B = \frac{1.12 \times 49 \times 10^{-3} \times 122}{0.56 \times 4.9}$$

Calculating:

$$W_B = 2.44\,\mathrm{g}$$

Therefore, the weight of benzoic acid, x, is $2.44\,\mathrm{g}$.

Question25

At $T(\mathrm{K})$ two liquids A and B form an ideal solution. The vapour pressures of pure liquid A and B at that temperature are 400 and 600 mm Hg respectively, If the mole fraction of liquid B is 0.3 in the mixture, the mole fractions of A and B in vapour phase respectively are

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Options:

A. $0.391, 0.609$

B. $0.509, 0.491$

C. $0.609, 0.391$

D. $0.491, 0.509$

Answer: C

Solution:

Given Data

Liquid A

Pure vapor pressure, $p_A^{\circ} = 400$ mm Hg

Mole fraction, $X_A = 0.7$ (since $X_B = 0.3$)

Liquid B

Pure vapor pressure, $p_B^{\circ} = 600$ mm Hg

Mole fraction, $X_B = 0.3$

Total Vapor Pressure Calculation

According to Dalton's law of partial pressures, the total vapor pressure (p_{total}) of the solution is given by:

$$p_{\text{total}} = X_A \cdot p_A^{\circ} + X_B \cdot p_B^{\circ}$$

Substituting the values:

$$p_{\text{total}} = 0.7 \times 400 + 0.3 \times 600 = 280 + 180 = 460 \text{ mm Hg}$$

Mole Fraction in Vapor Phase

Mole Fraction of A, Y_A

The mole fraction of A in the vapor phase is calculated as follows:

$$Y_A = \frac{p_A}{p_{\text{total}}} = \frac{0.7 \times 400}{460} = \frac{280}{460} = 0.609$$

Mole Fraction of B, Y_B

The mole fraction of B in the vapor phase can be found by:

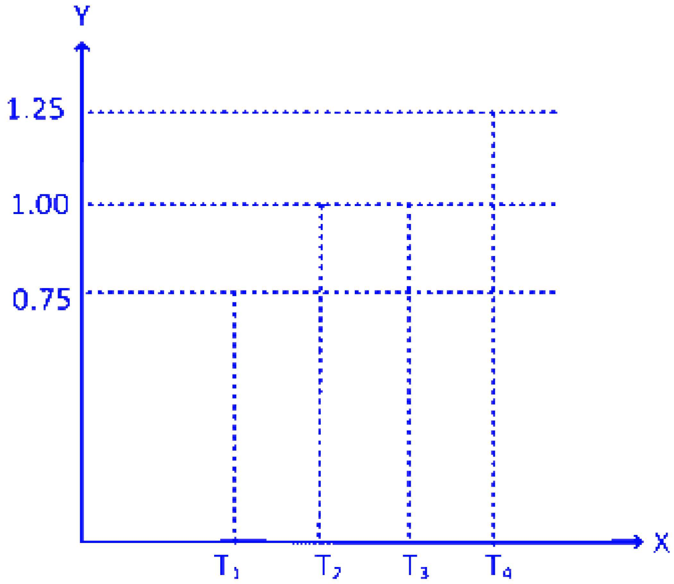
$$Y_B = 1 - Y_A = 1 - 0.609 = 0.391$$

Therefore, the mole fractions of A and B in the vapor phase are 0.609 and 0.391, respectively.

Question26

The following graph is obtained for vapour pressure (in atm) (on Y-axis) and T (in K) (on X-axis) for aqueous urea solution and water. What is the boiling point (in K) of urea solution?

(Atmospheric pressure $=1$ atm)



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Options:

A. T_1

B. T_2

C. T_3

D. T_4

Answer: C

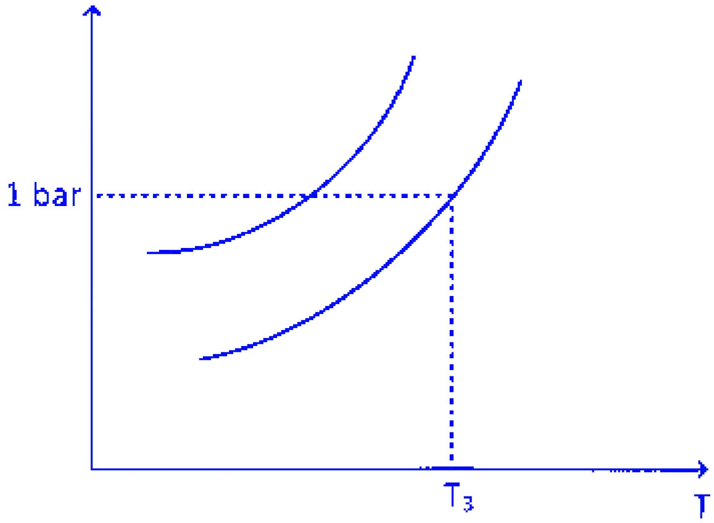
Solution:

Water boils at 100°C or 373.13 K becomes at this temperature, the vapour pressure of waer is equal to the atmospheric pressure.

However, vapour pressure of liquid decreases in presence of non-volatile solution like urea.

In order that the vapour pressure reaches the atomospheric pressure, the temperature is increase beyond normal boiling point of pure water.

Hence, urea solution temperature (boiling) is always higher than pure solvent (water).



At 1 bar, boiling point of urea is T_3.

Question27

Given below are two statements.

Statement I : Liquids A and B form a non-ideal solution with negative deviation. The interactions between A and B are weaker than $A-A$ and $B-B$ interactions.

Statement II : In reverse osmosis, the applied pressure must be higher than the osmotic pressure of solution

The correct answer is

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Options:

A. Both Statement I and Statement II are correct.

B. Both Statement I and Statement II are not correct.

C. Statement I is correct but Statement II is not correct.

D. Statement I is not correct but Statement II is correct

Answer: D

Solution:

Statement I is incorrect, while Statement II is correct. The corrected version of Statement I should state that for a non-ideal solution with negative deviation, the interactions between different components ($A-B$) are stronger than the interactions within the same components ($A-A$ and $B-B$). This results in a negative change in volume and enthalpy upon mixing ($\Delta V_{\text{mixing}} = -\text{ve}$ and $\Delta H_{\text{mixing}} = -\text{ve}$).

Statement II asserts that in reverse osmosis, the applied pressure must exceed the osmotic pressure of the solution for the process to occur, which is correct.

Question28

A solution was prepared by dissolving 0.1 mole of a non- volatile solute in 0.9 moles of water. What is the relative lowering of vapour pressure of solution?

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Options:

A. 0.9

B. 0.5

C. 0.1

D. 0.05

Answer: C

Solution:

According to Raoult's law, the relative lowering of vapor pressure is equal to the mole fraction of the solute.

The formula for calculating the relative lowering of vapor pressure is:

$\frac{p_0-p}{p_0} = \frac{x_A}{x_A+x_B}$

Where:

x_A is the moles of the solute.

x_B is the moles of the solvent.

For this solution, we have:

$x_A = 0.1$ moles (solute)

$x_B = 0.9$ moles (solvent)

Plugging these values into the formula, we get:

$\frac{p_0 - p}{p_0} = \frac{0.1}{0.1+0.9} = 0.1$

Therefore, the relative lowering of vapor pressure is 0.1.

Question29

A non- volatile solute is dissolved in water. The ΔT_{b} of resultant solution is 0.052 K . What is the freezing point of the solution (in K)?

(K_b of water $= 0.52 \mathrm{~Kkgmol}^{-1}$,

K_f of water $= 1.86 \mathrm{~Kkgmol}^{-1}$,

freezing point of water $= 273 \mathrm{~K}$)

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Options:

A. 272.628

B. 273.186

C. 273.000

D. 272.814

Answer: D

Solution:

To determine the freezing point of the solution, we use the concept of colligative properties. The elevation in boiling point (ΔT_b) is given as 0.052 K, with the elevation constant (K_b) for water at 0.52 K kg/mol. The freezing point depression constant (K_f) is 1.86 K kg/mol, and the original freezing point of water is 273 K.

First, calculate the molality (m) from the elevation in boiling point:

$$\Delta T_b = m \times K_b \quad \Rightarrow \quad m = \frac{\Delta T_b}{K_b}$$

$$m = \frac{0.052\ \mathrm{K}}{0.52\ \mathrm{K\ \ kg/mol}} = 0.1$$

Next, use this molality to find the depression in freezing point (ΔT_f):

$$\Delta T_f = m \times K_f = 0.1 \times 1.86 = 0.186\,\mathrm{K}$$

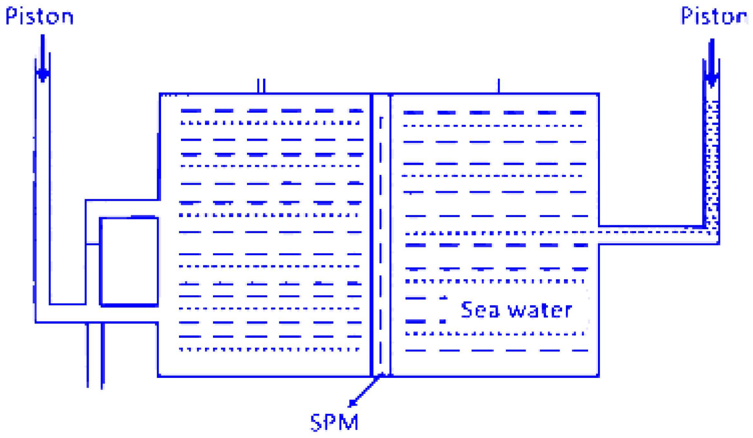
Finally, calculate the freezing point of the solution (T_f):

$$T_f = 273\,\mathrm{K} - \Delta T_f = 273\,\mathrm{K} - 0.186\,\mathrm{K} = 272.814\,\mathrm{K}$$

Thus, the freezing point of the solution is 272.814 K.

Question30

The osmotic pressure of sea water is 1.05 atm . Four experiments were carried as shown in table. In which of the following experiments, pure water can be obtained in part-II of vessel.



Table

Expt. No	Pressure applied in part-1 of vessel	Pressure applied in part-II of vessel
I.	10 atm	–
II.	–	10 atm
III.	15 atm	–
IV.	–	15 atm

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Options:

A. I, III only

B. II, IV only

C. I, II, III, IV

D. IV only

Answer: A

Solution:

From experiment I and III pure water can be obtained in vessel II. This is due to phenomenon of reverse osmosis.

Reverse osmosis is the process of movement of solvent through a SPM from the solution to the pure solvent by applying excess pressure ($>\pi$) on solution side (in this case sea water.)

Question31

The ΔT_b value for 0.01 m KCl solution is 0.01 K . What is the van't Hoff factor? $\left(K_b \text{ for water } = 0.52 \text{ K kg mol}^{-1}\right)$

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Options:

A. 1.92

B. 1.72

C. 0.96

D. 0.86

Answer: A

Solution:

To find the van't Hoff factor for the KCl solution, we start with the given data:

Elevation in boiling point, $\Delta T_b = 0.01\,\mathrm{K}$

Molal boiling point elevation constant for water, $K_b = 0.52\,\mathrm{K\,kg\,mol^{-1}}$

Molality of the solution, $m = 0.01\,\mathrm{m}$

The formula for the elevation in boiling point is:

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

Where i is the van't Hoff factor.

Rearranging the formula to solve for i, we get:

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

Substituting the given values:

$i = \frac{0.01}{0.52 \times 0.01}$

Calculating the result:

$i = 1.92$

Question32

200 g of $20\% \frac{\mathrm{w}}{\mathrm{w}}$ urea solution is mixed with 400 g of $40\% \frac{w}{w}$ urea solution. What is the weight percentage $\left(\frac{w}{w}\%\right)$ of resultant solution?

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Options:

A. 30.33

B. 33.33

C. 36.33

D. 28.33

Answer: B

Solution:

Let's break down the problem step by step:

Calculate the mass of urea in the first solution:

The first solution weighs 200 g and is 20% urea.

Mass of urea = $200\text{ g} \times 0.20 = 40\text{ g}$.

Calculate the mass of urea in the second solution:

The second solution weighs 400 g and is 40% urea.

Mass of urea = $400\text{ g} \times 0.40 = 160\text{ g}$.

Determine the total mass of urea and the total mass of the solution:

Total urea = $40\text{ g} + 160\text{ g} = 200\text{ g}$.

Total solution mass = $200\text{ g} + 400\text{ g} = 600\text{ g}$.

Find the weight percentage of urea in the resultant solution:

Weight percentage =

$$\frac{\text{Total mass of urea}}{\text{Total mass of solution}} \times 100\% = \frac{200\text{ g}}{600\text{ g}} \times 100\% \approx 33.33\%.$$

Thus, the weight percentage of urea in the final solution is approximately 33.33%, which corresponds to Option B.

Question33

At 293 K , methane gas was passed into 1 L . of water. The partial pressure of methane is 1 bar. The number of moles of methane dissolved in 1 L water is (K_{H} of methane $= 0.4\text{ K bar}$)

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Options:

A. 1.38

B. 138×10^{-2}

C. 138×10^{-3}

D. 1.38×10^{-1}

Answer: D

Solution:

To determine the number of moles of methane dissolved in 1 L of water at 293 K with a partial pressure of 1 bar, we employ Henry's law and the concept of mole fraction.

Given:

Partial pressure of methane, $p_0 = 1$ bar

Henry's constant for methane, $K_{\mathrm{H}} = 0.4\,\mathrm{K\,bar}$

Step 1: Apply Henry's Law

Henry's law can be expressed as:

$p_0 = K_{\mathrm{H}}\chi$

Solving for the mole fraction (χ) of methane, we have:

$1 = 0.4 \times 10^3 \times \chi$

$\chi = 2.5 \times 10^{-3}$

Step 2: Calculate Mole Fraction

The mole fraction of methane (x) in the solution is given by:

$x = \frac{n(\text{methane})}{n(\text{methane})+n(\text{water})}$

With 1 L of water, the number of moles of water is:

$n(\text{water}) = \frac{1000\text{ g}}{18\text{ g/mol}} = 55.5\,\text{mol}$

Assuming $n(\text{methane})$ is significantly smaller than 55.5 mol, substituting into Equation 2 simplifies to:

$2.5 \times 10^{-3} = \frac{n}{55.5}$

Solving for n, the moles of methane:

$n = 2.5 \times 10^{-3} \times 55.5$

$n = 1.38 \times 10^{-1}$

Thus, the number of moles of methane dissolved in 1 L of water is 1.38×10^{-1}.

Question34

0.05 mole of a non-volatile solute is dissolved in 500 g of water. What is the depression in freezing point of resultant solution?

$$\left(K_f\left(\mathrm{H}_2\mathrm{O}\right)=1.86\ \mathrm{K\ kg\ mol}^{-1}\right)$$

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Options:

A. 0.047 K

B. 0.372 K

C. 0.093 K

D. 0.186 K

Answer: D

Solution:

$$\begin{aligned}\text{Given, } n &= 0.05 \text{ mol} \\ V &= 500 \text{ g of water} \\ K_f &= 1.86 \text{ k kg/mol} \\ \Delta T_f &= ?\end{aligned}$$

$$\begin{aligned}\therefore \text{ Molality } &= \frac{\text{Number of moles}}{\text{Weight of solvent (kg)}} \\ &= \frac{0.05}{0.5} \\ \Delta T_f &= m \times K_f \\ &= \frac{0.05}{0.5} \times 1.86 \\ &= 0.186 \text{ K}\end{aligned}$$

Question35

Which of the following form an ideal solution?

I. Chloroethane and bromoethane

II. Benzene and toluene

III. n-hexane and n-heptane

IV. Phenol and aniline

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Options:

A. I and II

B. I, II and III

C. II, III and IV

D. I and IV

Answer: B

Solution:

Ideal solutions are the solutions which obey Raoult's law at all temperatures and pressures. The interactions of solute-solute, solvent-solvent and solute-solvent in these solutions are almost similar. No association/dissociation occurs in these solutions.

Mixture of phenol and aniline shows negative deviation from ideal behaviour and forms an azeotropic mixture. The phenol-aniline interactions are stronger than phenol-phenol or aniline-aniline interactions.

$$\Delta V_{\text{mixing}} \quad < 0$$
$$\Delta H_{\text{mixing}} \quad > 0$$

Hence, it does not form an ideal solution.

Rest options (I), (II) and (III) form ideal solution.

Question36

A solution containing 6.0 g of urea is isotonic with a solution containing 10 g of a non-electrolytic solute X. The molar mass of X (in g mol^{-1}) is

Options:

A. 50.0

B. 100

C. 75.0

D. 68.0

Answer: B

Solution:

Mass of urea, $m_U = 6$ g

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

Moles of urea, $n_v = 6/60 = 0.1$ mol

Mass of $X, m_X = 10$ g

As solutions are isotonic, so

$\pi_U = \pi_x$
$C_U RT = C_x RT$

Assuming volume of solvent to be same in both cases,

$n_v = n_X$
$0.1 = n_X \quad$ where, $m_x =$ mass of X

$\Rightarrow \quad \frac{m_X}{M_X} = 0.1 \quad M_X =$ Molar mass of X

$M_X = \frac{10}{0.1} = 100 \text{ g mol}^{-1}$

Question37

$x\%(w/V)$ solution of urea is isotonic with 4% (w/V) solution of a non-volatile solute of molar mass 120 g mol^{-1}. The value of x is

Options:

A. 2

B. 4

C. 3

D. 5

Answer: A

Solution:

$4\% w/V$ solution means 100 mL of solution contains 4 g of non-volatile solute.

Molar mass of solute $= 120 \text{ g mol}^{-1}$

Moles, $n_s = \frac{4}{120} = \frac{1}{30}$ mol

$x\% \left(\frac{w}{V}\right)$, so mass of urea is x g.

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

Using formula, number of moles $= \frac{\text{Mass}}{\text{Molar mass}}$

Moles, $n_U = \frac{x}{60}$ mol

For isotonic solution, $\pi_s = \pi_U$

$$\begin{aligned} C_o &= C_S \\ n_U &= n_S \\ \frac{x}{60} &= \frac{1}{30} \\ x &= 2 \end{aligned}$$

$\therefore$ $2\%(w/V)$ solution of urea is isotonic with $4\%\ (w/V)$ solution of non-volatile solute.

Question38

At $T(\text{ K}) \times$ g of a non-volatile solid (molar mass 78 g mol^{-1}) when added to 0.5 kg water, lowered its freezing point by 1.0°C. What is x (in g)? (K_f of water at $T(\text{ K}) = 1.86 \text{ K kg mol}^{-1}$)

Options:

A. 10.48

B. 20.96

C. 41.92

D. 5.24

Answer: B

Solution:

Given, mass of non-volatile solid = x g

Molar mass = 78 g mol^{-1}

Mass of water = 0.5 kg

K_f (water) $= 1.86\ \mathrm{K\ kg}^{-1}\ \mathrm{mol}^{-1}$

Lowering in freezing point, $\Delta T_f = 1^{\circ}\mathrm{C}$Using the equation;

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

Now, substituting the given values in above formula,

$m = \frac{1}{1.86}$ kg

We know that molality

$$= \frac{\text{Number of moles of solute}}{\text{Mass of solvent (in kg)}}$$

$$\frac{n}{M} = \frac{1}{1.86}; \quad [n = \text{ number of moles of solute}$$

$$\left.\text{So, } n = \frac{1}{1.86} \times 0.5 = \text{ mass of solvent in kg}\right]$$

$$\frac{x}{78} = \frac{0.5}{1.86}$$

$$\Rightarrow \quad x = \frac{0.5 \times 78}{1.86}$$

$x = 20.96$ g

Question39

Assertion (A) Blood cells collapse when suspended in saline water.

Reason (R) Cell membrane dissolves in saline water.

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Options:

A. Both A and R are correct and R is the correct explanation of A .

B. Both A and R are correct, but R is not the correct explanation of A.

C. A is correct but R is not correct.

D. A is incorrect but R is correct.

Answer: C

Solution:

Blood serum contains 0.9% salt. When blood cells placed in saline water, they lose water due to exosmosis and collapses. Thus, assertion is correct.

The water moves through cell membrane during this process. Thus, reason is not correct.

Hence, option (c) is correct. i.e. A is correct but R is not correct.

Question40

If two liquids A and B form a minimum boiling azeotrope at some specific composition, then which statement among the following is correct?

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Options:

A. $A-B$ molecular interactions are stronger than $A-A$ and $B-B$ interactions.

B. The total vapour pressure of the mixture is greater than that corresponding to an ideal solution.

C. There is a slight decrease in volume when the two components are mixed.

D. ΔH_{mix} will have a negative value.

Answer: B

Solution:

When a minimum boiling azeotrope is formed, the resulting solution will have maximum vapour pressure as it has minimum boiling point.

Interaction between $A-A$ or $B-B$ are stronger than $A-B$ interaction.

$\therefore \Delta V>0, \Delta H_{\text{mix}}>0$

$\therefore$ Only statement (b) is correct.

Question41

The vapour pressure of a solvent decreased by20 mm of Hg when a non-volatile solute wasadded to the solvent. The mole fraction of thesolute in the solution is 0.5. What should bethe mole fraction of the solvent for thedecrease in the vapour pressure needs to be10 mm of Hg?

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Options:

A. $\frac{3}{4}$

B. $\frac{2}{3}$

C. $\frac{1}{4}$

D. $\frac{3}{2}$

Answer: A

Solution:

The decrease in vapour pressure can be calculated as:

$p^{\Upsilon} - p_s = p^{\Upsilon} \times$ mole fraction of solute

We are told that the vapour pressure goes down by 20 mm of Hg when the mole fraction of solute is 0.5. So,

$20 = p^{\Upsilon} \times 0.5$ ... (i)

If we want the vapour pressure to drop by only 10 mm of Hg, let the new mole fraction of solute be χ_2:

$10 = p^{\Upsilon} \times \chi_2$ ... (ii)

Now, let's solve for χ_2:

From equation (i): $p^{\Upsilon} = 40$

Using this in equation (ii):

$10 = 40 \times \chi_2$

So, $\chi_2 = \frac{10}{40} = \frac{1}{4}$

The mole fraction of the solvent (χ_1) is what is left after subtracting the solute's mole fraction from 1:

$\chi_1 = 1 - \frac{1}{4} = \frac{3}{4}$

Question42

If the K_{H} values for $\mathrm{Ar}(g), \mathrm{CO}_2(g), \mathrm{HCHO}(g)$ and $\mathrm{CH}_4(\mathrm{g})$ respectively are $40.39, 1.67, 1.83 \times 10^{-5}$ and 0.413, then identify the correct increasing order of their solubilities.

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Options:

A. $\mathrm{HCHO} < \mathrm{CH}_4 < \mathrm{CO}_2 < \mathrm{Ar}$

B. $\mathrm{HCHO} < \mathrm{CO}_2 < \mathrm{CH}_4 < \mathrm{Ar}$

C. $\mathrm{Ar} < \mathrm{CO}_2 < \mathrm{HCHO} < \mathrm{CH}_4$

D. $\mathrm{Ar} < \mathrm{CO_2} < \mathrm{CH_4} < \mathrm{HCHO}$

Answer: D

Solution:

Higher the value of K_H, lower will be the solubility of the gas at a given pressure.

Solubility $\propto \frac{1}{K_H}$

So, the correct order is

$\mathrm{Ar} < \mathrm{CO_2} < \mathrm{CH_4} < \mathrm{HCHO}$

Question43

If 500 mL of $CaCl_2$ solution contains 3.01×10^{22} chloride ions, molarity of the solution will be

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Options:

A. 0.05 M

B. 0.01 M

C. 0.1 M

D. 0.02 M

Answer: A

Solution:

$$\mathrm{CaCl_2} \longrightarrow \mathrm{Ca^{2+}} + 2\mathrm{Cl^-}$$
$$\mathrm{Cl^-} = 3.01 \times 10^{22} \text{ ions}$$

For

$$2\mathrm{Cl}^- = \frac{3.01 \times 10^{22}}{2}$$
$$= 1.505 \times 10^{22}\mathrm{CaCl}_2 \text{ required}$$
$$\text{Mole} = \frac{1.505 \times 10^{22}}{6.02 \times 10^{23}} = \frac{0.25}{10} = 0.025 \text{ mol}$$

Molarity for 0.025 mole of CaCl_2

$= \frac{0.025}{500} \times 1000 = 0.05\mathrm{M}$

Question44

Which condition is not satisfied by an ideal solution?

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Options:

A. Both $\Delta_{\text{mix}} \quad H = 0$ and $\Delta_{\text{mix}} \quad S = 0$

B. Obeyance on Raoult's law

C. Both $\Delta_{\text{mix}} \quad H = 0$ and $\Delta_{\text{mix}} \quad V = 0$

D. $\Delta_{\text{mix}} \quad H = 0$

Answer: A

Solution:

An ideal solution is a homogeneous mixture which $\left(\Delta_{\text{mix}} \quad S = 0\right)$ obey Raoult's law.

For ideal solution, we need $\Delta_{\text{mix}} \quad H = 0$ and $\Delta_{\text{mix}} \quad V = 0$ but $\Delta_{\text{mix}} \quad S \neq 0$ i.e. $\Delta_{\text{mix}} \quad S > 0$ obey Raoult's law at every range of concentration for $\Delta_{\text{mix}} \quad S > 0$.

Number of molecule = entropy increase

$\Delta_{\text{mix}} \quad G = \Delta H - T\Delta S$
$\Delta_{\text{mix}} \quad G = -T\Delta S < 0$

e.g., benzene + toluene (show ideal solution)

Question45

A solution of urea (molar mass 60 g mol^{-1} **) boils at** $100.20^{\circ}\mathrm{C}$ **at the atmospheric pressure, if** K_f **and** K_b **for water are 1.86 and** $0.512 \mathrm{~K}$ kgmol^{-1} **respectively. The freezing point of the solution will be**

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Options:

A. $-0.654\Upsilon\mathrm{C}$

B. $+0.654\Upsilon\mathrm{C}$

C. $-0.726\Upsilon\mathrm{C}$

D. $+0.726\Upsilon\mathrm{C}$

Answer: C

Solution:

Formula used,

Freezing point of solution, $\Delta T_l = K_f \times$ molality of solution and $\Delta T_b = K_b \times$ molatity of solution

Now, ratio of K_f and K_b is $\frac{\Delta T_f}{\Delta T_b} = \frac{K_f}{K_b}$ ..... (i)

Given that, $\Delta T_b = T_2 - T_1 = 100.20 - 100 = 0.20^{\circ}\mathrm{C}$

K_f of water $= 1.86 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

K_b of water $= 0.512 \mathrm{~K} \mathrm{~kg} \mathrm{~mol}^{-1}$

Put value in Eq. (i),

$\therefore \quad \frac{\Delta T_f}{0.20} = \frac{1.86}{0.512} \Rightarrow \Delta T_f = 0.726$

Now, $\Delta T_f \Rightarrow T_1 - T_2$

$0.726^{\circ}\mathrm{C} = 0^{\circ}\mathrm{C} - T_2 \Rightarrow T_2 = -0.726^{\circ}\mathrm{C}$

Hence, freezing point of solution is $-0.726^{\circ}\mathrm{C}$

Question46

When difference in boiling points of twoliquids is too small, then the separation iscarried out by

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Options:

A. steam distillation

B. simple distillation

C. fractional distillation

D. vacuum distillation

Answer: C

Solution:

Fractional distillation is used for theseparation of a mixture of two or more miscibleliquids for which the difference in boiling points isless than 25K.

e.g. CCl_4 (Carbon tetrachloride) boiling point is350 K and anisole (boiling point is 427 K can beseparated by fractional distillation.

The process involves heating the mixture andpartial condensation of vapours along a column,which is set up such that components with lowerboiling points pass through the column and arecollected earlier than components with higherboiling points.

Question47

Which of the following will form an ideal solution?

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Options:

A. C_2H_5OH and H_2O

B. HNO_3 and H_2O

C. $CHCl_3$ and CH_3COCH_3

D. C_6H_6 and $C_6H_5CH_3$

Answer: D

Solution:

Solutions of benzene and toluene, which havevery similar molecular structure, are ideal. Anymixture of the two has a volume equal to the sumof the volumes of the separate components andthe mixing process occurs without absorption orevolution of heat.

Benzene and toluene solution over the entirerange of composition is ideal. The vapour pressureof pure benzene and toluene at 300 K are 50.

Question48

The molal elevation constant is the ratio ofelevation in boiling point to

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Options:

A. molarity

B. molality

C. mole fraction of solute

D. mole fraction of solvent

Answer: B

Solution:

The molal elevation constant is the ratio of elevation in boiling point to molality of solution.

$$\Delta T_b = K_b \times m$$

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

Here, K_b = molal elevation constant

ΔT_b = increase in boiling point

m = molality of solution

Due to solute presence in solution, the solution's boiling point is greater than the pure solvent's boiling point. This is called elevation in boiling point (ΔT_b).
