# SAI WISDOM ACADEMY

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## CLASS XII - CHEMISTRY

### Chapter 1: Solutions

Complete Notes - Worked Examples - Practice Questions

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# Chapter 1 : Solutions

A **solution** is a homogeneous mixture of two or more chemically non-reacting substances whose composition can be varied within certain limits. The component present in the largest amount is called the **solvent**; the other components are called **solutes**. This chapter develops the quantitative language used to describe such mixtures, the laws that govern the behaviour of dilute solutions, and the properties that depend only on the number of solute particles present.

## 1. Types of Solutions

Based on the physical state of the solute and solvent, nine binary solutions are possible. The most commonly studied ones are listed below.

| Type | Solute | Solvent | Example |
|---|---|---|---|
| Solid in liquid | Solid | Liquid | Sugar in water |
| Liquid in liquid | Liquid | Liquid | Ethanol in water |
| Gas in liquid | Gas | Liquid | Oxygen dissolved in water |
| Solid in solid | Solid | Solid | Brass (Cu + Zn) |
| Gas in gas | Gas | Gas | Air |
| Liquid in solid | Liquid | Solid | Hydrated salts |
| Gas in solid | Gas | Solid | Hydrogen in palladium |

## 2. Expressing Concentration of Solutions

Concentration tells us how much solute is dissolved in a given quantity of solvent or solution. Several measures are used depending on the context.

### 2.1 Mass percentage (w/w)

Mass of solute present in 100 g of solution.

$$\text{Mass \% of A} = (\text{Mass of A} / \text{Total mass of solution}) \times 100$$

### 2.2 Volume percentage (v/v)

Volume of solute present in 100 mL of solution.

$$\text{Volume \% of A} = (\text{Volume of A} / \text{Total volume of solution}) \times 100$$

### 2.3 Mass by volume percentage (w/v)

Mass of solute (in g) present in 100 mL of solution. Commonly used in medicine and pharmacy.

### 2.4 Parts per million (ppm)

Used when the solute is present in trace amounts.

$$\text{ppm of A} = (\text{Number of parts of A} / \text{Total number of parts of solution}) \times 10^6$$

### 2.5 Mole fraction (x)

Ratio of moles of one component to the total moles of all components.

$$x_A = n_A / (n_A + n_B), \quad x_A + x_B = 1$$

### 2.6 Molarity (M)

Number of moles of solute dissolved per litre of solution.

$$M = (\text{Moles of solute}) / (\text{Volume of solution in L}) \quad \text{Units: mol } L^{-1}$$

**Note:** Molarity is temperature dependent because the volume of a solution changes with temperature.

### 2.7 Molality (m)

Number of moles of solute dissolved per kilogram of solvent.

$$m = (\text{Moles of solute}) / (\text{Mass of solvent in kg}) \quad \text{Units: mol } kg^{-1}$$

Molality is independent of temperature because it involves only masses.

### Worked Example 2.1

Calculate the molarity of a solution prepared by dissolving 4.0 g of NaOH in enough water to make 250 mL of solution.

**Solution:** Moles of NaOH = 4.0 / 40 = 0.10 mol.
Volume of solution = 250 mL = 0.250 L.
Molarity = 0.10 / 0.250 = **0.40 mol $L^{-1}$**.

### Worked Example 2.2

Find the molality of a solution containing 18.25 g of HCl in 500 g of water.

**Solution:** Moles of HCl = 18.25 / 36.5 = 0.500 mol.
Mass of solvent = 500 g = 0.500 kg.
Molality = 0.500 / 0.500 = **1.00 mol $kg^{-1}$**.

## 3. Solubility

**Solubility** of a substance is the maximum amount of it (in grams) that can be dissolved in 100 g of solvent at a specified temperature to form a saturated solution.

### 3.1 Solubility of solids in liquids

Every solid does not dissolve in a given liquid. The rule "like dissolves like" applies: polar solutes dissolve in polar solvents, non-polar in non-polar.

**Effect of temperature:** If dissolution is endothermic ($\Delta_{sol}H > 0$), solubility increases with rising temperature. If it is exothermic ($\Delta_{sol}H < 0$), solubility decreases with rising temperature.

**Effect of pressure:** Negligible, because solids and liquids are highly incompressible.

### 3.2 Solubility of gases in liquids

Gases dissolve in liquids to varying extents. Oxygen, for instance, is only sparingly soluble in water - yet that small amount sustains aquatic life.

Solubility of a gas in a liquid depends on:

- nature of the gas and the solvent,
- temperature (solubility usually decreases as T rises - dissolution of a gas is exothermic),
- pressure of the gas above the solution (solubility increases with pressure).

### 3.3 Henry's Law

**Statement:** At a constant temperature, the partial pressure of a gas in the vapour phase (p) is directly proportional to the mole fraction (x) of the gas in the solution.

$$p = K_H \times x$$

Here $K_H$ is the **Henry's law constant**. Its value depends on the nature of the gas and on temperature. A higher $K_H$ at a given temperature means lower solubility of the gas.

**Applications of Henry's law:**

- To increase the solubility of $CO_2$ in soft drinks and soda water, the bottle is sealed under high pressure.
- Scuba divers breathing compressed air at depth dissolve more $N_2$ in their blood. If they ascend rapidly, $N_2$ bubbles form in the bloodstream causing **bends**. Diver tanks are therefore filled with helium-oxygen mixtures.
- At high altitudes the partial pressure of $O_2$ is lower, so less $O_2$ dissolves in blood, leading to **anoxia**.

**Worked Example 3.1**

If $N_2$ gas is bubbled through water at 293 K, how many millimoles of $N_2$ gas would dissolve in 1 L of water? Assume that $N_2$ exerts a partial pressure of 0.987 bar. Given: $K_H$ for $N_2$ at 293 K = 76.48 kbar.

**Solution:** $x(N_2) = p / K_H = 0.987 / 76480 = 1.29 \times 10^{-5}$.
Moles of water in 1 L ≈ 1000/18 = 55.5 mol. Since $x(N_2)$ is very small, $n(N_2) \approx x(N_2) \times n(H_2O) = 1.29 \times 10^{-5} \times 55.5 = 7.16 \times 10^{-4}$ mol = **0.716 mmol**.

## 4. Vapour Pressure of Liquid Solutions

The pressure exerted by the vapour of a liquid in equilibrium with the liquid at a given temperature is called its **vapour pressure**. When a non-volatile solute is dissolved in a volatile solvent, fewer solvent molecules occupy the surface, so the vapour pressure decreases.

### 4.1 Raoult's Law (volatile solute)

For a solution of two volatile, miscible liquids A and B, the partial vapour pressure of each component is directly proportional to its mole fraction in the solution.

$$p_A = p_A^{\circ} \cdot x_A, \quad p_B = p_B^{\circ} \cdot x_B$$

Total vapour pressure (by Dalton's law):

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

Here $p_A^{\circ}$ and $p_B^{\circ}$ are the vapour pressures of pure A and pure B at the same temperature.

### 4.2 Raoult's Law (non-volatile solute)

If solute B is non-volatile, only solvent A contributes to the vapour. The total vapour pressure of the solution is

$$p_{total} = p_A^{\circ} \cdot x_A$$

Since $x_A = 1 - x_B$, the relative lowering of vapour pressure equals the mole fraction of the solute:

$$(p_A^{\circ} - p_A) / p_A^{\circ} = x_B$$

Raoult's law for non-volatile solutes is in fact a special case of the more general Raoult's law applied to volatile components.

## 5. Ideal and Non-Ideal Solutions

An **ideal solution** obeys Raoult's law over the entire range of composition at all temperatures. For it,

- $\Delta_{mix}H = 0$ (no heat is absorbed or released on mixing),
- $\Delta_{mix}V = 0$ (no volume change on mixing),
- Solute-solvent interactions are similar in magnitude to solute-solute and solvent-solvent interactions.

Examples: benzene + toluene, n-hexane + n-heptane, chlorobenzene + bromobenzene.

### 5.1 Non-ideal solutions

Solutions that do not obey Raoult's law are called **non-ideal**. Depending on the nature of solute-solvent interactions, deviations may be positive or negative.

| Feature | Positive deviation | Negative deviation |
|---|---|---|
| Interaction A-B | Weaker than A-A and B-B | Stronger than A-A and B-B |
| Observed p | Higher than predicted by Raoult's law | Lower than predicted by Raoult's law |
| $\Delta_{mix}H$ | Positive (endothermic) | Negative (exothermic) |
| $\Delta_{mix}V$ | Positive (volume increases) | Negative (volume decreases) |
| Examples | Ethanol + water, acetone + $CS_2$ | HCl + water, $CHCl_3$ + acetone |

### 5.2 Azeotropes

An **azeotrope** is a binary mixture that boils at a constant temperature and distils without any change in composition. Azeotropes form when deviations from Raoult's law are large.

- **Minimum boiling azeotropes** arise from solutions showing large *positive* deviation. Example: ethanol-water mixture containing about 95% ethanol by volume boils at 351.15 K.
- **Maximum boiling azeotropes** arise from solutions showing large *negative* deviation. Example: $HNO_3$-water mixture containing about 68% $HNO_3$ by mass boils at 393.5 K.

## 6. Colligative Properties

Properties of dilute solutions of a non-volatile solute that depend only on the **number** of solute particles relative to the solvent, and not on the chemical nature of the solute, are called **colligative properties**.

The four colligative properties are:

- Relative lowering of vapour pressure
- Elevation of boiling point
- Depression of freezing point
- Osmotic pressure

### 6.1 Relative Lowering of Vapour Pressure

Already derived above:

$$(p_A^{o} - p_A) / p_A^{o} = x_B = n_B / (n_A + n_B)$$

For very dilute solutions $n_B << n_A$, so this reduces to:

$$(p_A^{o} - p_A) / p_A^{o} \approx n_B / n_A = (w_B M_A) / (M_B w_A)$$

where w and M denote masses and molar masses respectively. This expression allows determination of the molar mass $M_B$ of the solute.

## 6.2 Elevation of Boiling Point

A liquid boils at the temperature at which its vapour pressure equals the atmospheric pressure. Since the addition of a non-volatile solute lowers the vapour pressure, the solution must be heated to a higher temperature to boil. This rise is called **elevation of boiling point** $\Delta T_b$.

$$\Delta T_b = T_b - T_b^{\circ} = K_b \cdot m$$

$K_b$ is the **molal elevation constant** (or ebullioscopic constant). For water $K_b = 0.52$ K kg mol$^{-1}$.

Combining with the definition of molality, the molar mass of the solute is

$$M_B = (1000 \cdot K_b \cdot w_B) / (\Delta T_b \cdot w_A)$$

### Worked Example 6.1

A solution containing 18 g of a non-volatile solute in 200 g of water boils at 100.26 °C. Find the molar mass of the solute. ($K_b$ for water = 0.52 K kg mol$^{-1}$.)

**Solution:** $\Delta T_b = 0.26$ K.
$M_B = (1000 \times 0.52 \times 18) / (0.26 \times 200) = 9360 / 52 =$ **180 g mol$^{-1}$**.

## 6.3 Depression of Freezing Point

The freezing point of a pure liquid is the temperature at which the solid and liquid phases have the same vapour pressure. Because a non-volatile solute lowers the vapour pressure of the liquid, the vapour pressure curve of the solution meets that of the solid at a lower temperature. The resulting decrease in freezing point is $\Delta T_f$.

$$\Delta T_f = T_f^{\circ} - T_f = K_f \cdot m$$

$K_f$ is the **molal depression constant** (cryoscopic constant). For water $K_f = 1.86$ K kg mol$^{-1}$.

$$M_B = (1000 \cdot K_f \cdot w_B) / (\Delta T_f \cdot w_A)$$

### Worked Example 6.2

45 g of ethylene glycol ($C_2H_6O_2$) is dissolved in 600 g of water. Calculate the freezing point of the resulting solution. ($K_f$ for water = 1.86 K kg mol$^{-1}$.)

**Solution:** Molar mass of ethylene glycol = 62 g mol$^{-1}$. Moles = 45 / 62 = 0.7258.
Molality m = 0.7258 / 0.600 = 1.210 mol kg$^{-1}$.
$\Delta T_f = 1.86 \times 1.210 = 2.25$ K.
Freezing point of solution = 273.15 - 2.25 = **270.90 K** (i.e. about -2.25 °C).

## 6.4 Osmosis and Osmotic Pressure

**Osmosis** is the spontaneous flow of solvent molecules through a **semi-permeable membrane** from a region of lower solute concentration (or pure solvent) into a region of higher solute concentration. The hydrostatic pressure that must be applied to the solution side to just prevent this flow is the **osmotic pressure** ($\pi$).

For dilute solutions, osmotic pressure obeys an equation analogous to the ideal gas law:

$$\pi = C\,R\,T = (n_B / V)\,R\,T$$

where C is the molar concentration of the solute, R is the gas constant, and T is the absolute temperature.

Hence the molar mass of the solute is:

$$M_B = (w_B\,R\,T) / (\pi\,V)$$

**Isotonic, hypertonic and hypotonic solutions:**

- Two solutions with the same osmotic pressure at the same temperature are **isotonic** (e.g. 0.9 % NaCl solution is isotonic with blood).
- A solution of higher osmotic pressure than another is **hypertonic**; cells placed in it shrink due to exosmosis.
- A solution of lower osmotic pressure is **hypotonic**; cells placed in it swell, and may burst.

**Reverse osmosis** occurs when a pressure larger than the osmotic pressure is applied to the solution side. Solvent then flows from the solution into the pure solvent compartment. This principle is used in **desalination of sea water** using cellulose acetate membranes.

**Worked Example 6.3**

200 $cm^3$ of an aqueous solution of a protein contains 1.26 g of the protein. The osmotic pressure of the solution at 300 K is found to be $2.57 \times 10^{-3}$ bar. Calculate the molar mass of the protein. (R = 0.0831 L bar $mol^{-1}$ $K^{-1}$.)

**Solution:** V = 0.200 L.
$M_B = (w_B\,R\,T) / (\pi\,V)$
$= (1.26 \times 0.0831 \times 300) / (2.57 \times 10^{-3} \times 0.200)$
$= 31.40 / 5.14 \times 10^{-4}$
$= \mathbf{6.11 \times 10^4\ g\ mol^{-1}}$.

## 7. Abnormal Molar Masses and the van't Hoff Factor

Colligative properties depend on the number of solute particles. When the solute undergoes **association** (e.g. benzoic acid dimerising in benzene) or **dissociation** (e.g. NaCl in water giving $Na^+$ and $Cl^-$), the actual number of particles in solution is different from the number calculated assuming no association/dissociation. The molar masses obtained from colligative measurements are then "abnormal".

To account for this, van't Hoff introduced the factor **i**:

**i = (Observed colligative property) / (Calculated colligative property assuming no change in particles)**
**= (Normal molar mass) / (Observed molar mass)**
**= (Total moles of particles after association/dissociation) / (Moles of particles before)**

Modified colligative-property equations become:

**Relative lowering of vapour pressure: $(p_A^o - p_A) / p_A^o = i \cdot x_B$**
**Elevation of boiling point: $\Delta T_b = i K_b m$**
**Depression of freezing point: $\Delta T_f = i K_f m$**
**Osmotic pressure: $\pi = i C R T$**

**Interpretation of i:**

- i = 1 → solute neither associates nor dissociates (ideal behaviour).
- i > 1 → solute dissociates (e.g. i ≈ 2 for KCl, i ≈ 3 for $K_2SO_4$ if dissociation were complete).
- i < 1 → solute associates (e.g. i ≈ 0.5 for benzoic acid dimerising in benzene).

### Worked Example 7.1

Calculate the freezing point of an aqueous solution containing 10.50 g of $MgBr_2$ in 200 g of water, assuming $MgBr_2$ is completely dissociated. (Molar mass of $MgBr_2$ = 184 g $mol^{-1}$; $K_f$ of water = 1.86 K $kg\ mol^{-1}$.)

**Solution:** Moles of $MgBr_2$ = 10.50 / 184 = 0.0571.
Molality m = 0.0571 / 0.200 = 0.285 mol $kg^{-1}$.
$MgBr_2 \rightarrow Mg^{2+} + 2\ Br^-$, so i = 3.
$\Delta T_f = i K_f m = 3 \times 1.86 \times 0.285 = 1.59$ K.
Freezing point of solution = 273.15 - 1.59 = **271.56 K**.

## 8. Quick Recap

| Quantity | Formula | Units |
|---|---|---|
| Mole fraction | $x_A = n_A / \Sigma n$ | - |
| Molarity | $M = n_B / V_{solution}$ | mol $L^{-1}$ |
| Molality | $m = n_B / w_A(kg)$ | mol $kg^{-1}$ |
| Henry's law | $p = K_H x$ | - |
| Raoult's law (volatile) | $p_i = p_i^o x_i$ | - |
| Lowering of v.p. | $(p_A^o - p_A)/p_A^o = x_B$ | - |
| Elevation of b.p. | $\Delta T_b = i K_b m$ | K |
| Depression of f.p. | $\Delta T_f = i K_f m$ | K |
| Osmotic pressure | $\pi = i C R T$ | bar / atm |

## 9. Practice Question Bank

Attempt all questions on your own before looking up the answers at the end of this section. Where numerical values are needed, use: $K_f$(water) = 1.86, $K_b$(water) = 0.52 (both in K kg $mol^{-1}$); R = 0.0831 L bar $K^{-1}$ $mol^{-1}$.

**Q1. Define the terms (i) mole fraction, (ii) molality, (iii) molarity. Which of these is independent of temperature, and why?**

**Q2. What is meant by the term solubility? How does the solubility of a gas in a liquid vary with (a) temperature and (b) partial pressure?**

**Q3. State and explain Henry's law. Give two important applications.**

**Q4. Why is the air dissolved in water richer in oxygen than the air above the water?**

**Q5. State Raoult's law for a solution of two volatile liquids and derive the expression for the total vapour pressure.**

**Q6. Why does the vapour pressure of a solvent decrease when a non-volatile solute is added to it?**

**Q7. Distinguish between ideal and non-ideal solutions with two examples of each.**

**Q8. What are azeotropes? Give one example each of minimum and maximum boiling azeotropes.**

**Q9. Define colligative properties. Why are they so called?**

**Q10. Why is it that the boiling point of a 1 m KCl(aq) solution is higher than that of a 1 m glucose(aq) solution?**

**Q11. Calculate the molarity of a solution prepared by dissolving 5.85 g of NaCl in water and making the final volume 250 mL.**

**Q12. Compute the molality of a solution prepared by dissolving 9.8 g of $H_2SO_4$ in 500 g of water.**

**Q13. An aqueous solution contains 0.10 mole of glucose in 250 g of water. Find its molality and the mole fraction of glucose.**

**Q14. The vapour pressure of pure benzene at a certain temperature is 0.850 bar. A non-volatile, non-electrolyte solid weighing 0.5 g is dissolved in 39.0 g of benzene; the vapour pressure of the solution becomes 0.845 bar. Find the molar mass of the solid. ($M_{benzene}$ = 78)**

**Q15. If $N_2$ gas is bubbled through water at 293 K, how many moles of $N_2$ will dissolve in 1 L of water? (Partial pressure of $N_2$ = 0.987 bar; $K_H$ = 76.48 kbar.)**

**Q16. A 5% solution by mass of cane sugar (M = 342) is isotonic with a 0.877% solution by mass of a substance X. Find the molar mass of X.**

**Q17. Calculate the boiling point of a solution prepared by dissolving 60 g of urea (M = 60) in 1 kg of water at 1 atm pressure.**

**Q18. 20 g of an organic compound dissolved in 500 g of water lowers the freezing point of water by 1.24 K. Calculate the molar mass of the compound.**

**Q19. Define osmotic pressure. Why is the colligative property of osmotic pressure preferred over others for determining the molar mass of biomolecules?**

**Q20. Give reasons: (a) Aquatic species are more comfortable in cold water than in warm water. (b) Lower partial pressure of $O_2$ at high altitude causes climbers to feel weak. (c) A cold drink bottle is sealed under high pressure.**

**Q21. What is the van't Hoff factor? How is it related to the degree of dissociation $\alpha$ of a binary electrolyte AB?**

**Q22. 2 g of benzoic acid ($C_6H_5COOH$) dissolved in 25 g of benzene shows a depression in freezing point equal to 1.62 K. Molal depression constant for benzene is 4.9 K kg $mol^{-1}$. What is the percentage association of acid if it forms a dimer in solution?**

**Q23. 0.6 mL of acetic acid ($CH_3COOH$) having density 1.06 g/mL is dissolved in 1 L of water. The depression in freezing point observed is 0.0205 K. Calculate the van't Hoff factor and the dissociation constant of acid. ($K_f$(water) = 1.86; M($CH_3COOH$) = 60.)**

**Q24. Determine the osmotic pressure of a solution prepared by dissolving 25 mg of $K_2SO_4$ in 2 L of water at 25 °C, assuming $K_2SO_4$ is completely dissociated. (M = 174.)**

**Q25. State four important industrial / biological applications of osmosis and reverse osmosis.**

## 10. Answers and Solution Hints

**A1.** Mole fraction is the ratio of moles of one component to total moles; molality is moles of solute per kg of solvent; molarity is moles of solute per litre of solution. Mole fraction and molality are **temperature independent** because they involve only masses or mole counts. Molarity changes with T because volume changes with T.

**A2.** Solubility is the maximum mass of solute that dissolves in 100 g of solvent at a given temperature to form a saturated solution. (a) Solubility of a gas usually **decreases** as T rises (dissolution is exothermic). (b) It **increases** with increasing partial pressure of the gas (Henry's law).

**A3.** $p = K_H\, x$ at constant T. Applications: pressurising soft drinks; use of He-$O_2$ mixtures in scuba tanks; treatment of anoxia at high altitude with $O_2$-enriched air.

**A4.** $K_H$ for $O_2$ is smaller than for $N_2$, so at a given partial pressure a larger mole fraction of $O_2$ dissolves; the dissolved air is therefore richer in oxygen than atmospheric air.

**A5.** $p_A = p_A^{\circ}\, x_A$, $p_B = p_B^{\circ}\, x_B$; $p_{total} = p_A + p_B = p_B^{\circ} + (p_A^{\circ} - p_B^{\circ})\, x_A$.

**A6.** Surface molecules of solvent are partially replaced by solute molecules. Fewer solvent molecules escape into the vapour phase, lowering the vapour pressure.

**A7.** Ideal: obeys Raoult's law; $\Delta_{mix}H = 0$, $\Delta_{mix}V = 0$. Examples: benzene + toluene; n-hexane + n-heptane. Non-ideal: deviates from Raoult's law. Positive deviation example: ethanol + water. Negative deviation example: $CHCl_3$ + acetone.

**A8.** Azeotropes are constant-boiling mixtures that distil unchanged. Minimum boiling: ethanol-water (~95% ethanol). Maximum boiling: $HNO_3$-water (~68% $HNO_3$).

**A9.** They depend only on the number of solute particles, not their identity. The four are: relative lowering of vapour pressure, elevation of b.p., depression of f.p. and osmotic pressure.

**A10.** KCl dissociates: $KCl \rightarrow K^+ + Cl^-$, giving ~ twice as many particles as glucose at the same molality. Hence $\Delta T_b$ is larger.

**A11.** Moles NaCl = 5.85/58.5 = 0.10; V = 0.250 L; M = **0.40 mol $L^{-1}$**.

**A12.** Moles $H_2SO_4$ = 9.8/98 = 0.10; m = 0.10/0.500 = **0.20 mol $kg^{-1}$**.

**A13.** m = 0.10/0.250 = **0.40 mol $kg^{-1}$**; $n(H_2O)$ = 250/18 = 13.89; x(glucose) = 0.10/(0.10+13.89) = **$7.15 \times 10^{-3}$**.

**A14.** $(\Delta p)/p_A^{\circ} = (0.850 - 0.845)/0.850 = 5.88 \times 10^{-3}$. n(benzene) = 39/78 = 0.500. n(solute) ≈ $5.88\times10^{-3} \times 0.500 = 2.94\times10^{-3}$. M(solute) = 0.5 / $2.94\times10^{-3}$ = **170 mol $g^{-1}$**.

**A15.** $x(N_2) = p/K_H = 0.987/76480 = 1.29\times10^{-5}$. $n(H_2O)$ in 1 L ≈ 55.5. $n(N_2)$ ≈ **$7.16\times10^{-4}$ mol**.

**A16.** Isotonic solutions have equal molar concentration. (5/342) = ($0.877/M_X$) (per 100 g solution). $M_X = 0.877 \times 342 / 5$ = **60 mol $g^{-1}$**.

**A17.** m = 60/60 / 1 = 1 mol $kg^{-1}$. $\Delta T_b = 0.52 \times 1 = 0.52$ K. $T_b = 373.15 + 0.52$ = **373.67 K** (100.52 °C).

**A18.** $M = (1000\, K_f\, w_B) / (\Delta T_f\, w_A) = (1000 \times 1.86 \times 20)/(1.24 \times 500)$ = **60 mol $g^{-1}$**.

**A19.** Osmotic pressure is the pressure that must be applied to the solution side to stop solvent flow across a semi-permeable membrane. It is preferred for biomolecules because: (i) it is measured at room temperature, avoiding decomposition; (ii) the magnitude is large even for dilute solutions, giving good accuracy.

**A20.** (a) Gas solubility decreases with T; warm water holds less $O_2$. (b) Lower partial pressure of $O_2$ $\rightarrow$ less $O_2$ dissolves in blood (Henry). (c) Higher pressure increases $CO_2$ solubility, keeping the drink fizzy.

**A21.** i is the ratio (observed colligative property)/(calculated value assuming no association/dissociation). For AB $\rightarrow A^+ + B^-$, $i = 1 + \alpha$.

**A22.** Observed $\Delta T_f$ = 1.62 K. Molality m = (2/$M_{obs}$) / 0.025. i = $\Delta T_f/(K_f\ m_{theo})$ = 1.62 / (4.9 × (2/122)/0.025) = 0.504. For dimerisation $2A \rightarrow A_2$: i = 1 - $\alpha$/2; $\alpha$ = 2(1 - 0.504) = 0.992 $\rightarrow$ **~99.2%**.

**A23.** n($CH_3COOH$) = (0.6 × 1.06)/60 = 0.0106. m = 0.0106/1 = 0.0106 mol $kg^{-1}$. $\Delta T_f$(calc) = 1.86 × 0.0106 = 0.0197 K. i = 0.0205/0.0197 = 1.041. $\alpha$ = i - 1 = 0.041 = **4.1%** ($K_a \approx C\alpha^2 \approx 1.79\times10^{-5}$).

**A24.** n($K_2SO_4$) = $25\times10^{-3}/174 = 1.44\times10^{-4}$. C = $1.44\times10^{-4}/2 = 7.18\times10^{-5}$ mol $L^{-1}$. i = 3. $\pi$ = i C R T = 3 × $7.18\times10^{-5}$ × 0.0831 × 298 = **$5.33\times10^{-3}$ bar**.

**A25.** Applications: desalination of sea water by reverse osmosis; preservation of meat / fruits / pickles by adding salt or sugar; movement of water from soil into plant roots; intravenous fluids in hospitals must be isotonic with blood.

**End of Chapter 1.** Keep this note handy for revision. For any doubts, write to your faculty or visit Sai Wisdom Academy, Plot No 8, 3rd Main Road, Babu Nagar.