CHAPTER # 10

SOLUTIONS AND COLLOIDS

Q1. Define mixture with the help of example also write their properties.

Ans: Mixture:

The combination of difference substances is called mixture. Most of the substances we encounter more often in daily life are mixtures.

Examples:

For example milk, wood, air, gasoline etc. are mixtures of substances.

Note:

Such substances have properties different from those of the pure material they contain.

Characteristics:

i. Its composition is variable

ii. It retains some properties of its components.

Q2. Define different types of mixture with the help of example.

Ans: Following are the two types of mixture.

i. Homogeneous mixture

ii. Heterogeneous mixture

i. Homogeneous mixture:

A homogeneous mixture is simply any mixture that is uniform in composition throughout.

OR

The mixture whose properties are uniform throughout the liquid is called homogeneous mixture.

Example:

- Water itself is an example of a homogeneous mixture.
- A bottle of alcohol is a man-made homogeneous mixture.
- In the human body, blood plasma is an example of a homogeneous mixture. This is a colorless fluid that holds the blood cells in suspension.
- The bottle of vegetable oil so often used in cooking is a homogeneous mixture.
- Salt in water.

ii. Heterogeneous mixture:

The mixture whose properties are not uniform throughout the liquid is called heterogeneous mixture.

A heterogeneous mixture is made of different substances that remain physically separate. Heterogeneous mixtures always have more than one phase (regions with uniform composition and properties).

Example:

- A mixture of sand and sugar is a heterogeneous mixture.

- A glass of iced tea is a heterogeneous mixture.
- Blood is a heterogeneous mixture.

Q3. Define phase with the help of example.

Ans: Phase:

A sample of matter having a fixed composition and uniform properties throughout is called a **phase.**

Example:

For example pure sample of water under standard conditions exists as a single liquid phase. The properties of water e.g. **density, vapour pressure** etc are uniform throughout this liquid phase.

Q4. Define solution and also give its examples.

Ans: Solution:

A homogeneous mixture of two or more pure substances, which has uniform composition throughout is called solution.

Examples:

Some examples of solutions are salt water, rubbing alcohol, and sugar dissolved in water.

Q5. Why benzene and cyclohexane, oils etc. do not dissolve in water?

Ans: There are some substances that do not dissolve in water. For instance substances like benzene, cyclohexane, oils etc. do not dissolve in water.

Reason:

This is because these substances have non-polar molecules which do not have any interaction with water molecules.

Q6. How many types of molecules may exist in nature?

Ans: They are two types of molecules that exist in nature.

i. Hydrophilic molecules.
ii. Hydrophobic molecules.

i. Hydrophilic molecules:

Hydro mean water, philic means loving. So hydrophilic literally means water loving.

Molecules that are miscible with water are known as hydrophilic molecules. Such molecules can form hydrogen bond with water molecules.

Example:

For example, molecules of methanol, acetone, acetic acid etc. are called hydrophilic molecules.

ii. Hydrophobic molecules.

Phobic means disliking, therefore hydrophobic literally mean water disliking. **Molecules that do not dissolve in water are known as hydrophobic molecules**.

Example:

For example, molecules of organic fats and oils are called hydrophobic molecules.

Q7. What is critical solution temperature? Explain by giving the example of phenol-water system.

Ans: Phenol-water system:

i. When equal volumes of phenol and water are mixed into each other, two liquid layers are formed.

ii. The lower layer consists of a small amount of water dissolved in phenol, while the upper layer consists of small amount of phenol dissolved in water.

iii. It is observed that at 25°C upper layer is 5% solution of phenol in water and the lower layer is 30% water in phenol.

iv. This means water is solute in lower layer and phenol is solute in the upper layer.

v. Such solutions are called conjugate solutions.

vi. As the temperature increases, the mutual solubility of two liquids increases.

vii. Water starts moving from upper to the lower layer and phenol from lower layer to the upper layer thus composition of both the layers changes.

viii. When the temperature finally reaches 69.5°C, the composition of both the layers become identical. Each layer contains 34% phenol and 66% water.

ix. Above this temperature two solutions merge into one another and two liquids become completely miscible in all proportions.

Critical solution temperature:

The temperature at which two conjugate solutions merge into one another is called **critical solution temperature or upper consulate temperature.**

Examples:

Critical solution temperature for water-aniline system is 167°C with 15% water and that for aniline-hexane is 59.6 °C with 52% aniline. And that for water-phenol is 69.5°C with 34% phenol.

Q8. Explain the dissolution process and the factors affecting the dissolving rate.

Ans: Dissolution Process:

i. When a solute is added in a suitable solvent, it dissolves forming a solution. In the formation of a solution three types of interactions are involved. These are **solute - solute, solvent-solvent and solute-solvent interactions**.

ii. A solution forms only when the interactions between solute-solvent molecules are **equal to or greater** than the interactions between solute-solute and solvent-solvent molecules.

iii. This means, the process of dissolution can occur only when the interactions that hold together solute particles are weekend appreciably by the solute-solvent interactions.

Dissolving rate:

The speed with which a solute dissolves in a solvent is called dissolving rate.

Factors influence the rate:

Four factors influence the rate at which substances dissolve.

(i) Particle size **(ii)** Temperature

(iii) Solution Concentration **(iv)** Stirring.

(i) Particle size:

When a sugar cube in placed into water, it dissolves slowly than does an equal amount of finely granulated sugar. A sugar cube exposes less surface area to the water molecules than do the tiny particles of granulated sugar. Inner regions of cube dissolves only after the outer layers are in solution.

There are fewer unexposed particles when sugar is granulated, so it dissolves faster. Thus, to increase the rate of dissolving, large piece of solid must be grind into small particles.

(ii) Temperature:

Temperature is another factor which changes the rate at which solutes dissolve in a solvent. At higher temperature, increased molecular motion increases the interaction of solute and solvent particles which increases the dissolving rate.

(iii) Solution Concentration:

As the concentration of solute in solution increases, the time needed for more solute to dissolve increases. Initially the dissolving rate is maximum. When the first solute is added to the solvent.

With each addition of solute, the dissolving rate decreases until no more solute is observed to dissolve. At this stage solution is said to be saturated. Any addition of solute to a saturated solution remain un-dissolved and settle to the bottom of the container.

(iv) Stirring:

Stirring a solution increases the rate at which a solid dissolves decreasing the concentration of solute in the immediate region surrounding the solid solute. Stirring also increases the amount of exposed solute surface to the solvent.

Q9. Explain the salvation process in Ionic and Molecular Substances.

Ans: Solvation of Ionic and Molecular Substances:

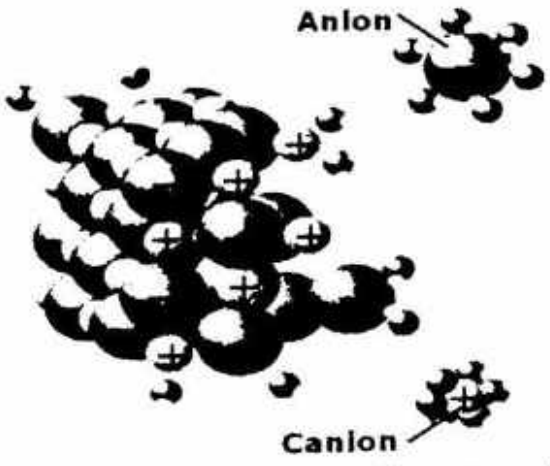
The process in which solvent molecules interact and surround solute ions or molecules is known as **solvation**. When water is the solvent, this process is known as **hydration**.

Ionic solids:

Ionic solids are soluble only in solvents having polar molecules. When an ionic compound is dissolved in a polar solvent, it splits up into its ions.

Example:

For instance, when NH_4NO_3 dissolves in water the resulting solution contains NH_4^+ and NO_3^- ions floating around independently. Solvent molecules surround these ions by directing their negative poles towards positive ions and their positive poles towards negative ions. This interaction is called as **ion-dipole interaction**. Circles with positive sign represent NH_4^+ ions and circles with negative sign represent NO_3^- ions.



Dissolving of ammonium nitrate crystal in water and hydration

Molecular solids:

Molecular solids are held together by dispersion forces, dipole-dipole forces and sometimes hydrogen bonds. Such solids dissolve readily in solvents with similar types of inter-molecular forces.

Example:

When a polar molecular substance is mixed with a polar solvent. Polar ends of solvent molecules interact with the opposite polar ends of molecules of solute and break attractive forces between them. Finally solvent molecules surround these molecules. This interaction is called dipole-dipole interaction.

Q10. Write down the daily life applications of heat of solution.

Ans: Daily Life Applications of Heat of Solution:

Instant hot and cold packs are in common use today.

Cold packs:

Cold packs are used for the treatment of injuries and reduction of swelling.

Hot packs:

Hot packs are used for instant warmth for hikers and skiers and treatment of pulled muscles.

These packs are excellent examples of basic science producing a technologically useful product.

Principle:

These packs are based on heat of solutions.

Construction:

i. These packs contain two separate compartments.

ii. One contains water and the other contains a salt, NH_4NO_3 for cold packs and $CaCl_2$ or $MgSO_4$ for hot packs.

iii. When required these packs are kneaded, the wall between the compartments breaks, allowing the salts to mix with water.

iv. Heat is absorbed in cold packs and released in hot packs.

v. Gradually these packs attained room temperature.

Q11. How can we determine percentage of water in the hydrate?

Ans: Percentage of water in a hydrate:

The percentage of water in a hydrate can be determined by adopting following steps.

i. For this purpose a known mass of hydrate is heated to expel water completely.

ii. Mass of anhydrous solid is determined.

iii. Difference in the two masses gives the mass of water present. From these masses, percentage of water in the hydrate is determined.

$$Percent\ water\ in\ a\ hydrate = \frac{Mass\ of\ water\ in\ the\ hydrate}{mass\ of\ hydrate} \times 100$$

Example:

An ion having higher charge density has greater ability to attract water molecules. In $CuSO_4.5H_2O$, the charge density of Cu^{++} ion in greater than that of SO_4^{-2} ion. Thus Cu^{++} ion has greater ability to attract water molecules than SO_4^{-2} ion. That is why, out of five water of hydration four are associated with Cu^{++} ion and

only one with SO_4^{-2} ion. In hydrates, although new bonds between ions and water molecules are formed but no hydrogen oxygen bond of water is broken.

Example 10.1:

250g of $CuSO_4\,xH_2O$ on heating produced 159.82g $CuSO_4$. Calculate the percent of water in the $CuSO_4,\,x\,H_2O$. Also determine the value of x.

Solution:

Mass of $CuSO_4\,xH_2O$ = 250g

Mass of $CuSO_4$ = 159.82g

Mass of water in the hydrate = 250 -159.82 = 90.18g

$$\text{Percent of water in the hydrate} = \frac{90.18\text{g}}{250\text{g}} \times 100 = 36.07\%$$

$$x = \text{No. of moles of water} = \frac{\textit{Mass of water in the hydrate}}{\textit{molar mass of water}}$$

$$= \frac{90.18\text{g}}{18\text{g/mole}}$$

$$x = 5.01$$

$$x = 5$$

Q12. Define following terms.

i. Unsaturated solution. ii. Saturated solution.

iii. Solubility

Ans: i. Unsaturated solution.

A solution which can dissolve more solute under existing conditions is called **unsaturated solution**.

ii. Saturated solution

A solution which contains maximum amount of dissolve solute under existing conditions is called **saturated solution**.

iii. Solubility

The maximum amount of solute that dissolve in the given quantity of a solvent under given conditions is called **solubility**.

Q13. Explain the concept of solubility with the help of examples.

Ans: Solubility:

The maximum amount of solute that dissolve in the given quantity of a solvent under given conditions is called **solubility**.

OR

Amount of a substance (called the solute) that dissolves in a unit volume of a liquid substance (called the solvent) to form a saturated solution under specified conditions of temperature and pressure is **called solubility.**

Unit:

Solubility is commonly expressed in number of grams of solute per cm^3 of solution or in terms of number of moles of solute per dm^3 of solution.

Explanation:

On molecular level a saturated solution is in equilibrium with excess of solute. At equilibrium the rate of solute going into the solution equals the rate of solute returning from the solution.

Un-dissolved solute $_{(s)} \rightleftharpoons$ Dissolve solute $_{(aq)}$

In a saturated solution dissolving of solute infact continues. At the same time many dissolved particles moving freely in solution strike the solid solute. Such particles are recaptured by the solid solute present at the bottom of the container. Such particles are thus re-crystallized. A stage is reached when a dynamic equilibrium is established between dissolved and undissolved solute.

Rule:

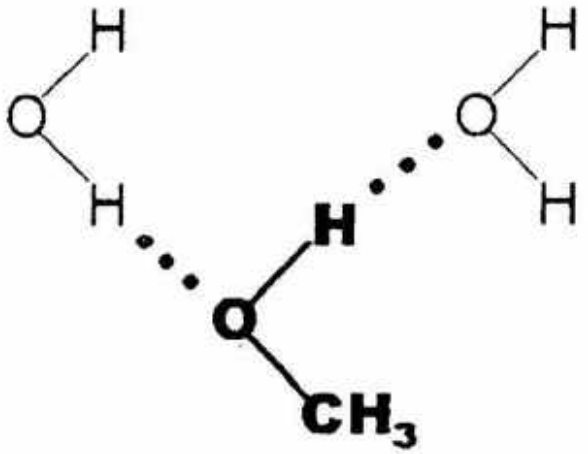
A general rule useful for predicting solubilities is **"like dissolves like."** Substances which have similar **structures and intermolecular** forces tend to be **soluble.**

While substances which have **dissimilar structures and intermolecular** forces are **insoluble.**

Examples:

Water and methanol:

For instance, water and methanol have similar structures and have hydrogen-bonding between their molecules. They can form **hydrogen bonds** with each other, when they are mixed. Thus water and methanol are **miscible**.



H-bonding between water and methanol

Benzene, C_6H_6 and carbon tetrachloride CCl_4:

Molecules of benzene, C_6H_6 and carbon tetrachloride CCl_4 are non-polar and have **London dispersion forces** between their molecules. When these two liquids are mixed, their molecules attract each other with London dispersion forces, thus these are **miscible liquids**.

Ionic solids:

Ionic solids have a crystal lattice structure composed of oppositely charged ions.

Examples:

i. When an ionic solid e.g. NaCl is placed in water which is a polar solvent. These ions are attracted by polar molecules. Water molecules break the crystal lattice of Na^+Cl^- and then surrounds the resulting Na^+ and Cl^- ions. These ions are called **hydrated ions**.

ii. When NaCl crystals are placed in CCl_4 or C_6H_6. Non-polar molecules of these liquids are unable to attract ions in Na^+Cl^- and cannot break apart the crystal lattice. Thus NaCl is **insoluble** in these solvents.

SELF-CHECK EXERCISE 10.1

Which solvent, liquid ammonia, NH_3 or benzene, C_6H_6 is more likely to dissolve each of the following solute. Give reasons?

(a) AgCl (b) Wax (c) H_2O (d) NH_4OH

Solution:

The concept of solubility is based on the principle **"Like dissolve like"** as NH_3 is a polar solvent. Therefore, it is more readily dissolve in polar substances like AgCl and NH_4OH.

Benzene C_6H_6 is a non-polar solvent. Therefore, it will dissolve non-polar substance, wax.

Q14. Why dissolving process of gases in water is always exothermic?

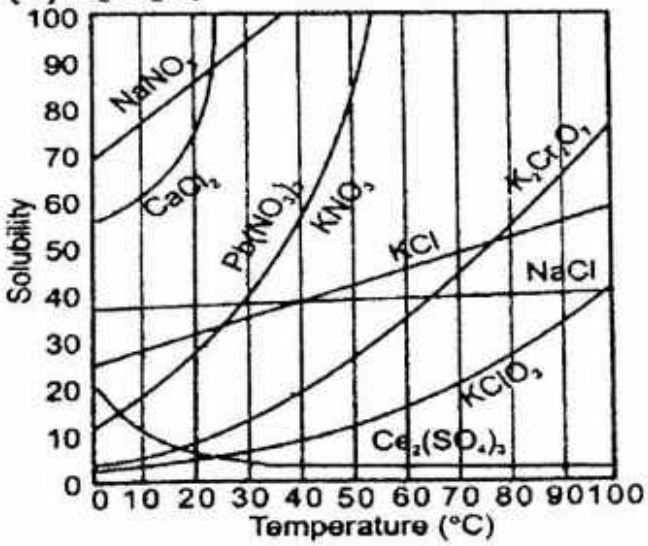
Ans: Water has open structure due to which little work is required to accommodate gas molecules, this means dissolving process is exothermic. **Thus solubilities of gases in water decrease with temperature**

SELF-CHECK EXERCISE 10.2

Consult solubility curve shown in figure 10.3 and explain the following

1. **$\Delta H_{solution}$ of $Ce_2(SO_4)_3$ and that of $KClO_3$ is positive or negative.**
2. **$\Delta H_{solution}$ of $CaCl_2$ is greater or lesser than that of $NaNO_3$ at 20°C.**
3. **Indicate the temperature at which solubilities of NaCl and KCl are same.**
4. **Which of the following salts has rapid increase in solubility with the increase in temperature:**

(a) $CaCl_2$ (b) $K_2Cr_2O_7$



Solution:

In exothermic reaction $\Delta H_{Solution} = negative$

In endothermic reaction $\Delta H_{Solution} = positave$

i. $\Delta H_{Solution}$ of $Ce_2(SO_4)_3$ and that of $KClO_3$ is positive or negative.

- The curve of $Ce_2(SO_4)_3$ decreases with increase in temperature, it shows that the solubility of $Ce_2(SO_4)_3$ decreases so it is an exothermic process and its $\Delta H_{Solution}$ **is negative**.
- $KClO_3$ has rising curve so its solubility increases with increase in temperature and hence, it is an endothermic process with positive $\Delta H_{Solution}$.

ii. $\Delta H_{Solution}$ of $CaCl_2$ is greater or lesser than that of $NaNO_3$ at 20°C.
The solubility curve shows that the solubility of $NaNO_3$ is greater than $CaCl_2$ at 20°C, so $\Delta H_{Solution}$ of $CaCl_2$ is lesser than that of $NaNO_3$ at 20°C.

iii. Indicate the temperature at which solubilities of NaCl and KCl are same.
At 40°C, their solubilities are same

iv. **Which of the following salts has rapid increase in solubility with the increase in temperature: (a) $CaCl_2$ (b) $K_2Cr_2O_7$**
$CaCl_2$ has more rising curve than $K_2Cr_2O_7$ thus $CaCl_2$ has rapid increase in solubility with the increase in temperature

Q15. Briefly describe the effect of pressure on solubility and also state Henry's law to explain your answer.

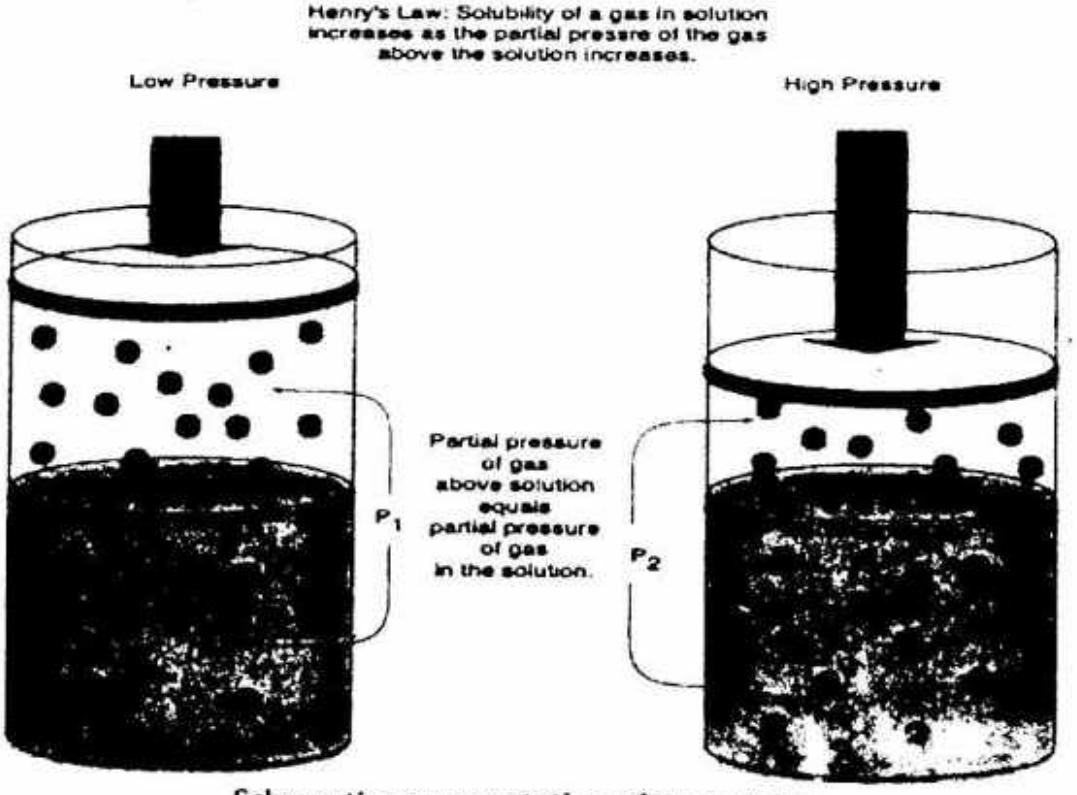
Ans: Effect of Pressure on Solubility:

Pressure has negligible effect on the solubilties of solids or liquids. On the other hand it does significantly **increase the solubility of a gas**.

Henry's Law:

At a given temperature, the solubilities of gases which do not react with the solvent are directly proportional to the partial pressures of the gases above the solution. This relationship is known as **Henry's Law**.

It is obeyed most accurately by dilute solutions of gases: Henry's law can be understood from Fig.



Schematic representation of Henry's law

Explanation:

Equilibrium between the gas above a liquid and the dissolved gas within the liquid is reached when the rates of evaporation and condensation of gas molecules become equal.

When the pressure in suddenly increased by pushing the piston, the number of molecules per unit volume increase in the gaseous state, this causes an increase in the rate at which the gas enter the solution, so the concentration of the dissolved gas increases.

The greater gas concentration in the solution causes an increase in the rate of evaporation, until a new equilibrium is reached.

Examples:

Carbonated drinks are bottled at high pressures of carbon dioxide. When the cap is removed, the effervescences results from the fact that the partial pressure of carbon dioxide in the atmosphere is much less than used in the bottling process. As a result, the equilibrium quickly shifts to one of the lower gas solubility. **This causes effervescences.**

Information:

We sometimes speak of dilute or concentrated solutions, but these terms have no precise meaning.

A dilute solution is one which contains relatively small amount of solute per unit volume of solution.

Whereas a concentrated solution is one which contains relatively greater amount of solute per unit volume of solution.

Example 10.2:

Calculate mass percent of a solution containing 10g sugar dissolved in 100g of water.

Solution:

$$Mass\ Percent = \frac{grams\ of\ sugar}{grams\ of\ solution} \times 100$$

$$= \frac{10g\ sugar}{10g\ sugar + 100\ g\ water} \times 100$$

$$= 9.09\%$$

Q16. Define percent weight by volume and also give its formula.

Ans: Percent weight by volume:

It is the mass of solute dissolved per 100 part by volume of solution. In this case volume of solvent is not exactly known.

$$\text{Percent w/v} = \frac{Mass\ of\ solute\ (g)}{vol.\ of\ solution\ (cm^3)} \times 100$$

For example, a 10% (w/v) NaCl solution contains 10g NaCl dissolved in 100 cm^3 of solution.

Q17. Define percent volume by volume and also give its formula.

Ans: Percent volume by volume:

It is the volume of solute dissolved per 100 parts by volume of solution. In such solutions volumes of solute and solvent may not be necessarily equal to the volume of solution.

$$\text{Percent v/v} = \frac{cm^3\ of\ solute}{cm^3\ of\ solution} \times 100$$

For example a 5% (v/v) Ethanol solution means 5cm^3 ethanol dissolved per 100 cm^3 of solution.

Q18. Define percent volume by weight and also give its formula.

Ans: Percent volume by weight:

The volume of a solute dissolved per 100g of solution is called percent volume by weight. In this case total volume of solution is not known.

For example. A 5% (v/w) Ethanol contains 5 cm^3 ethanol per 100g of solution.

$$\%\ \text{v/w} = \frac{volume\ of\ solute\ (cm^3)}{Mass\ of\ solution\ (g)} \times 100$$

Q19. Differentiate between molarity and molality.

Ans: Difference between molarity and molality:

Molarity (M)	Molality (m)
i. It is defined as the number of moles of solute dissolved per dm^3 of solution.	**i.** It is defined as the number of moles of solute dissolved per kilogram of solvent.
ii. Molarity is moles of solute per liter of solution.	**ii.** Molality is moles of solute per kilogram of solvent.
iii. Molarity is expressed as the moles per **Liter/s of solution.**	**iii.** Molality is moles per **Kilogram/s of solvent.**
iv. $M = \frac{\text{moles of solute}}{\text{dm}^3\text{of solution}}$ $M = \frac{\text{grams of solute}}{\text{molar mass of solute} \times \text{dm}^3\text{ of solution}}$	**iv.** $m = \frac{\text{moles of solute}}{\text{kg. of solvent}}$
Example: If 0.5 mole of NaOH (20g) is dissolved in enough water to make one dm^3 of solution, 0.5 molar or 0.5 M NaOH solution is obtained.	**Example:** When 58.5g NaCl (1 mole) is dissolved in one kilogram of water, the resulting solution would be one molal or 1m NaCl solution.

Example 10.3:

What is the molarity of a solution of 0.25g of $NaHCO_3$ in 100 cm^3 of solution?

Solution:

Mass of $NaHCO_3$ = 0.25g

Molar mass of $NaHCO_3$ = 23 + 1 + 16 × 3

= 72g / mole

Volume of solution = 100 cm^3

= 100/1000 = 0.1 dm^3

$$\text{Molarity} = \frac{grams\ of\ NaHCO_3}{molar\ mass\ of\ NaHCO_3 \times dm^3\ of\ solution}$$

$$= \frac{0.25g}{72g/mole \times 0.1\ dm^3} = 34.72\ M$$

Example 10.4:

Ethanol is an excellent organic solvent. It is used to prepare tinctures and in the extraction of medicinal compounds from plants. For this purpose either pure ethanol or its aqueous solutions are used. A

solution is prepared by mixing 1.00g of ethanol (C_2H_5OH) with 100g of water. Calculate molality of this solution.

Solution:

Mass of ethanol	=	1.00 g
Molar mass of ethanol	=	$12 \times 2 + 1 \times 6 + 16 = 46$ g/mole
Moles of ethanol	=	$\frac{1.00g}{46g/mole}$
	=	2.17×10^{-2} mole
Mass of water	=	100g = 100/1000 = 0.1 Kg
Molality of solution	=	moles of C_2H_5OH / Kg of water
	=	2.17×10^{-2} mole / 0.1 kg
	=	0.217 m.

Notice that molality is independent of temperature whereas molarity depends on temperature. This is because molarity is based upon volume of solution. As the temperature increases, the amount of solute remains constant but the volume of solution increases slightly, thus molarity decreases slightly.

SELF-CHECK EXERCISE 10.3

Any fluids infused intravenously into an individual must be isotonic with the blood cells and the blood plasma. Such infusions are either 5% dextrose (glucose) or 9% normal saline. The first solution is composed of 5.0 gram of glucose per 100 cm^3 of solution and the other of 9.0 gram NaCl per 100 cm^3 of solution. Calculate the molarity of these intravenous solutions. (Ans: 0.28M, 0.15M)

(i) 5% dextrose means 5g glucose in 100 cm^3 solution.

$$Mass\ of\ glucose\ (C_6H_{12}O_6) = 5\ g$$

$$Molecular\ Mass\ of\ glucose\ (C_6H_{12}O_6) = 12 \times 6 + 1 \times 12 + 16 \times 6 = 180\ g\ mole^{-1}$$

$$Volume\ of\ solution = V = 100\ cm^3 = \frac{100}{1000} = 0.1\ dm^3$$

$$Molarity = \frac{Mass\ of\ glucose\ (C_6H_{12}O_6)}{Molecular\ Mass\ of\ glucose\ (C_6H_{12}O_6)} \times \frac{1}{V(dm^3)}$$

As

$$Molarity = \frac{5}{180} \times \frac{1}{0.1} = 0.28\ M$$

(ii) 9% norma/ saline; 9g NaCl in 100 cm^3 solution.

$$Mass\ of\ NaCl = 9\ g$$

$$Molar\ Mass\ of\ NaCl = 23 + 35.5 = 58.5\ g\ mole^{-1}$$

$$Volume\ of\ solution = V = 100\ cm^3 = \frac{100}{1000} = 0.1\ dm^3$$

$$Molarity = \frac{Mass\ of\ NaCl}{Molar\ Mass\ of\ NaCl} \times \frac{1}{V(dm^3)}$$

$$Molarity = \frac{9}{58.5} \times \frac{1}{0.1} = 1.54\ M$$

Example 10.5:

An aqueous solution containing 100 g ethanol per dm^3 of solution has a density of 0.984 g cm^{-3}. Calculate mole fraction of each component of solution.

Solution:

Volume of solution = 1dm^3 = 1000 cm^3

Mass of ethanol = 100g

Mass of solution = Density × Vol of solution of solution

= 0.984 of cm^{-3}× 1000 cm^3

= 984g

Mass of water (solvent) = 984g – 100g = 884g

No of moles of H_2O = $\frac{884g}{18g/mole}$ = 49.1 mole

No. of moles of ethanol = $\frac{100g}{46\,g/mole}$ = 2.17 mole

Total No. of moles in solution = 49.1 + 2.17 = 51.27 mole

X_{H_2O} = $\frac{49.1\ mole}{51.27\ mole}$ = 0.958

$X_{ethanol}$ = $\frac{2.17\ mole}{51.27\ mole}$ = 0.042

Example 10.6:

An atmospheric chemist reports that one dm^3 of air in an urban area contained 3.5×10^{-4} cm^3 of CO. What was the concentration of CO in ppm?

$$ppm = \frac{Volume\ of\ CO\ (cm^3)}{Volume\ of\ air\ (cm^3)} \times 10^6$$

$$= \frac{3.5\times10^{-4}\ cm^3}{10^3\ cm^3} \times 10^6$$

= 0.35 ppm.

Example 10.7:

If the concentration of ozone in atmosphere reaches 0.5 ppb. What mass of ozone would be present per kg of air.

Solution:

$$ppb = \frac{Mass\ of\ ozone\ (g)}{Mass\ of\ air\ (g)} \times 10^9$$

$$0.5 = \frac{Mass\ of\ ozone}{1000\ g} \times 10^9$$

$$Mass\ of\ ozone = \frac{0.5 \times 1000}{10^9}$$
$$= 0.5 \times 10^3 \times 10^{-9}$$
$$= 5 \times 10^{-7} g$$

Note: ppb: parts per billion

SELF-CHECK EXERCISE 10.4

1. **Calculate the ppm by mass of calcium in a 2.5g tablet that contains 500mg calcium. (Ans: 0.2 ppm)**
2. **Waste water from a cement factory contains 22mg Ca^{+2} ions and 0.006g Mg^{+2} ions per Kg of solution. Calculate the concentration of both these ions in ppm. (Ans: 0.0000229, 0.000006)**
3. **Gold occurs in sea water at an average concentration of 1.09×10^{-2} ppb. How many Kg of water must be processed to recover 1gm of Gold. (Ans: 9.17×10^3 Kg)**

Solution:

1. $Mass\ of\ calcium = 500mg = \frac{500}{1000} = 0.5g$

$Mass\ of\ tablet = 2.5g$

$$ppm = \frac{Mass\ of\ calcium}{Mass\ of\ tablet} \times 10^6$$

$$ppm = \frac{0.5\ g}{2.5\ g} \times 10^6 = 0.2 \times 10^6\ ppm.$$

2. $Mass\ of\ solution = 1\ kg = 1000\ g$

$$Mass\ of\ Ca^{2+} = 22mg = \frac{22}{1000} = 0.022g$$

$$Mass\ of\ Mg^{2+} = 0.006g$$

$$Concentration\ of\ Ca^{2+}\ ions\ In\ ppm = \frac{Mass\ of\ Ca^{2+}}{Mass\ of\ solution} \times 10^6$$
$$= \frac{0.22\ g}{1000\ g} \times 10^6 = 22\ ppm$$

$$Concentration\ of\ Mg^{2+}\ ions\ in\ ppm = \frac{Mass\ of\ Mg^{2+}}{Mass\ of\ solution} \times 10^6$$
$$ppm = \frac{0.006\ g}{1000\ g} \times 10^6 = 6\ ppm$$

3. $Concentration\ of\ gold\ in\ sea\ water = 1.09 \times 10^{-2}\ ppb$

$Mass\ of\ gold\ to\ recover = 1\ g$

$Mass\ of\ water = ?$

$$ppb = \frac{Mass\ of\ gold\ (g)}{Mass\ of\ water\ (g)} \times 10^9$$

$$1.09 \times 10^{-2} = \frac{1\ g}{Mass\ of\ water\ (g)} \times 10^9$$

$$or \quad Mass\ of\ water = \frac{1\,g}{1.09 \times 10^{-2}} \times 10^9$$
$$= 0.917 \times 10^2 \times 10^9\,g$$
$$= 0.917 \times 10^{11}\,g$$
$$= 9.17 \times 10^{10}\,g$$
$$= 9.17 \times 10^7\,kg$$

Example 10.8:

Calculate the molality of 15% (w/w) glucose ($C_6H_{12}O_6$) solution.

Solution:

Since solution of glucose is 15% w/w

Mass of glucose = 15g

Mass of solution = 100g

Mass of water = 100g – 15g = 85g.

= 85/1000 = 0.085 kg

Molar mass of glucose = $12 \times 6 + 1 \times 12 + 16 \times 6$ = 180 g / mole

$$\text{Molality} = \frac{\text{Mass of solute}_{(g)}}{\text{Molar mass of solute g/mole} \times \text{Mass of solvent(kg)}}$$

$$\text{Molality} = \frac{150g}{180\ g/mole \times 0.085kg}$$

Molality = 0.98 moles per kg.

Molality = 0.98 m

Example 10.9:

Sulphuric acid is known as king of chemicals it is used in the manufacture of many chemicals, drugs, dies, plastics, paints, disinfectants, explosives, synthetic fibres etc. It is prepared commercially by contact process and is normally 98% by weight. If its density is 1.84g cm^{-3}, what is its molarity.

Solution:

Molar Mass of H_2SO_4 = 1×2+32×1+16×4 = 98 g/mole

Percentage of H_2SO_4 solution = 98%

Density of H_2SO_4 solution = 1.84 gcm^{-3}

1 cm^3 of H_2SO_4 solution contains 1.84g H_2SO_4 solution

1000 cm^3g H_2SO_4 solution would have = 1.84×1000 = 1840g H_2SO_4 solution

Of this total mass 98% is H_2SO_4

Mass of H_2SO_4 in solution = $1840 \times 98 / 100$ = 1803.2g

Moles of H_2SO_4 in solution = 1803.2g / 98 g/mole = 18.4 moles

Since solution has 18.4 moles of H_2SO_4 per 1000 cm^3 or 1dm^3 of solution, the molarity of this solution is 18.4 M.

Example 10.10:

Commercial HCl is 12 molar (density = 1.17 g cm^{-3}).

Calculate the mass percent of HCl in this solution.

Solution:

Molarity of HCl = 12 M

As d = m / v.

m = d × v

Thus mass of HCl solution present in its 1000 cm^3 (1dm^3) would be

1.17 g / cm^3 × 1000 = 1170g

Molar mass of HCl = 1 + 35.5 = 36.5g/mole

As HCl solution is 12M

Mass of HCl in solution = 12 × 36.5

= 438g

Mass percent of HCl = $\frac{438g}{1170g}$

= 37.44%

Often more concentrated solutions are diluted to obtain less concentrated solutions. Such dilutions are done after calculating the amount of water that is mixed with the more concentrated solution to give the desired lower concentration.

Example 10.11:

Sodium hydroxide solutions are used to neutralize acids, to treat cellulose in the preparation of rayon and to remove potato peals in commercial use. 250 cm^3 of 2M NaOH is mixed with 250 cm^3 of water. Calculate the molarity of resulting solution.

Solution:

250 Cm^3 of 2M NaOH contain = $\frac{250cm^3 \times 2}{1000cm^3}$ = 0.5 moles of NaOH

Total volume of solution = 250 cm^3 + 250 cm^3

= 500 cm^3 = 0.5 dm^3

Molarity of resulting solution = 0.5 moles / 0.5 dm^3 = 1M

SELF-CHECK EXERCISE 10.5

1. **Urea (NH_2CONH_2) is a white solid produced commercially as a fertilizer and starting material for plastics. It is a non-volatile and non-ionizable compound. What is the mole fraction of Urea in an aqueous solution that is 0.25m Urea. (Ans: 4.49×10^{-3})**
2. **Potassium hydroxide (KOH) is used in the manufacture of liquid soaps, in paints and varnish removers. An experimenter needs 25cm^3 of 0.015M KOH. What mass of KOH, he will dissolve to make this solution. (Ans: 0.021 g)**

Solution:

1. $Molality\ of\ urea = 0.25\ m$

$Number\ of\ moles\ of\ urea = n_2 = 0.25\ moles$

$Mass\ of\ water = n_1 = 1\ kg = 1000\ g$

$Molecular\ mass\ of\ water = 18\ g\ mole^{-1}$

$Number\ of\ moles\ of\ water = n_1 = \frac{1000}{18} = 55.56\ moles$

$Mole\ fraction\ of\ urea = X_2 = ?$

$$X_2 = \frac{n_2}{n_1 + n_2}$$

$$Or \qquad X_2 = \frac{0.25}{55.56 + 0.25}$$

$$X_2 = 4.48 \times 10^{-3}$$

2. $Volume\ of\ solution = 25\ cm^3 = 0.025\ dm^3$

$Molarity\ of\ solution = M = 0.015M$

$Molar\ Mass\ of\ KOH = 39 + 16 + 1 = 56\ g\ mole^{-1}$

$Mass\ of\ KOH = ?$

As

$$Molarity = \frac{Mass\ of\ KOH}{Molar\ Mass\ of\ KOH} \times \frac{1}{Volume\ of\ solution\ in\ dm^3}$$

$$Mass\ of\ KOH = Molarity \times Molar\ Mass\ of\ KOH \times Volume\ of\ solution\ in\ dm^3$$

$$Mass\ of\ KOH = 0.015 \times 56 \times 0.025 = 0.021\ g$$

Q20. State and derive Raoult's law.

Ans: Raoult's Law:

A French chemist Raoult discovered a quantitative relationship between vapour pressure of solution and concentration of solution. This relationship is known as Raoult's law.

Statement:

It states that when a nonvolatile solute is dissolved in a solvent, the vapour pressure of solution (P) is directly proportional to the mole fraction of solvent (X_1)

Mathematically:

$P \propto X_1$

$P = P^\circ X_1$ ---------- (1)

Where P° is constant of proportionality and is vapour pressure of the pure solvent. A more common form of Raoult's law is obtained by simple substitution. In a binary solution $X_1 + X_2 = 1$

And $X_1 = 1 - X_2$ Where X_2 is mole fraction of solute.

Substituting value of X_1 in equation (1), we get

$P = P^\circ (1 - X_2)$

$P = P^\circ - P^\circ X_2$

$P - P^\circ = - P^\circ X_2$

$P^\circ - P = P^\circ X_2$ ----------(2)

$\Delta P = P^\circ X_2$

Where $P^\circ - P$ = lowering in vapour of solvent (ΔP).

Therefore Raoult's law can also be stated as the lowering in vapour pressure is directly proportional to the mole fraction of solute.

Rearranging equation (2) we get another form of Raoult's law.

$$\frac{P^\circ - P}{P^\circ} = X_2$$

$$\frac{\Delta P}{P^o}$$

$\frac{P^o - P}{P}$ or $\frac{\Delta P}{P^o}$ is relative lowering in vapour pressure of the solvent.

Raoult's law can also be stated as the relative lowering in vapour pressure is equal to the mole fraction of solute.

Note: Lowering in vapour pressure depends on temperature whereas relative lowering in vapour pressure is independent of temperature.

Roult's law for ideal solution of two volatile components (liquids):

i. Raoult's Law also governs the vapor pressure of mixtures of two or more volatile components.

ii. The vapor pressure of each is determined by its mole fraction and the vapor pressure of the pure component.

iii. The total vapor pressure always lies between the vapor pressures of the pure components, and is determined by the mixture composition

Example 10.12:

What are the partial pressures of benzene and toluene in a solution in which the mole fraction of benzene is 0.6? What is the total vapor pressure? The vapor pressure of pure benzene is 95.1 mm Hg and the vapor pressure of pure toluene 28.4 mm Hg at 25°C.

Solution: If $x_{benzene} = 0.6$, than $x_{toluene} = 0.4$ because 1 - 0.6 = 0.4.

Now that we know the mole fractions and vapor pressures.

$P_{benzene} = x_{benzene}P_{benzene} = (0.6)(95.1 \text{ mm Hg}) = 57.1 \text{ mm Hg}$

$P_{toluene} = x_{toluene}P_{toluene} = (0.4)(28.4 \text{ mm Hg}) = 11.4 \text{ mm Hg}$

The total vapor pressure is simply the sum of the partial pressures:

$P_{total} = P_{benzene} + P_{tuolene} = 57.1 \text{ mm Hg} + 11.4 \text{ mm Hg} = 68.5 \text{ mm Hg}$

Q21. Define colligative properties and also write down the names of colligative properties.

Ans: Colligative Properties of Solutions Containing Non-Electrolyte Solutes:

There are certain properties of solutions depend only on number of solute particles and not on their nature. These properties are called as **colligative properties**.

Explanation:

Colligative properties are properties of solutions that depend upon the ratio of the number of solute particles to the number of solvent molecules in a solution, and not on the type of chemical species present.

The word colligative means **"to collect"**. Four such related properties are:

(i) Vapour pressure lowering
(ii) Boiling point elevation
(iii) Freezing point depression
(iv) Osmotic pressure.

Q22. Derive equation for lowering of vapour pressure from Raoult's law and how molar mass of a nonvolatile solute can be calculated.

Ans: Lowering of Vapour Pressure:

When a non-volatile solute is dissolved in a solvent, the escaping tendency of solvent molecules from the surface of the solution decreases. Thus its vapour pressure is lowered.

Mathematically:

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

$$\frac{\Delta P}{P^{\circ}} = X_2$$

If n_1 and n_2 are the number of moles of the solvent and solute respectively, then

$$X_2 = \frac{n_2}{n_1 + n_2}$$

So $$\frac{\Delta P}{P^{\circ}} = \frac{n_2}{n_1 + n_2}$$

Since for dilute solutions $n_2 << n_1$, n_2 can be ignored in the denominator

$$\frac{\Delta P}{P^{\circ}} = \frac{n_2}{n_1}$$

If W_1 and W_2 are masses of solvent and solute while M_1 and M_2 are their molecular masses respectively, then

$$n_1 = \frac{W_1}{M_1} \quad \text{and} \quad n_2 = \frac{W_2}{M_2}$$

So $$\frac{\Delta P}{P^{\circ}} = \frac{W_2\ M_2}{W_1\ M_1}$$

$$\frac{\Delta P}{P^{\circ}} = \frac{W_2 x M_1}{M_2 x W_1}$$

Or $$M_2 = \frac{W_2 \times M_1 \times P^{\circ}}{W_1 \times \Delta P}$$

Conclusion:

Thus molecular mass of a non-volatile solute can be calculated from this equation.

Example 10.12:

When 106.3g of an organic compound M dissolved in 863.5g of benzene, the vapour pressure of benzene is lowered from 98.6 to 86.7torr. Calculate the molecular mass of M.

Solution:

P° = 98.6 torr

P = 86.7 torr

$\Delta P = P^{\circ} - P$

$= 98.6 - 86.7 = 11.9$ torr

W_2 = 106.3g

W_1 = 863.5g

$$M_1 = 78\text{g/mole}$$

$$M_2 = \frac{W_2 \times M_1 \times P^\circ}{W_1 \times \Delta P} = \frac{106.3\text{g} \times 78\text{g/mole} \times 98.6\text{torr}}{863.5\text{g} \times 11.9\text{torr}}$$

$$= 79.5\text{ g/mole}$$

Q23. What are the causes of boiling point elevation and freezing point depression?

Ans: Causes of Boiling Point Elevation:

i. When a non-volatile and non-electrolyte solute is added to a solvent, its vapour pressure is decreased.

ii. This is because in solution both solute and solvent particles occur on the surface.

iii. Thus solute particles decrease the number of solvent surface particles.

iv. This decreases the rate of evaporation of solvent, which decreases the vapour pressure.

v. Therefore, a solution must be heated to a higher temperature than the boiling point of pure solvent to equalize its vapour pressure to the atmospheric pressure.

Conclusion:

Thus addition of solute to a pure solvent causes an elevation of the boiling point of solution.

Causes of Freezing Point Depression:

i. The decrease in vapour pressure of a pure solvent on the addition of a solute also affects the freezing point of the solution.

ii. The solution will freeze at a temperature at which vapour pressure of both solution and solid solvent are the same.

iii. This means solution will freeze at a lower temperature than that of the pure solvent.

Conclusion:

Thus addition of a non-volatile solute also causes a decrease or depression in freezing point of solution.

Q24. Give quantiative and graphical explanation of boiling point elevation.

Ans: Quantiative Aspects of Boiling Point Elevation:

When one mole of non-electrolyte and nonvolatile solute is dissolved in 1 kg of water boiling point is raised 0.52°C. This value is known as the molal boiling point elevation constant (ebullioscopic constant) for water.

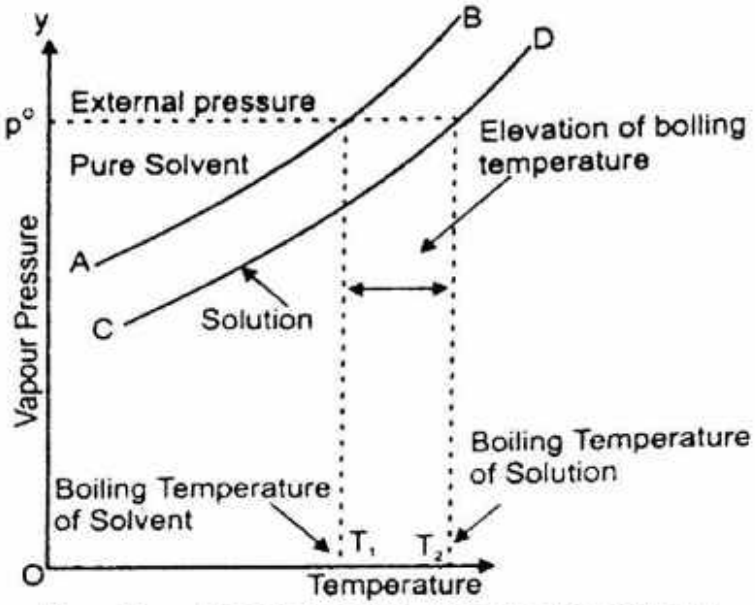
Graphically:

I. Figure shows how the vapour pressure influences the boiling point.

II. Curve AB represents variation in vapour pressure of pure solvent with temperature.

III. The solvent boils at temperature T_1 when its vapour pressure becomes equal to the external pressure P°.

Iv. The curve CD represents variation in vapour pressure of solution with temperature.



Elevation of boiling point – temperature curve

v. This curve must lie below that of pure solvent.

vi. This is because the vapour pressure of solution at all temperatures is lower than that of the pure solvent.

vii. The solution will boil at higher temperature T_2 to equalize its vapour pressure to external pressure P°.

Mathematically:

The difference of two boiling points gives the elevation of boiling point (ΔT_b)

$$\Delta T_b = T_2 - T_1$$

The magnitude of the boiling point elevation is directly proportional to the molality of solution.

$$\Delta T_b \propto m$$

$$\Delta T_b = K_b m \quad \ldots\ldots\ldots\ldots\ldots\ldots (1)$$

The constant K_b is called molal boiling point elevation constant or ebullioscopic constant. For a 1 molal solution.

$$\Delta T_b = K_b$$

Thus elevation of boiling point when 1 mole of non-volatile, non-electrolyte solute is dissolved in 1 kg of solvent is called molal boiling point elevation constant.

The value of K_b depends only on the nature of the solvent and is **independent** of nature of solute used. The nature of solute does not affect boiling point of elevation as long as the solute **does not ionize.**

The molality m of the solution containing W_2g of solute of molecular mass M_2 dissolved in W_1g of the solvent is given by

$$m = \frac{W_2 \times 1000}{M_2 \times W_1}$$

Substituting the value of m in eq (1) we have

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

This equation is used to determine the molecular mass of solute.

SELF-CHECK EXERCISE 10.6

Glycerol ($C_3H_8O_3$) is a syrupy, sweat tasting liquid used in cosmetics and candy. It is a non-volatile and non-electrolyte compound. What is the freezing point of an aqueous solution that is 0.25m glycerol?
(Ans: 0.465°C)

Solution:

$$Molality\ of\ solution\ = 0.25m$$
$$K_f\ for\ water\ =\ K_f\ =\ 1.86^\circ\ C$$
$$\Delta T_f\ =\ K_f \times m\ =\ 1.86 \times 0.25\ =\ 0.465^\circ\ C$$

Let,

$$Freezing\ point\ of\ pure\ water\ =\ T_1\ =\ 0^\circ\ C$$
$$Freezing\ point\ of\ solution\ \ =\ T_2\ =?$$
$$\Delta T_f\ =\ T_1 - T_2$$
$$T_2\ =\ T_1 - \Delta T_f$$
$$=\ 0 - 0.465 = -0.465^\circ\ C$$

Q25. Describe Lands Berger's method to measure boiling point elevation.

Ans: Lands Berger's Method:

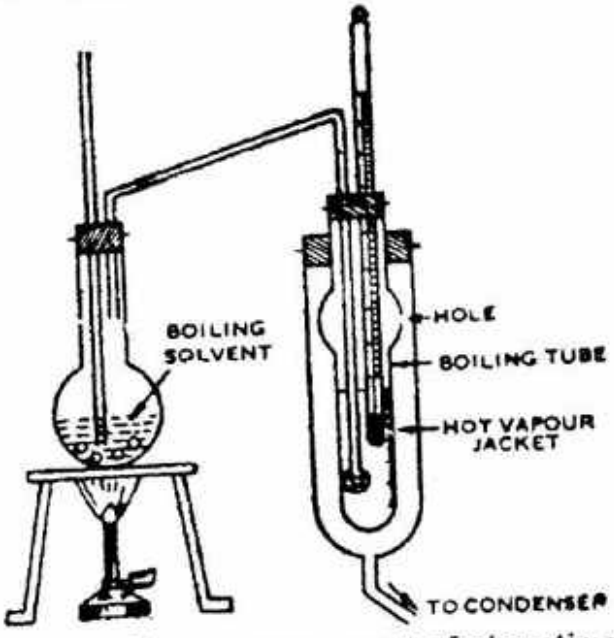
The apparatus used is shown in Fig. It consists of four major parts.

(a) A graduated inner tube with a hole in its side.

(b) A boiling flask which sends the solvent vapours into the inner tube though a rosehead.

(c) An outer tube, which receives solvent vapours coming out from the side hole of the inner tube.

(d) A Beckmann thermometer which can read upto 0.01K



Landsbeger's apparatus for measurement of elevation of boiling points.

Activity

Cholesterol is an important compound in our body. Its excess has been implicated as a cause of heart disease. Take 100 cm^3 of pure benzene (solvent) in the inner tube. Boil benzene in the boiling flask and pass its vapours through the benzene in the inner flask. These vapours will boil benzene in the inner tube by its latent heat of condensation. Record the temperature at which benzene is boiled in the inner tube. You will observe that benzene will boil at 5.5°C. Now stop the supply of vapours temporarily. Drop 4.5g of cholesterol in the inner tube. Pass vapours of the benzene from boiling flask again to boil the solution. Record boiling point of the solution. Thermometer will show 6.07 °C. Now stop the supply of benzene vapours. Remove thermometer and rose head from the inner tube and note the volume of solution. It will be 121.5 cm^3 Determine density of this solution. Its density will be 0.897g cm^{-3} From the volume and density determine the mass of solution. From the mass of solution determine the mass of benzene. From this data calculate the molecular mass of Cholesterol.

Solution:

Mass of Cholesterol = $W_2 = 4.5$ g
Mass of Benzene = $W_1 = ?$
Boiling Point of Benzene = $T_1 = 5.5°C$
Boiling Point of Solution = $T_2 = 6.07°C$
K_b for Benzene = 5.12
Volume of solution = $V = 121.5\ cm^3$
Density of solution = $d = 0.897\ g\ cm^{-3}$
Molecular mass of Cholesterol = $M_2 = ?$

As, $d = \frac{m}{V}$

$M = d \times V$

Mass of Solution = $0.897 \times 121.5 = 108.99g$

Mass of Benzene (W_1) = Mass of solution – Mass of Cholesterol

$W_1 = 108.99 - 4.5 = 104.486g$

As, $\Delta T_b = T_2 - T_1$

$\Delta T_b = 6.07 - 5.5 = 0.57°C$

Now

$$M_2 = \frac{K_b \times W_2 \times 1000}{\Delta Tb \times W_1} = \frac{4.5 \times 5.12 \times 1000}{0.57 \times 104.486} = 386.86 g\ mole^{-1}$$

SELF-CHECK EXERCISE 10.7

Calculate the boiling point of a solution containing 12.5 g of benzoic acid, $C_7H_6O_2$ in 110 g of benzene. Boiling point and K_b of benzene are 80.1°C and 2.53 respectively. (Ans: 82.637°C)

Solution:

$Mass\ of\ benzoic\ acid = 12.5\ g$

$Molecular\ mass\ of\ benzoic\ acid\ (C_7H_6O_2) = 12 \times 7 + 1 \times 6 + 16 \times 2$

$= 122\ g\ mole^{-1}$

$Mass\ of\ Benzene = 110\ g$

$Boiling\ Point\ of\ Benzene = T_1 = 80.1\ ^\circ C$

$$Boiling\ Point\ of\ Solution\ = T_2 = ?$$
$$K_b\ for\ Benzene\ = 2.53$$

As we know that,

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

$$\Delta T_b = \frac{2.53 \times 12.5 \times 1000}{122 \times 110} = 2.357^\circ C$$

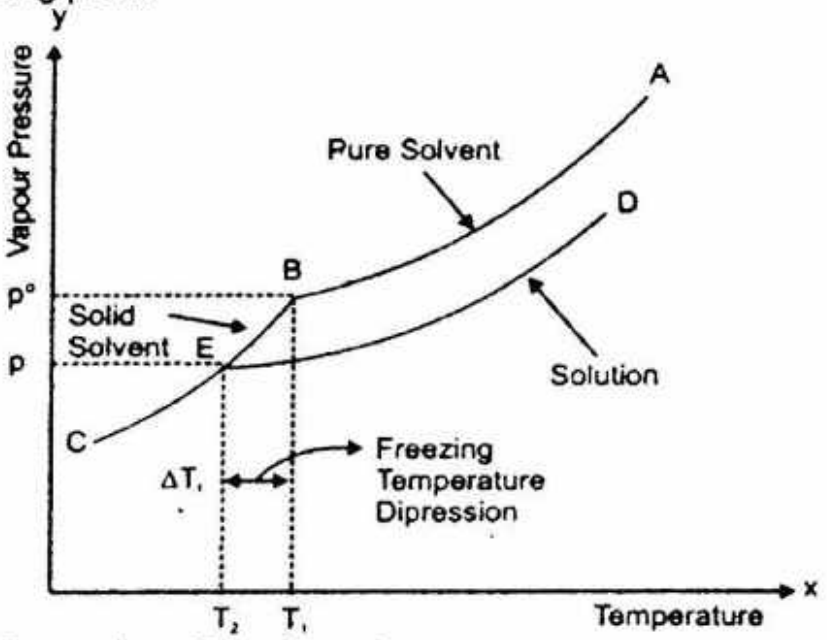
Thus,

$$\Delta T_b = T_2 - T_1$$
$$T_2 = T_1 + \Delta T_b$$
$$= 80.1 + 2.357 = 82.457^\circ C$$

Q26. Write quantitative and graphical explanation of freezing point depression for a solution.

Ans: Quantitative Aspects of Freezing Point Elevation:

The difference in the freezing points of pure solvent and solution is called the depression of freezing point.



Depression of freezing point – temperature curve

Graphically:

Figure shows how the vapour pressure influences the freezing point.

i. Curve ABC is for the pure solvent.

ii. The solvent freezes at temperature T_1 corresponding to the point B when its vapour pressure is P°.

iii. The portion of the cure BC is for the solid solvent.

iv. Greater slope of curve BC indicates a rapid change of vapour pressure with temperature.

v. The curve DEC is for the solution.

vi. DEC meets curve BC at point E which is freezing point of solution T_2 when its vapour pressure is P which is lower than P°.

vii. This is because vapour pressure of solution is always less than that if the pure solvent.

Mathematically:

The Difference between the two freezing points gives the depression of the freezing point ΔT_f.

$$\Delta T_f = T_1 - T_2$$

Experimentally $\Delta T_f \propto m$

$$\Delta T_f = K_f m \text{ ---------- (1)}$$

Where K_f is constant called the molal freezing point constant or the cryoscopic constant. For I molal solution.

$$\Delta T_f = K_f$$

Thus the depression of freezing point when I mole of non-volatile, non-electrolyte solute is dissolved in 1 kg of solvent is called **molal freezing point depression constant.**

The value of K_f depends upon the nature of the solvents and is independent of the solute used.

The nature of solute does not affect the freezing point depression as long as the solute doesn't ionize.

The molality m of the solution containing w_2g of solute of molecular mass M_2 dissolved in W_1 of the solvent is given by

$$m = \frac{W_2 \times 1000}{M_2 \times W_1}$$

Substituting the value of m in aq (1)

We have

$$\Delta T_f = \frac{K_f \times W_2 \times 1000}{M_2 \times W_1}$$

Rearranging $M_2 = \frac{K_f \times W_2 \times 1000}{\Delta T_f \times W_1}$

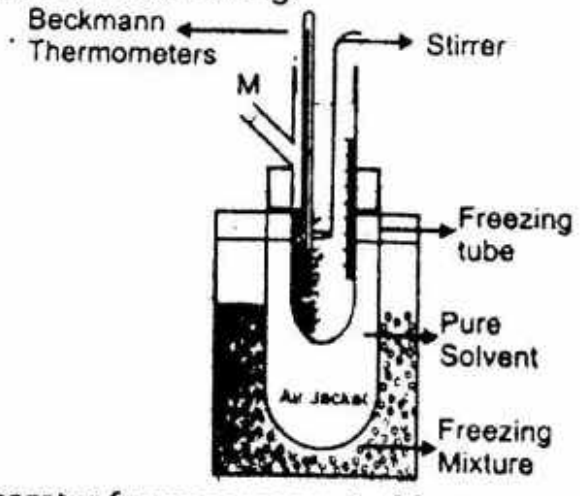
This equation is used to determine the molecular mass of solute.

Q27. Describe the method use for the measurement of freezing point depression.

Ans: Measurement of Freezing Point Depression:

Beckmann's Method:

The apparatus used is shown in Fig.



Beckmann's apparatus for measurement of depression of freezing point.

It consists of four parts.

(a) An inner freezing tube with a side arm. It is fitted with a stirrer.

(b) An outer tube in which freezing tube is adjusted. This tube serves as air jacket and helps to achieve a slower and more uniform rate of cooling.

(c) A large vessel containing a freezing mixture.

(d) A Beckmann thermometer which can read up to 0.01K.

Q28. Write daily life applications of depression of freezing point and elevation of boiling point.

Ans: Daily life applications of depression of freezing point and elevation of boiling point:

i. Ethylene glycol is used as an **antifreeze.**

ii. Ethylene glycol is nonvolatile in character and completely miscible with water.

iii. When mixed with water it lowers the freezing point but also raises the boiling point. In **winter** it protects a car by preventing the liquid from freezing whereas in **hot summer** it protects the radiator from over heating.

iv. The principle of freezing point depression is also used to prepare a freezing mixture for use in an ice cream machine. For this purpose **NaCl or $NaNO_3$** is used to **lower melting point of ice.**

Activity

In large quantities, the nicotine is a deadly poison. Take 13.4 g of pure water (solvent) in the freezing tube. Now fix thermometer in it, such that the bulb of the thermometer immerse in the solvent. Place freezing tube in the air jacket and air jacket in the large vessel containing a freezing mixture. Record accurate freezing point of solvent. After this remove freezing tube from the air jacket and re-melt the solvent. Add 3.62g of nicotine in the solvent through the side arm. Place freezing tube again in the air jacket and record freezing point of solution while stirring the solution constantly. Thermometer will show –0.563°C. Find the difference of the two freezing points. This will give the value of ΔT_f. From this data determine molecular mass of nicotine.

Solution:

Mass of Nicotine = W_2 = 3.62 g

Mass of water = W_1 = 73.4g

K_f for water = 1.86

Freezing point of water = T_1 = 0°C

Freezing point of solution = T_2 = -0.563°C

Molecular mass of nicotine = M_2 = ?

$\Delta T_f = T_1 - T_2$

$= 0 - (-0.563)$ $= 0.563°C$

Now

$$M_2 = \frac{K_1 \times W_2 \times 1000}{\Delta T_f \times W_1}$$

$$M_2 = \frac{1.86 \times 3.62 \times 1000}{73.4 \times 0.563}$$

$M_2 = 162.93 g\ mole^{-1}$

Thus molecular mass of nicotine is $162.93 g\ mole^{-1}$

SELF-CHECK EXERCISE 10.8

1. **Vitamin K_1 is a substance found in green leafy vegetables. It is needed by the body to produce a blood-clotting factor. A solution of 55.8 mg of vitamin K_1 in 1.048g of Benzene (freezing point of benzene is 5.455°C) has a freezing point of 4.850 °C. K_f for benzene is 5.12. What is the molecular weight of Vitamin K_1? (Ans: 450.596 g.mole^{-1})**

2. **Ethylene glycol $C_2H_6O_2$ is added to automobile radiator to prevent cooling water from freezing. Calculate the freezing point of ethylene glycol solution that contains 2kg ethylene glycol dissolved in 5 kg water. (K_f = 1.86). Also calculate the boiling point of this solution (K_b = 0.52). (Ans: F.P = -5412°C, B.P = 103.35°C)**

Solution:

1. $Mass\ of\ Benzene = 1.048\ g$

$K_f\ for\ benzene = 5.12$

$Mass\ of\ Vitamin\ K_1 = 58\ mg = \frac{55.8}{1000} = 0.0558\ g$

$Freezing\ Point\ of\ water = 5.455\ ^\circ C$

$Freezing\ Point\ of\ Solution = 4.850\ ^\circ C$

$Molecular\ mass\ of\ Vitamin\ K_1 = M_2 = ?$

As we know that

$\Delta T_f = T_1 - T_2$

$\Delta T_f = 5.455 - 4.850 = 0.605\ ^\circ C$

$$M_2 = \frac{K_f \times W_2 \times 1000}{\Delta T_f \times W_1}$$

$$M_2 = \frac{5.12 \times 0.0558 \times 1000}{0.605 \times 1.048} = 450.6\ g\ mole^{-1}$$

2.

$Mass\ of\ water = 5\ kg = 5000\ g$

$Mass\ of\ ethylene\ glycol = 2\ kg = 2000\ g$

$Boiling\ Point\ of\ water = 100\ ^\circ C$

$Boiling\ Point\ of\ Solution = T_2 = ?$

$K_b\ for\ water = 0.52$

Molecular mass of ethylene glycol $(C_2H_6O_2) = 12 \times 2 + 1 \times 6 + 16 \times 2$

$= 12 \times 2 + 1 \times 6 + 16 \times 2$

$= 62\ g\ mole^{-1}$

As we know that,

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

$$\Delta T_b = \frac{2.52 \times 2000 \times 1000}{62 \times 5000} = 3.355\ ^\circ C$$

Thus,

$\Delta T_b = T_2 - T_1$

$T_2 = T_1 + \Delta T_b$

$T_2 = 100 + 3.355 = 103.355\ ^\circ C$

For Freezing point of solution:

Mass of ethylene glycol $= 2\ kg = 2000\ g$

Molecular mass of ethylene glycol $(C_2H_6O_2) = 12 \times 2 + 1 \times 6 + 16 \times 2$

$= 12 \times 2 + 1 \times 6 + 16 \times 2$

$= 62\ g\ mole^{-1}$

Mass of water $= 5\ kg = 5000\ g$

Boiling Point of water $= T_1 = 100\ ^\circ C$

Boiling Point of Solution $= T_2 = ?$

K_f *for water* $= 1.86$

As we know that,

$$M_2 = \frac{K_f \times W_2 \times 1000}{\Delta T_f \times W_1}$$

$$\Delta T_f = \frac{K_f \times W_2 \times 1000}{M_2 \times W_1}$$

$$\Delta T_f = \frac{1.86 \times 2000 \times 1000}{62 \times 5000} = 12\ ^\circ C$$

Thus,

$\Delta T_f = T_1 - T_2$

$T_2 = T_1 - \Delta T_f$

$T_2 = 0 - 12 = -12\ ^\circ C$

Do You Know?

De-icing of aeroplane is based on freezing point depression. Think how?

Ice on frozen roads and sidewalks melts when sprinkled with salts such as NaCl and $CaCl_2$. This is because it depresses the freezing point of the water.

Q29. Explain osmotic pressure and daily life application of osmosis.

Ans: Osmotic Pressure:

Osmotic pressure is the pressure which needs to be applied to a solution to prevent the inward flow of water across a semipermeable membrane.

It is also defined as the minimum pressure needed to nullify osmosis.

The osmotic pressure of an ideal solution with low concentration can be approximated using the Morse equation (named after Harmon Northrop Morse).

$$\pi = \mathbf{MRT}$$

Where,

M is the molarity

$R = 0.08205746$ L atm K^{-1} mol^{-1} is the gas constant

T is the thermodynamic (absolute) temperature

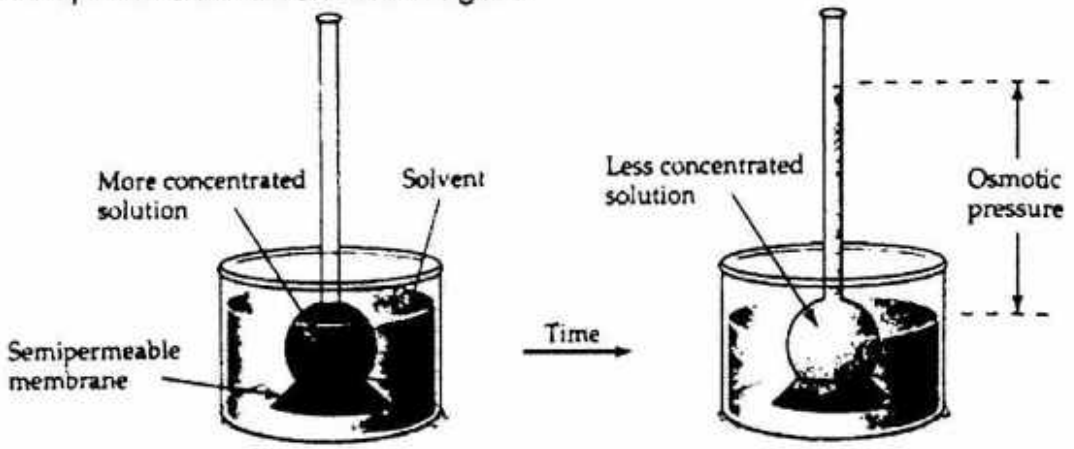
Examples:

i. Certain animal membranes such as that of bladder or outer covering of intestines are semipermeable. They allow the passage of water but not any solute dissolved in the water.

ii. A man made material such as cellophane can be used as semipermeable membrane.

Explanation:

A solution inside the bulb is separated from pure solvent in the beaker by a semipermeable as shown in figure.



The phenomenon of osmosis

As time passes the volume of solution increases and that of solvent decreases. The process continues until the hydrostatic pressure due to the extra height of the solution prevents further osmosis. Therefore the osmotic pressure of a solution is defined as the pressure, which must be applied above the solution to prevent passage of solvent through a semipermeable membrane into the solution.

Information:

i. Most important applications of osmosis are found in living organisms.

ii. If red blood cells are placed in pure water, the cells expand and ultimately rupture as a result of water entering the cells through **osmosis.**

iii. The osmotic pressure of the fluid inside the cell is equivalent to that of **0.95%** NaCl solution. Thus if cells are placed in **0.95% solution of NaCl**, there is no net flow of water through the cell walls and the cell remain stable this solution is said to be **isotonic.**

iv. Solutions having same osmotic pressures are **called isotonic solution.**

v. If concentration of NaCl solution is greater than 0.95%, water flows out of the cell walls and the cell shrinks. This NaCl **Solution is hypertonic.**

vi. If concentration of NaCl solution is less than 0.95%, water flows into the cell and the solution is called **hypotonic**.

Q30. What happened if red blood cells are placed in pure water?

Ans: If red blood cells are placed in pure water, the cells expand and ultimately rupture as a result of water entering the cells through **osmosis.**

SELF-CHECK EXERCISE 10.9

Rank the following aqueous solutions:

i) 0.02 m Urea ii) 0.02 m Glucose in order of decreasing

(a) boiling point (b) freezing point

(c) vapour pressure (d) osmotic pressure.

Solution:

Colligative properties depend upon the number of solute particles as the molality of urea and glucose is same i.e. 0.02 m thus they have same number of particles.
So for both solution boiling points, freezing points and osmotic pressures will remain same.

Q31. Define colloids with the help of examples.

Ans: Colloids:

A heterogeneous mixture of tiny particles of a substance dispersed through a medium is called colloidal dispersion or a colloid.

Dispersion medium or continuous phase:

The particles are called colloidal particles or the dispersed phase and the medium is called the dispersion medium or continuous phase (solvent).

Explanation:

i. In a colloid dispersed phase is like solute in a true solution and dispersion medium corresponds to the solvent in the true solution.

ii. The dispensed or suspended particles are single large molecules or aggregates of molecules or ions ranging in size from **1 to 10^3 nm.**

iii. This means these particles are bigger in size than those of true solutions.

iv. These particles are visible in electron microscope but invisible in ordinary microscope.

Example:

Milk, mayonnaise, smoke, rubber, foam, soap, fog, mist, gem stones like black diamond, opal etc are colloids.

Q32. Define lyophillic molecules.

Ans: Lyophillic molecules:

i. When a protein crystal in dropped into water.

ii. The polymer molecules dissolve to produce a colloidal dispersion.

iii. Such colloidal dispersions in which dispersed phase shows an affinity or attraction for the dispersion medium are called lyophillic **(means solvent loving).**

iv. The molecules of dispersed phase are called **lyophillic molecules.**

Examples:

Examples of lyophillic colloids are proteins, gum, gelatin, starch etc in water.

Q33. Define lyophobic.

Ans: Lyophobic:

i. When solid AgCl is brought in contact with water, it does not spontaneously disperse to form a colloid.

ii. Such a colloid, which cannot be made by spontaneous dispersion is called lyophobic (means solvent hating).

iii. The molecules of dispersed phase of such colloids have very little or no attraction for the dispersion medium, are called **lyophobic molecules or particles**.

Examples:

Examples of lyophobic colloids are sulphur, gold, Iron (III) hydroxide in water.

Q34. Define suspensions and also give its examples.

Ans: Suspensions:

i. A dispersions in which particles of the dispersed substance are visible to the naked eye are called **suspensions.**

ii. Sizes of these particles are greater than 10^3 nm.

iii. These particles are bigger than colloids.

iv. These particles cannot pass through ordinary filter paper.

v. They settle down under the influence of gravity.

vi. Suspensions scatter and reflect light.

Examples:

Mud or sand in water are suspensions.

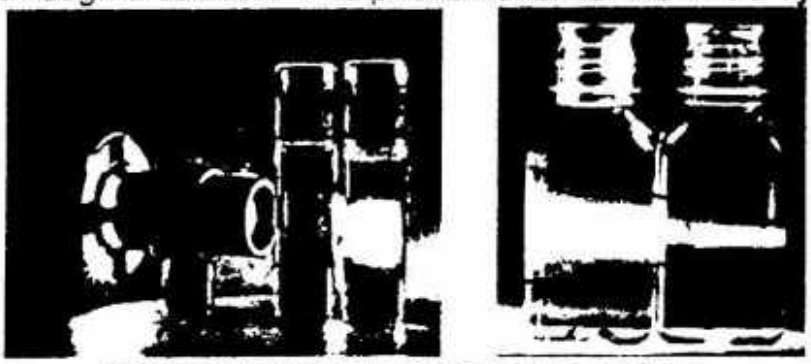
Q35. Write down the properties of colloids.

Ans: Properties of Colloids:

General properties of colloids are as follows:

1. Most colloids are **cloudy or opaque**, but some are transparent to the naked eye.

2. When light passes through a colloid, it is scattered by the dispersed particles because their sizes are similar to the **wavelength of visible light.** When viewed from the side, the scattered beam is visible and broader that one passing through a solution. This phenomenon is known as **Tyndall effect**.



Light scattering by colloidals suspensions

3. Colloidal particles exhibit **Brownian motion**. This can be seen under low magnification. This is because colloidal particles are pushed this way and that way by molecules of dispersing medium.

4. Colloidal particles **do not coagulate into larger particles** and settle out. This is due to the fact that particles have charged surfaces. These surfaces interact with the molecules of dispersing medium. For example, aqueous

protein has charge amino acid groups facing the water and uncharged groups buried within the molecule.

5. **Temperature changes affect colloids**. Increase in temperature makes the colloidal particles move faster and colloid more often and convert little particles into a lump. This means increase in temperature causes coagulation. For example, heating milk causes coagulation of casein in milk.
6. Colloidal particles have **little power of diffusion**. This is because the colloidal particles have very large size as compare to ordinary solute particles.
7. Colloidal particles **can pass through ordinary filter paper** but **cannot pass through ultra filter papers.**
8. Colloidal particles have **high ratio of surface area to volume** as compared to the particles of true solutions.
9. Colloidal particles **do not settle under the influence of gravity.**

Q36. How can you explain common types of colloids on the basis of their nature and composition?

Ans: Types of Colloids:

Colloids are classified according to whether the dispersed and dispersing substances are gases, liquids or solids.

Types of colloids

Colloid Type (Common Name)	Nature and Composition	Example
Sol	Solid particles dispersed through a liquid	Milk of magnesia, starch dispersed in water, paints, coloured glassed, gems
Gel	Continuous network of solid throughout the liquid medium	Halwa, jellies, gelatin
Aerosol	Either a solid or a liquid dispersed in a gas	Smoke (solid expressed in air), fog (liquid dispersed in air)
Emulsion	Liquid dispersed in an another liquid	Milk, mayonnaise
Foam	Bubbles of a gas suspended in a liquid or a solid	Canned shaving cream, soap leather, whipped cream

Q37. How can you compare colloids, suspensions and true solutions.

Ans: Comparison of Colloids, Suspensions and True Solutions:

The properties of a colloid depend mainly on the size, shape and charge of dispersed particles.

Table: Some properties of colloids, suspensions and true solutions:

Sr. No.	Properties	Colloids	Suspensions	True solutions
1.	Size of particles	1-10^3 nm	> 10^3 nm	0.1 – 1nm
2.	Phase	Heterogeneous	Heterogeneous	Homogenous

3.	Aggregates	Particles are composed of 10^3 to 10^9 atoms	Particles are composed of more than 10^9 atoms	Particles are composed of 1 to 10^3 atoms
4.	Charge on the particles	Positive or negative	Positive or negative or may be neutral	Both positive and negative or may be neutral
5.	Visibility of particles	Invisible by naked eye and ordinary micro-scope but visible in electron microscope	Visible by the naked eye and in ordinary micro-scope	Indivisible by the naked eye, ordinary micro-scope as well as electron micro-scope
6.	Filterability	Particles can pass through ordinary filter paper, but cannot pass through ultra filter paper	Particles cannot pass through ordinary as well as ultra filter paper.	Particles can pass through ordinary as well as ultra filter paper.
7.	Dispersion of light	Scatter light	Scatter light	Cannot scatter light
8.	Effect of gravity	Particles do not settle under the influence of gravity	Particles settle under the influence of gravity	Particles do not settle under the influence of gravity

SUMMARY OF KEY TERMS

1. A solution is a homogenous mixture of two or more substances, which may be solids, liquids or gases. The solubility of a solute in a given solvent depends on the interactions between the solute and the solvent particles. Substances that have similarly intermolecular forces tend to dissolve in each other, leading to the generalization "like dissolves like".
2. A rise in temperature usually increases the solubility of solids and liquids but decreases the solubility of gases.
3. Raoult's Law states that when a non-volatile solute is dissolved in a solvent, the vapour pressure of the solution (P) is directly promotional to the mole fraction of the solvent. (X_1) $P = P^0 X_1$
4. The concentration of a solution can be expressed as percent by mass, percent by volume, molarity, molality, mole fraction, ppm, ppb and ppt. The unit of concentration most commonly used by the chemist is molarity,

which is defined as the number of moles of solute dissolved per dm^3 of solution.

5. Sometimes it is useful to know the ratio of the amount of solute to the amount of solvent. For this purpose units of molality and mole fraction are used. These units are particularly important in the study of colligative properties, that depend only on the number of solute particles that are present and not on their nature.
6. Vapour pressure lowering, boiling point elevation, freezing point depression and osmotic pressure are colligative properties of solution. Adding a solute to a solvent lowers the vapour pressure of the solvent. This causes an elevation in boiling point and depression in freezing point of solution.
7. Osmotic pressure causes the solvent to flow across a semi-permeable membrane when the concentrations of solute on either side of the membrane are unequal.
8. Colligative properties can be used to determine the molecular weight of the solute. This is because these properties depend only on the number of solute particles in a solution and not on their identity.
9. Sea water can be made drinkable by the process of reverse osmosis.
10. A colloid is a heterogeneous mixture of tiny particles of a substance dispersed through a medium. These particles are only visible in electron microscope. When light passes through a colloid, it is scattered by the dispersed particles. This phenomenon is known as Tyndall effect.
11. A dispersion in which particles of the dispersed substance are visible to the naked eye is called suspensions.

EXERCISE

MULTIPLE CHOICE QUESTIONS

1: Choose the correct answer

(i) 18g glucose is dissolved in 180g of water. The relative lowering of vapour pressure is

(a) 1 (b) 1.8 (c) 0.01 (d) 0.001

(ii) 100g of a 10% (W/W) NaOH contains 10g of NaOH in

(a) 100g of H_2O (b) 110g of H_2O
(c) 10g of H_2O (d) 100g solution.

(iii) Which of the following W/W solutions has the lowest freezing point?

(a) 18% glucose (b) 6% urea
(c) 34.2% sucorose (d) 9% glucose

(iv) All solutions containing 1g of non-volatile solutes will have

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(c) 34.2% sucorose (d) 9% glucose

(iv) All solutions containing 1g of non-volatile solutes will have

(a) Same vapour pressure (b) Same boiling point
(c) Same freezing point (d) all of a, b, c

(v) Molarity of pure water is

(a) 1 (b) 1.8 (c) 55.5 (d) 18

(vi) A solution of urea (Mol.Wt=60) is 10% (W/V). the volume in which 1 mole of it is dissolved will be

(a) $6dm^3$ (b) $0.6dm^3$ (c) $60\ cm^3$ (d) $0.54\ dm^3$

(vii) The solubility of a substance decreases with increase in temperature if heat of solution is

(a) positive (b) negative
(c) zero (d) cannot be predicted

(viii) Molarity of glucose solution when 9g of it is dissolved in 250 cm^3 of solution is

(a) 0.25M (b) 0.2m (c) 0.5M (d) None of these

(ix) Sea water has about 6 ppm dissolved oxygen. What mass of dissolved oxygen is present in one kg of sea water

(a) 6g (b) 6×10^{-3}kg (c) 6×10^{-3}g (d) 6×10^{-3}mg

(x) Which of the following is volatile

(a) glucose (b) urea (c) sucrose (d) water

(xi) Which of the following is not true of a colloid.

(a) it is heterogeneous (b) scatter light
(c) can pass through ultra filter paper
(d) Its particles cannot be seen under ordinary microscope.

(xii) Aerosol are colloids which contain

(a) A solid dispersed in a liquid
(b) A solid dispersed in a gas
(c) A liquid dispersed in a gas
(d) A liquid in another liquid

1. a
2. a, b
3. b, c
4. c, d

Answers

i. c	ii. d	iii. d	iv. d	v. c	vi. b	vii. b
viii. d	ix. c	x. d	xi. c	xii. 3		

2. Determine the molality of the following solutions

i) 3.5% (W/W) glucose (aq) (Ans: 0.2 g dm^{-3})

ii) 9.9g $NaNO_3$ dissolved in 940g of water (Ans: 0.124 g dm^{-3})

Solution: I) 3.5% (W/W) glucose (aq)

Mass of Glucose = 3.5g

Molar mass of Glucose ($C_6H_{12}O_6$) = 180

Mass of Water = 100 − 3.5 = 96.5g

Molality is given by

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

$$\text{Molality} = \frac{\text{Mass of Solute}}{\text{Molar mass of solute}} \times \frac{1000}{\text{mass of solvant in gram}}$$

$$\text{Molality} = \frac{3.5}{180} \times \frac{1000}{96.5} = 0.20\ \text{m}$$

ii) 9.9g $NaNO_3$ dissolved in 940g of water

Mass of water = 940 g

Mass of $NaNO_3$ = 9.9 g

Formula Mass of $NaNO_3$ = 23 + 14 + 48 = 85 g mole^{-1}

Using Molarity relation

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

$$\text{Molarity} = \frac{\text{Mass of Solute}}{\text{Molar mass of solute}} \times \frac{1000}{\text{mass of solvant in gram}}$$

$$\text{Molarity} = \frac{9.9}{85} \times \frac{1000}{940} = 0.124\ \text{m}$$

3. What are the freezing and boiling points of a solution prepared by dissolving 13.3g of ethylene glycol in 100g of water. (Ans: F.P =-3.99°C, B.P = 101.115°C)

Solution:

Mass of Solute (ethylene glycol) = W_2 = 13.3g

Molar mass of Solute = 62 g mole^{-1}

Mass of solvent (water) = W_1 = 100g

Molar mass of solvent (water H_2O) = 18g mole^{-1}

Molal F. P constant, K_f for water = 1.86 °C

Mol. Elevation B. P constant, K_b for water = 0.52° C

Freezing point of Solution = ?

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

$$= \frac{1.86 \times 13.3 \times 1000}{62 \times 100} = 3.99°C$$

Let Freezing point of pure water = T_1 = 0°C

Freezing point of solution = T_2 = ?

$$\Delta T_f = T_1 - T_2$$

or $$T_2 = T_1 - \Delta T_f = 0 - 3.99$$

$$= -3.99°C$$

Boiling point of Solution = ?

As

$$\Delta T = \frac{K_b \times K_b \times 1000}{M_2 \times W_1} = \frac{0.52 \times 13.3 \times 1000}{62 \times 100} = 1.115°C$$

Let Boiling point of pure water = T_1 = 100°C

Boiling point of solution = T_2 = ?

$\Delta T_b = T_2 - T_1$

$T_2 = T_1 + \Delta T_b$ $= 100 + 1.116° = 101.116°C$

4. When cooking, what effect does adding salt to water have on the time required to boil food?

Ans: When the salt is added, a phenomenon known as "boiling point elevation" take place. Boiling point elevation happens when a non-volatile solute (or a dissolvable substance—in this case, the salt) is added to a pure solvent (or a substance that dissolves a solute—in this case, the water itself) to create a solution (the salt water).

By adding salt, the vapour pressure of water decreases due to which its boiling point increases. Hence salt water requires more exposure to the heat in order to boil than water alone, so the boiling point is elevated and more time is required to boil food.

5. Concentrated sulphuric acid is 98% (W/W) H_2SO_4 (aq). Its density is 1.84g.cm^{-3}. Calculate

i) The molarity of the solution (Ans:18.4 M)

ii) The molality of the solution (Ans:10molal)

iii) What quantity of conc.H_2SO_4 is required to prepare 500cm^3 of 0.1M H_2SO_4.

Solution: (i) Molarity of Solution:

98% (W/W) solution means 98 g H_2SO_4 is dissolved in 100 g of solution

$Mass\ of\ H_2SO_4 = 98g$

$Density\ of\ H_2SO_4 = 1.84\ gcm^{-3}$

$Molar\ mass\ of\ H_2SO_4 = 2 + 32 + 64 = 98gmole^{-1}$

$$Number\ of\ moles = \frac{Given\ mass}{Molar\ mass}$$

$$Number\ of\ moles = \frac{98}{98} = 1\ mole$$

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

$$Volume = \frac{Mass}{density} = \frac{100}{1.84} = 54.3478\ cm^3$$

$54.3478\ cm^3\ contains\ H_2SO_4 = 1\ mole$

$$1000\ cm^3 contains\ H_2SO_4 = \frac{1}{54.3478} \times 1000 = 18.4\ moles$$

$So\ molarity\ of\ H_2SO_4 = 18.4\ moles$

(ii) The molality of the solution:

98% (W/W) solution means 98 g H_2SO_4 is dissolved in 100 g of solution

Mass of H_2SO_4 = 98 g

Mass of Solution = 100 g

Mass of Solvent = 100 g - 98 g

= 2 g = 0.002 kg

Molar mass of H_2SO_4 = 98g mole^{-1}

As, $$Molality = \frac{m_{solute}}{M_{solute}} \times \frac{1}{W_{solvent(kg)}}$$

$$Molality = \frac{98}{98} \times \frac{1}{0.002} = 500m$$

(iii) What quantity of Concentrated H_2SO_4 is required to prepare 500cm^3 of 0.1M H_2SO_4.

Molarity of H_2SO_4 present = M_1 = 18.4 M
Volume of H_2SO_4 required = V_1 = ?
Molarity of H_2SO_4 prepared M_2 = 0.1 M
Volume of H_2SO_4 prepared V_2 = 500 cm^3
As we know that

$M_1 V_1 = M_2 V_2$

$$V_1 = \frac{M_2 \times V_2}{M_1} = \frac{0.1 \times 500}{18.4} = 2.717cm^3$$

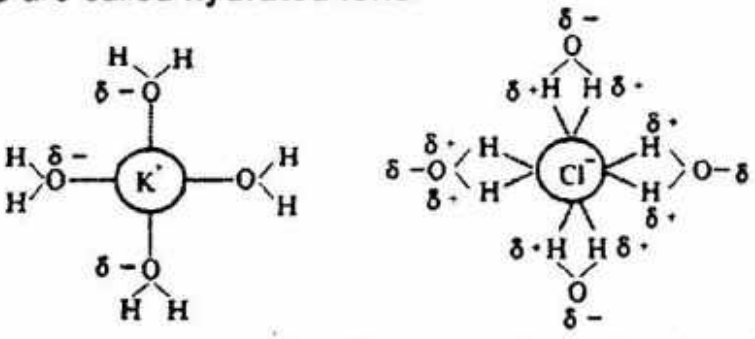
Thus $2.717cm^3$ of 18.389M H_2SO_4 will be taken in a 500 cm^3 flask and the volume is made up to the mark. This is required 0.1M H_2SO_4 solution.

6. What would you expect potassium chloride, KCl, an ionic solid to be more soluble in H_2O or CCl_4? Explain your prediction.

Ans: The basic principle of solubility is **"Like dissolves like"**. KCl is an ionic solid. Ionic solids have a crystal lattice structure composed of oppositely charged ions.

Examples:

i. When an ionic solid e.g. KCl is placed in water which is a polar solvent. These ions are attracted by polar molecules. Water molecules break the crystal lattice of k^+Cl^- and then surrounds the resulting k^+ and Cl^- ions. These ions are called **hydrated ions**.



ii. When kCl crystals are placed in CCl_4. Non-polar molecules of these liquids are unable to attract ions in Na^+Cl^- and cannot break apart the crystal lattice. Thus kCl is **insoluble** in these solvents

7. List three factors that accelerate the dissolution process.

Ans: Following are the three factors that accelerate the dissolution process:

(1) Particle size:

When a sugar cube in placed into water, it dissolves slowly than does an equal amount of finely granulated sugar. A sugar cube exposes less surface area to the water molecules than do the tiny particles of granulated sugar. Inner regions of cube dissolves only after the outer layers are in solution.

There are fewer unexposed particles when sugar is granulated, so it dissolves faster. **Thus, to increase the rate of dissolving, large piece of solid must be grind into small particles.**

(2) Temperature:

Temperature is another factor which changes the rate at which solutes dissolve in a solvent. At higher temperature, increased molecular motion increases the interaction of solute and solvent particles which increases the dissolving rate.

(4) Stirring:

Stirring a solution increases the rate at which a solid dissolves decreasing the concentration of solute in the immediate region surrounding the solid solute. Stirring also increases the amount of exposed solute surface to the solvent.

8. Explain the nature of solutions in liquid phase giving examples of completely miscible, partially miscible and immiscible liquid-liquid solutions.

Ans: Nature of Solution in Liquid phase:

Solutions of liquids in liquids may be divided into three classes.

(a) Completely miscible **(b)** Partially miscible **(c)** Immiscible.

(a) Completely Miscible Liquids:

Liquids which are miscible in all proportions are called completely miscible.

Examples:

Water and methanol, Acetone and water, Benzene and cyclohexane are completely miscible liquid pairs. Substances that dissolve in each other possess similar types of inter-molecular interactions.

Water and methanol:

Water and methanol are miscible because molecules in the pure liquids and their mixtures form hydrogen bonds. The degree of hydrogen bonding in the solution is same as that in the pure liquids.

Benzene and cyclohexane:

Benzene and cyclohexane are completely miscible because the molecules in the pure liquids and their mixtures interact through London dispersion forces. These dispersion forces between benzene and cyclohexane are about the same as those in pure liquids.

(b) Partially Miscible Liquids:

Liquids that dissolve in each other to a very small extent are called partially miscible liquids.

On mixing such liquids two layers are formed. Each layer is a saturated solution of the other liquid.

For example ether and water are partially miscible liquids. The mutual solubilities of these solutions change by temperature changes.

Examples:

Typical examples of such systems are:-

(i) Phenol-water
(ii) Aniline-water
(iii) Aniline-n-hexane
(iv) Nicotine-water

(c) Completely Immiscible Liquids:

Liquids which do not dissolve into each other in any proportion at any temperature are called completely immiscible.

Examples:

For instance water and benzene, carbon disulphide and water, cyclohexane and water are completely immiscible in one another.

Benzene, carbon disulphide or cyclohexane:

Benzene, carbon disulphide or cyclohexane are non-polar in nature. The only forces of attraction between their molecules are dispersion forces. In water hydrogen bonds are much stronger than dispersion forces. Since hydrogen bonds cannot be disrupted by the molecules of benzene, carbon disulphide or cyclohexane, these liquids are immiscible with water.

9. Express solution concentration in terms of mass percent, molality, molarity, parts per million, parts per billion and parts per trillion and mole fraction.

Ans: Mass Percent:

It is defined as the mass of solute present in 100g of solution. It is also referred as percent weight / weight.

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

15% NaCl solution means 15g of NaCl dissolved per 100 g of solution or 85g of water.

Molarity (M):

i. It is defined as the number of moles of solute dissolved per dm^3 of solution.

ii. Molarity is moles of solute per liter of solution.

iii. Molarity is expressed as the moles per **Liter/s of solution.**

iv. $M = \frac{moles\ of\ solute}{dm^3 of\ solution}$

$$M = \frac{grams\ of\ solute}{molar\ mass\ of\ solute \times dm^3\ of\ solution}$$

Example:

If 0.5 mole of NaOH (20g) is dissolved in enough water to make one dm^3 of solution, 0.5 molar or 0.5 M NaOH solution is obtained.

Molality (m):

i. It is defined as the number of moles of solute dissolved per kilogram of solvent.

ii. Molality is moles of solute per kilogram of solvent.

iii. Molality is moles per **Kilogram/s of solvent**.

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

Example:

When 58.5g NaCl (1 mole) is dissolved in one kilogram of water, the resulting, solution would be one molal or 1m NaCl solution.

Parts per million (ppm):

It is defined as the number of parts by weight (or volume) of a solute per million parts by weight (or volume) of the solution.

$$ppm = \frac{Mass\ or\ volume\ of\ solute}{Mass\ or\ volume\ of\ solution} \times 10^6$$

Parts per billion (ppb):

It is defined as the number of parts by weight (or volume) of a solute per billion parts by weight (or volume) of the solution.

$$ppb = \frac{Mass\ or\ volume\ of\ solute}{Mass\ or\ volume\ of\ solution} \times 10^9$$

Parts per trillion (ppt):

It is defined as the number of parts by weight (or volume) of a solute per trillion parts by weight (or volume) of the solution.

$$ppt = \frac{Mass\ or\ volume\ of\ solute}{Mass\ or\ volume\ of\ solution} \times 10^{12}$$

Mole Fraction (X):

It is defined as the ratio of the number of moles of a given component to the total number of moles of solution.

Examples:

Suppose a solution contains n_A and n_B moles of two components A and B. Mole fraction of each component is given by

Mole fraction of component A = $X_A = \frac{n_A}{n_A + n_B}$

Mole fraction of component B = $X_B = \frac{n_B}{n_A + n_B}$

Note:

Rare components of a solution are expressed by ppm, ppb and ppt. A pollutant concentration of 1 ppm in the atmosphere means that out of one million molecules of air there is one molecule of the pollutant, whereas 1 ppb concentration of it means that out of one billion molecules of air these is one molecule of that pollutant.

10. Distinguish between the solvation of ionic species and molecular substances.

Ans: Similarities:

The process of solvation of ionic and molecular substances is similar in the sense in that both the cases the process of solvation involves the disruption of solute particles from the crystal lattice. Then these particles are surrounded by solvent molecules.

Differences:

The heat of solvation in both the cases is much different since ion-dipole interactions are much stronger than dipole-dipole or dispersion forces, heat of solvation for ionic species is generally larger than those for molecular solvents.

11. Describe the role of solvation in the dissolving process.

Ans: Role of Solvation in the Dissolving Process:

i. Solution making process is accompanied by either the absorption or evolution of heat.

ii. Energy change in the formation of a solution depends upon three types of interactions between solute-solute, solvent-solvent and solute-solvent particles.

iii. In the formation of solution, interactions between solute particles are broken. At the same time solvent molecules also move apart to accommodate the solute particles.

iv. Since energy is needed to break interactions. Therefore, both these processes are endothermic.

v. Simultaneously interactions between the particles of solute and solvent are established i.e., solvation occurs.

vi. Solvent molecules surround solute particles from all sides.

vii. In these interactions energy is released, so it is exothermic.

Result:

Thus the strength of these two types of interactions determines whether the process of dissolution is endothermic or exothermic.

12. Define the term water of hydration.

Ans: Hydrate:

A crystalline substance which has associated with each formula unit a definite number of water molecules is called a hydrate.

Such water molecules are called **water of crystallization** or water of hydration

Water of Hydration:

The water molecules that combine with compounds as they are crystallized from aqueous solution are called **water of crystallization or water of hydration**.

Example:

(i) $CuSO_4.\ 5H_2O$ **(ii)** $BaCl_2.\ 2H_2O$ **(iii)** $MgSO_4.\ 7H_2O$

(iv) $(COOH)_2.\ 2H_2O$ **(v)** $Na_2CO_3.\ 10H_2O$

13. Define the term colligative.

Ans: The properties of solutions which depends only on number of solute particles and not on their nature. Those properties are called as colligative properties.

Examples:

(i) Lowering of Vapour pressure

(ii) Elevation of Boiling point

(iii) Depression of Freezing point

(iv) Osmotic pressure

14. List the characteristic of colloids and suspensions.

Ans: Characteristic of colloids and suspensions:

Sr. No.	Properties	Colloids	Suspensions
1.	Size of particles	$1\text{-}10^3$ nm	$> 10^3$ nm
2.	Phase	Heterogeneous	Heterogeneous
3.	Aggregates	Particles are composed of 10^3 to 10^9 atoms	Particles are composed of more than 10^9 atoms
4.	Charge on the particles	Positive or negative	Positive or negative or may be neutral
5.	Visibility of particles	Invisible by naked eye and ordinary micro-scope but visible in electron microscope	Visible by the naked eye and in ordinary micro-scope
6.	Filterability	Particles can pass through ordinary filter	Particles cannot pass through ordinary as

		paper, but cannot pass through ultra filter paper	well as ultra filter paper.
7.	Dispersion of light	Scatter light	Scatter light
8.	Effect of gravity	Particles do not settle under the influence of gravity	Particles settle under the influence of gravity

15. Define hydrophilic and hydrophobic molecules.

Ans: Hydrophilic molecules:

Hydro mean water, philic means loving. So hydrophilic literally means water loving.

Molecules that are miscible with water are known as hydrophilic molecules. Such molecules can form hydrogen bond with water molecules.

Example:

For example, molecules of methanol, acetone, acetic acid etc. are called hydrophilic molecules.

Hydrophobic molecules.

Phobic means disliking, therefore hydrophobic literally mean water disliking.

Molecules that do not dissolve in water are known as hydrophobic molecules.

Example:

For example, molecules of organic fats and oils are called hydrophobic molecules.

16. Describe on a particle basis why a solution has a lower vapour pressure than the pure solvent.

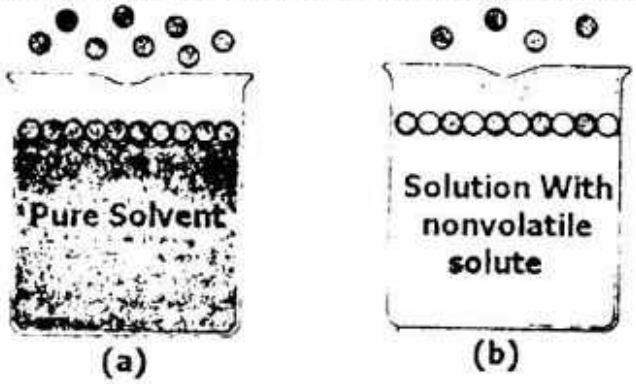
Ans: Causes of Lowering in Vapour Pressure:

i. Every liquid has a definite vapour pressure at a particular temperature.

ii. In a pure liquid, all the surface particles are that of the liquid.

iii. But in a solution containing a nonvolatile solute, both solute and solvent particles occur on the surface.

iv. The solute particles decrease the number of solvent surface particles.



Lowering of vapour pressure

v. Decreasing the number of surface solvent particles decreases the rate of evaporation of the solvent, which decreases the vapour pressure.

Conclusion:

Thus the vapour pressure of a solution containing a nonvolatile solute is always less than the vapour pressure of the pure solvent.

17. Explain osmotic pressure, reverse osmosis and give their daily life applications.

Ans: Osmotic Pressure:

Osmotic pressure is the pressure which needs to be applied to a solution to prevent the inward flow of water across a semipermeable membrane.

It is also defined as the minimum pressure needed to nullify osmosis.

The osmotic pressure of an ideal solution with low concentration can be approximated using the Morse equation (named after Harmon Northrop Morse).

$$\pi = MRT$$

Where,

M is the molarity

$R = 0.08205746$ L atm K^{-1} mol^{-1} is the gas constant

T is the thermodynamic (absolute) temperature

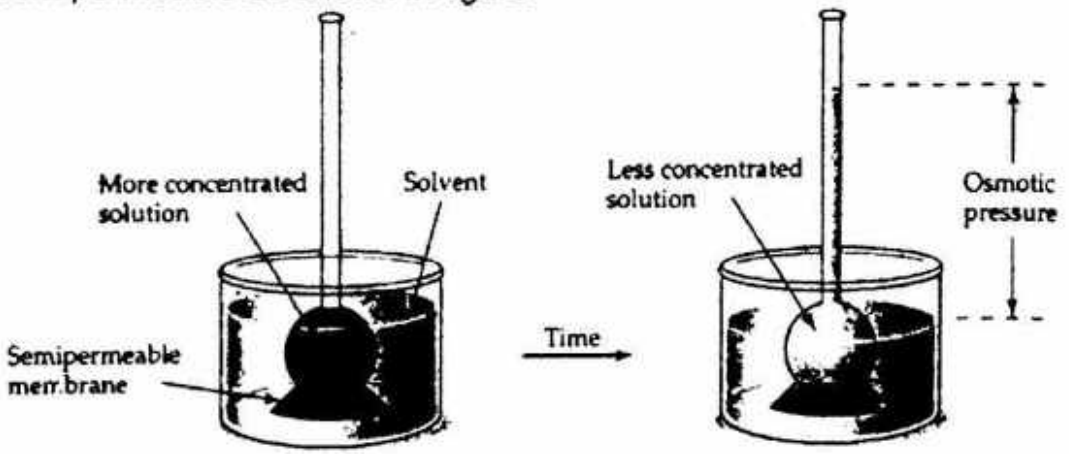
Examples:

i. Certain animal membranes such as that of bladder or outer covering of intestines are semipermeable. They allow the passage of water but not any solute dissolved in the water.

ii. A man made material such as cellophane can be used as semipermeable membrane.

Explanation:

A solution inside the bulb is separated from pure solvent in the beaker by a semipermeable as shown in figure.



The phenomenon of osmosis

As time passes the volume of solution increases and that of solvent decreases. The process continues until the hydrostatic pressure due to the extra height of the solution prevents further osmosis. Therefore the osmotic pressure of a solution is defined as the pressure, which must be applied above the solution to prevent passage of solvent through a semipermeable membrane into the solution.

Daily Life Applications of Osmosis:

The osmosis phenomenon manifests itself in many interesting daily life applications.

1. Biochemists use a technique called hemolysis to study the contents of red blood cells. These cells are protected from the external environment by a semi permeable membrane. Red blood cells are placed in a hypotonic solution. Since hypotonic solution is less concentrated than the interior of the

cell, water moves into the cells. The cells swell and eventually bursts releasing hemoglobin and other molecules.

2. Osmosis is the major mechanism for transporting water upward in plants. Because leaves constantly loose water to the air by a process called transpiration. As a result, leaf fluids become more concentrated than the ground water. Thus water enters membranes in the roots and rises into the tree. It can create an osmotic pressure that can exceed 20 atm in the tallest trees.
3. A large quantity of sugar is essential to preserve jam and jelly. This is because the sugar helps to kill bacteria that may cause botulism. When a bacterial cell is in a hypertonic solution. The intracellular water tends to move out of the bacterial cell to the more concentrated sugar solution by osmosis. This causes the bacterial cell to shrink and eventually to cease functioning
4. Food is preserved by coating with salt, which produces hypertonic solution. Thus food coated with salt causes microbes on the surface to shrivel and die from loss of cell water.

Reverse Osmosis:

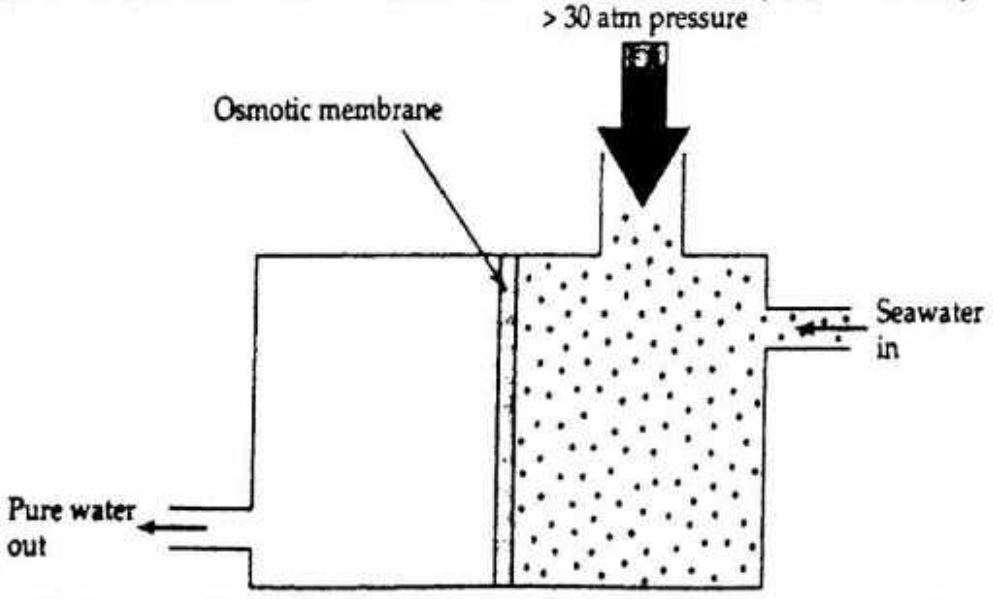
If a solution in contact with pure solvent across a semipermeable membrane is subjected to an external pressure equal to osmotic pressure, it stops osmosis. If external pressure is greater than solution's osmotic pressure, it will force solvent to flow from solution to the solvent. This process is called reverse osmosis.

Daily Life Application of Reverse Osmois:

i. Sea water is highly hypertonic to body fluids and thus is not drinkable.
ii. By reverse osmosis it is subjected to desalination (removal of dissolved salts) to make it drinkable.
iii. Desalination plant (as shown in figure), remove large amounts of dissolve salts from sea water.

Working:

The sea water is pumped under high pressure (20 atm) through the semipermiable membrane, which allow water molecules to pass and stop ions.



A schematic for the desalination of seawater by reverse osmosis

18. Define heat of solution and apply this concept to the hydration of ammonium nitrate crystals.

Ans: The enthalpy of solution, enthalpy of dissolution or heat of solution is the enthalpy change associated with the dissolution of a substance in a solvent at constant pressure resulting in infinite dilution.

Unit:

The enthalpy of solution is most often expressed in kJ/mol at constant temperature.

Explanation:

i. When a solute dissolves in water, attractive forces among these solute particles must be overcome to break solute particles from the solid lattice.

ii. A strong force must bind these separated solute particles to water molecules in the solution.

iii. Therefore, heat of solution measures the **net energy flow** that occurs as a substance dissolves.

iv. The energy needed to break solute particles is equal to lattice energy of solute ($\Delta H_{lattice}$). On the other hand the energy released when solute particles bind to the water molecules in the solution is called heat of hydration (ΔH_{hyd}).
Heat of hydration increase with increasing charge on the ions and decreases with increasing size of ions.

v. For other solvents (except water) this energy is known as **heat of solvation**.

vi. Heat of solution includes both the energy needed to move solvents molecules move apart and energy release when they surround solute particles. Thus, heat of solution $\Delta H_{solutoin}$ is the difference in $\Delta H_{lattice}$ and ΔH_{hyd}.

$$\Delta H_{solutoin} = \Delta H_{lattice} - \Delta H_{hyd}$$

Because of the different attractive forces in both solute and solution heat, heats of solution vary from significantly exothermic to significantly endothermic.

vii. If the crystal binds solute particles more tightly then the solution, solute must absorb energy as it dissolves and we have endothermic solution process.

viii. When this $\Delta H_{solution}$ is very large, the solute is unlikely to be soluble. But if this $\Delta H_{solution}$ moderately positive, then solute dissolves.

Examples:

Hydration of ammonium nitrate:

When ammonium nitrate dissolves in water 25.7 KJmole^{-1} energy is absorbed and the flask becomes cold. This is because crystal binds NH_4^+ and NO_3^- ions more tightly than the solution.

$NH_4NO_3(s) + xH_2O(l) \longrightarrow NH_4^+(aq) + NO_3^-(aq)$ $\Delta H_{solution} = 25.7$ KJmole^{-1}

On the other hand, if solution binds solute particles more tightly then the crystal, energy is released as the solute dissolves and we have an exothermic solution process.

Hydration of NaOH:

For example, when NaOH dissolves in water 44.5 KJmole^{-1} energy is released and the flask becomes hot

$NaOH(s) + xH_2O(l) \longrightarrow Na^{+}(aq) + OH^{-}(aq)$ $\quad \Delta H_{soln} = 44.5$ KJmole^{-1}

Thus solution process can be endothermic or exothermic depending upon the difference in the lattice energy for solute and heat of solvation for solute particles.

Other examples are given below:

$CaCl_2(s) + xH_2O(l) \longrightarrow Ca^{+2}(aq) + 2Cl^{-}(aq)$ $\quad \Delta H_{soln} = -83.0$ KJmole^{-1}

$MgSO_4(s) + xH_2O(l) \longrightarrow Mg^{+2}(aq) + SO_4^{-2}(aq)$ $\quad \Delta H_{soln} = -91.2$ KJmole^{-1}

The values of standard enthalpies of solution of some ionic solids in water are given in table.

Enthalpies of Solutions of some ionic solids.

Substance	Enthalpy of Solution (KJ mole^{-1})
LiCl	-37.0
NaCl	+2.98
KCl	+17.2
KI	+20.3
NH_4NO_3	+25.7
$AlCl_3$	-321.0

19. Explain the phenomenon of freezing in a mixture of ice and salt.

Ans: Melting point is a property of solids. The melting point is simply the temperature at which the solid turns into a liquid. For example, the melting point of normal ice is 0°C at standard pressure.

Freezing point is a property of liquids. The freezing point is simply the temperature at which the liquid turns into a solid. Thus the freezing point of normal freshwater is 0°C.

The value of the melting and freezing points is the same (ie. both are 0°C for water).

When salt is added to ice it lowers the freezing point by decreasing the vapour pressure of the solution. Thus the solution will freeze at a lower temperature than that of the pure solvent. Therefore addition of a non-volatile solute causes a depression in freezing point of solution.

20. Explain how solute particles may alter the colligative properties.

Ans: Causes of Boiling Point elevation:

i. When a non-volatile and non-electrolyte solute is added to a solvent, its vapour pressure is decreased.

ii. This is because in solution both solute and solvent particles occur on the surface.

iii. Thus solute particles decrease the number of solvent surface particles.

iv. This decreases the rate of evaporation of solvent, which decreases the vapour pressure.

v. Therefore, a solution must be heated to a higher temperature than the boiling point of pure solvent to equalize its vapour pressure to the atmospheric pressure.

Conclusion:

Thus addition of solute to a pure solvent causes an elevation of the boiling point of solution.

Causes of Freezing Point Depression:

i. The decrease in vapour pressure of a pure solvent on the addition of a solute also affects the freezing point of the solution.

ii. The solution will freeze at a temperature at which vapour pressure of both solution and solid solvent are the same.

iii. This means solution will freeze at a lower temperature than that of the pure solvent.

Conclusion:

Thus addition of a non-volatile solute also causes a decrease or depression in freezing point of solution.

Magnitude of colligative:

Moreover, the magnitude of colligative properties increases with increase in number of particles i.e., when concentration of solute is increased.

21. Explain the effect of temperature on solubility.

Ans: Q. Describe the effect of temperature on solubility.

Ans: Effect of Temperature on Solubility:

Generally, an **increase in temperature increases solubility** of a solid in a liquid. At higher temperature greater masses of solutes dissolve in a fixed mass of water than at lower temperature.

Solubility curve:

A curve drawn between solubility and temperature is called **solubility curve**.

Exceptions:

i. The solubilites of solid solutes generally increase with the increase in temperature. However this is not always the case.

ii. The solubilties of some solids e.g. $Na_2 SO_4$ and $Ce_2 (SO_4)_3$ decrease with increasing temperature.

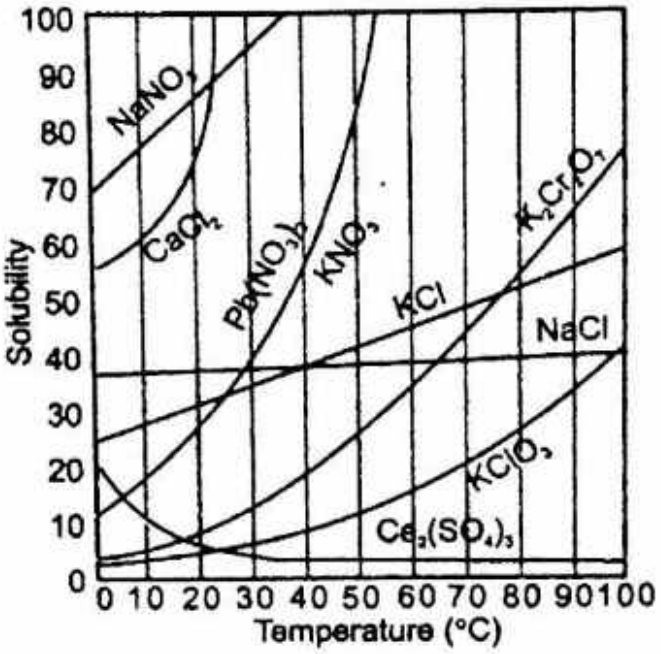
iii. Some substances like NaCl have relatively constant solubility with increasing temperature.

Prediction by enthalpy of solution:

The variation of solubility with temperature can be predicted from a consideration of the enthalpy of solution.

i. If ΔH°_{Soln} in negative, the solubility of solute decreases as the temperature increases.

ii. If ΔH°_{Soln} in positive, the solubility increases with temperature.



Solubility curves

Explanation of Solubility curves:

Solubility of KNO_3 rapidly increases with increase in temperature, this is because it has large positive value of ΔH°_{Soln} (+36 KJ/mole) The solubility of NaCl does not vary much with temperature because of relatively small magnitude of ΔH°_{Soln} (+3 KJ/mole).

Solubility of gases:

It is generally observed that the solubility of a gas **decreases with increasing temperature.**

The dissolving of gases in liquids can be understood in terms of two processes

(a) The condensation of the gas which is exothermic.

(b) The creation of holes in the liquid to accommodate the condensed gas molecules, it is endothermic.

Because of the open structure of water, little work is required to accommodate gas molecules, this means dissolving process is expthermic. **Thus solubilities of gases in water decrease with temperature.**

22. Compare molal and molar solutions.

Ans:

Molal Solutions	Molar Solutions
i. A molal solution is a solution that contains 1 molecular weight of solute in a kilogram of solvent.	**i.** Molar solution is an aqueous solution that contains 1 mole (gram-molecular weight) of solute in 1 liter of the solution.
ii. It is a concentration of a solution expressed in moles or molality (m).	**ii.** It is expressed in molarity (M).
iii. It is independent of temperature.	**iii.** Its value depends upon temperature.
iv. 1 molal solution is dilute than 1 molar solution.	**iv.** 1 molar solution is concentrated then 1 molal solution.
Example: When 58.5g NaCl (1 mole) is dissolved in one kilogram of water, the resulting, solution would be one molal or 1m NaCl solution.	**Example:** If 0.5 mole of NaOH (20g) is dissol.ved in enough water to make one dm^3 of solution, 0.5 molar or 0.5 M NaOH solution is obtained.

23. Explain the following

(a) Molality is independent of temperature but malarity depends on it.

(b) One molal solution of glucose in water is dilute as compared to one molar solution of glucose, but the number of particles of solute is the same.

(c) The total volume of solution by mixing $50cm^3$ of ethanol and $50cm^3$ of water may not be equal to $100cm^3$, why?

(d) NaCl and $NaNO_3$ are used to lower the melting point of ice.

(e) In summer the antifreeze solutions protect the radiator from boiling over.

(f) One molal and two molal solutions of urea boil at different temperatures.

(g) Relative lowering of vapour pressure is independent of the temperature.

(h) The sum of mole fractions of all the components is always equal to one.

Ans:

(a) Molality is independent of temperature but molarity depends on it.

Molality:

In case of molality the concentration of solvent is expressed in terms of mass. The mass of substance is not affected by the change in temperature. Hence molality is independant of temperature.

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

Molarity:

In case of molarity the concentration of solution is expressed in terms of volume which change with the change of temperature. Hence molarity is temperature dependant.

$$M = \frac{moles\ of\ solute}{dm^3 of\ solution}$$

$$M = \frac{grams\ of\ solute}{molar\ mass\ of\ solute \times dm^3\ of\ solution}$$

OR

Molality is independent of temperature but molarity depends on it is because volume of a solution changes with temperature and Molarity = no. of moles / volume of solution.

But Molality = Number of moles / mass of solvent and hence molality is not altered by temperature.

(b) One molal solution of glucose in water is dilute as compared to one molar solution of glucose, but the number of particles of solute is the same.

The molal aqueous solution of a solute say glucose is dilute in comparison to its molar solution.

Reason:

The reason is that in molal solution the quantity of the solvent is comparatively greater. While in molar in molar solution the quantity of the solute is comparatively greater.

But both one molal and one molar solutions contain one mole of solute particles. Thus, both solutions contain same number of particles (i.e. 6.02×10^{23}).

OR

Molarity:

Molarity is the measure of concentration of a solution in terms of moles solute per liter of solution.

Molality:

Molality is a measure of concentration in terms of moles solute per kg of solvent.

Reason:

As the concentration of the solution increases, the volume of the solute will start to become appreciable and will start to take a measurable amount of space that would otherwise be taken by the water. Whereas in the case of molality, the denominator is always the mass of the solvent (not solvent plus solute) so it will remain unchanged regardless of the concentration.

In practical terms:

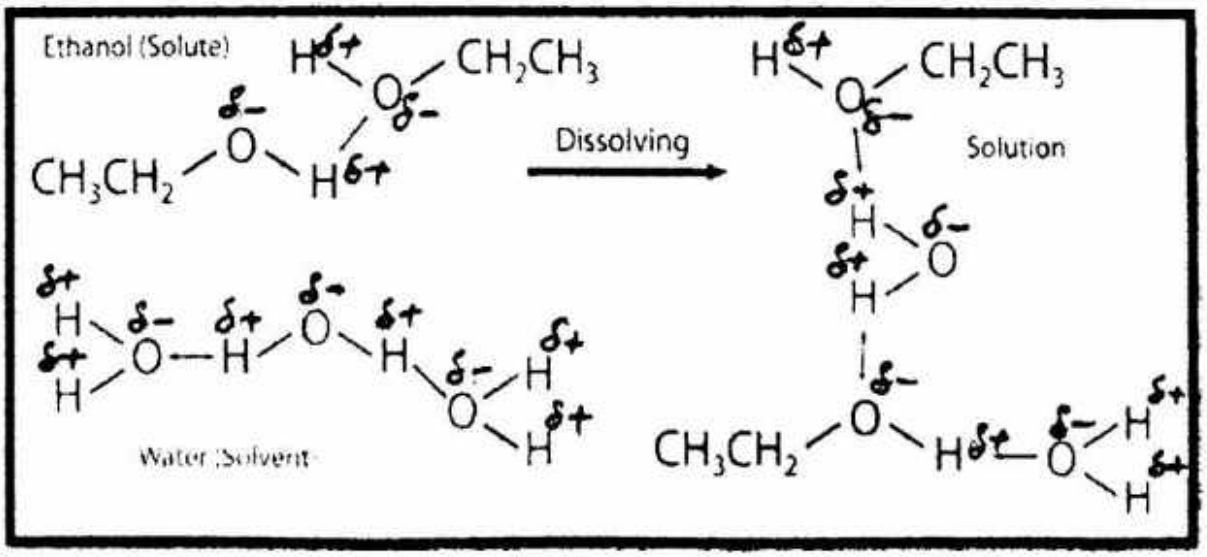
In practical terms, dilute solutions in water will have the same value of molarity and molality since the density of water is 1 g/mL, thus 1 liter of water will weigh 1 kg.

(c) The total volume of solution by mixing $50cm^3$ of ethanol and $50cm^3$ of water may not be equal to $100cm^3$, why?

The total volume of solution by mixing $50cm^3$ of ethanol and $50cm^3$ of water may not be equal to $100cm^3$

i. It is because of the phenomenon of association or dissociation of solute particles in solvent.

ii. Ethanol and water form strong hydrogen bonding due to which their molecules come closer and therefore the total volume of the solution decrease.



(d) NaCl and $NaNO_3$ are used to lower the melting point of ice.

The addition of NaCl and $NaNO_3$ decreases the vapour pressure of water due to which the freezing point of water decreases.

OR

When salt is added to ice it lowers the freezing point by decreasing the vapour pressure of the solution. Thus the solution will freeze at a lower temperature than that of the pure solvent.

(e) In summer the antifreeze solutions protect the radiator from boiling over.

Antifreeze lowers the vapour pressure due to which freezing point decreases and boiling point increases thus in summer due to the increase in boiling point the antifreeze solutions protect the radiator from boiling over.

(f) One molal and two molal solutions of urea boil at different temperatures.

As we know that colligative properties depends on the number of the solute particles and not on their nature.

As one molal solution contain one mole of particles and two molal solution contain two moles of particles therefore one molal and two molal sclutions of urea boil at different temperatures.

(g) Relative lowering of vapour pressure is independent of the temperature.

Relative lowering of vapour pressure, depends upon mole-fraction of solute in a solution.

According to Raoult's law

$$\frac{\Delta P}{P^o} = \text{Relative lower of V. P}$$

And

$$\frac{\Delta P}{P^o} = X_2 \quad \ldots\ldots\ldots (i)$$

Where X_2 = Mole fraction of solute

From equation (i) it is concluded that relative lowering of vapour pressure depends upon mole fraction and number of moles do not depend upon temperature. Thus relative lowering of vapour pressure is independent of the temperature.

(h) The sum of mole fractions of all the components is always equal to one.

Suppose a solution contain three components A,B and C with 5,8 and 10 moles respectively.

$n_A = 5$ mole

$n_B = 8$ mole

$n_C = 10$ mole

Total number of moles = $n_t = 23$ moles

Mole fraction of component A = $X_A = \frac{n_A}{n_t} = \frac{5}{23} = 0.217$

Mole fraction of component B = $X_B = \frac{n_B}{n_t} = \frac{8}{23} = 0.347$

Mole fraction of component C = $X_C = \frac{n_C}{n_t} = \frac{10}{23} = 0.43$

$X_A + X_B + X_C = 0.217 + .347 + 0.43 = 1$

24. 1.89g of an organic compound, A was dissolved per 85cm^3 of water (d=0.998 gcm^{-3}). The boiling point under one atmospheric pressure of this solution is increased to 100.106°C. What is the molecular weight of A. (Ans: 109.2978.4)

Solution:

Density of water = d =0.998 g cm^{-3}

Volume of water = V = 85 cm^3

Mass of water = W_1 = ?

$$density = \frac{Weight}{volume}$$

Mass of water = $W_1 = 0.998 \times 85 = 84.83$ g

Mass of an organic compound = $W_2 = 1.89$ g

Mass of water (solvent) = $W_1 = 84.83$ g

K_b for water = $K_b = 0.52°C$

Boiling point of water = $T_1 = 100°C$

Increase in the boiling point of water = $T_2 = 100.106°C$

Elevation of boiling point = $\Delta T_b = T_2 - T_1 = 100.106°C - 100°C = 0.106°C$

Molecular mass of organic compound = M_2 = ?

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

$$= \frac{0.52 \times 1.89 \times 1000}{0.106 \times 84.83}$$

$$= 109.297\ g\ mole^{-1}$$

25. What freezing point do you expect for water in which 17.9g sucrose $C_{12}H_{22}O_{11}$ is dissolved per 47.6g of H_2O. (Ans: -2°C)

Solution:

Mass of sucrose = $W_2 = 17.9$g

Mass of water = $W_1 = 47.6$g

K_f for water = $K_f = 1.86°C$

Molar Mass of sucrose ($C_{12}H_{22}O_{11}$) = $M_2 = (12 \times 12)+(1\times22)+(16\times11)=342$ gmole^{-1}

Depression of freezing point = ΔT_f = ?

As we know that,

$$\Delta T_f = \frac{K_f \times W_2 \times 1000}{M_2 \times W_1} = \frac{1.86 \times 17.9 \times 1000}{342 \times 47.6} = 2.045°C$$

Let,

Freezing point of pure water = $T_1 = 0°C$

Freezing point of solution = T_2 = ?

As we know that,

$$\Delta T_f = T_1 - T_2$$

or

$$T_2 = T_1 - \Delta T_f$$

$$= 0 - 2.045$$

$$= -2.045°C$$

26. The vapour pressure of pure water is 23.756 torr at 25°C. How much glucose would be added to 100g of water to bring the vapour pressure down to 23.00 torr? (Ans: 31.82g)

Solution:

Mass of glucose = W_2 = ?
Mass of water = W_1 = 100 g
Vapour pressure of pure water = P° = 23.756 torr
Vapour pressure of solution = P = 23 torr
Lowering of vapour pressure = $\Delta P = P^\circ - P = 23.756 - 23 = 0.756$ torr
Molar Mass of glucose ($C_6H_{12}0_6$) = $M_2 = 12\times6+1\times12+16\times6 = 180$ g mole^{-1}
Molar Mass of water (H_2O) = M_1 = 18 g mole^{-1}
As we know that

$$\frac{\Delta P}{P^\circ} = \frac{W_2 \times M_1}{W_1 \times M_2}$$

$$W_2 = \frac{\Delta P}{P^\circ} \times \frac{W_1 \times M_2}{M_1}$$

$$W_2 = \frac{0.756}{23.756} \times \frac{100\times180}{18} = 31.824g$$

27. You are provided 98%(W/W) H_2SO_4 having density 1.84 gcm^{-3}. How much volume of this H_2SO_4 is required to obtain one dm^3 of 30% (W/W) H_2SO_4 which has density 1.25g cm^{-3}. This solution is used in lead storage batteries as electrolyte.

Solution: For 98% sulphuric acid:

Density of sulphuric acid = 1.84 g/cm^3
Mass of 1 cm^3 of solution = 1.84 g
Mass of 1000 cm^3 of solution = 1.84 × 1000 = 1840 g

Therefore,
100 g of solution contain sulphuric acid = 98 g
1840 g of solution contain sulphuric acid = $\frac{98}{100} \times 1840 = 1803.2$ g.
Hence, 1803.2 g of H_2SO_4 are present in 1000 cm^3 of solution.
Mass of H_2SO_4 = m_{solute} = 1803.2 g
Molecular Mass of H_2SO_4 = M_{solute} = 98 g mole^{-1}
Volume of solution = V = 1000 cm^3 = 1 dm^3
Therefore, molarity is given by

$$\text{Molarity} = \frac{m_{solute}}{M_{solute}} \times \frac{1}{V} = \frac{1803.2}{98} \times \frac{1}{1} = 18.4M$$

For 30% sulphuric acid:

% of sulphuric acid = 30 %
Density of sulphuric acid = 1.25 g/cm^3
Thus
Mass of 1 cm^3 of solution = 1.25 g
Mass of 1000 cm^3 of solution = 1.25 × 1000 = 1250 g
Therefore
100 g of solution contain sulphuric acid = 30 g
1250 g of solution contain sulphuric acid = $\frac{30}{100} \times 1250 = 375g$

Hence, 375 g of H_2SO_4 are present in 1000 cm^3 of solution.

Mass of H_2SO_4 = m_{solute} = 375 g

Mol. Mass of H_2SO_4 = M_{solute} = 98 g $mole^{-1}$

Volume of solution = V = 1000 cm^3 = 1 dm^3

Therefore, molarity is given by

$$\text{Molarity} = \frac{m_{solute}}{M_{solute}} \times \frac{1}{V} = \frac{375}{98} \times \frac{1}{1} = 3.83M$$

To prepare 1 dm^3 (1000 cm^3) of 3.83 M H_2SO_4 from 18.4 M H_2SO_4, we have

Conc. H_2SO_4 **Dil. H_2SO_4**

$M_1V_1 = M_2V_2$

Or $V_1 = \frac{M_2}{M_1} \times V_2$

$$V_1 = \frac{3.38}{18.4} \times 1000 = 208.15cm^3$$

28. Caffeine is about 10 times more soluble in hot water than in cold water. Is ΔH°_{soln} of caffeine in water is positive or negative?

Ans: Caffeine is about 10 times more soluble in hot water than in cold water it means that it needs heat (energy) to dissolve. As an increase in temperature adding heat favours the endothermic processes therefore dissolve of caffeine in water is endothermic process.

Thus ΔH°_{soln} of caffeine in water is positive.

29. PCl_5 can react with itself in a reversible reaction to form a salt that contains the PCl_4^+ and PCl_6^- ions.

$$2PCl_5 \rightleftharpoons PCl_4^{1+} + PCl_6^{1-}$$

The extent to which this reaction occurs depends on the solvent in which it is run. Predict whether a non polar solvent such as CCl_4 favours the product or the reactants of this reaction. Predict what would happen to this reaction if we use a polar solvent.

Ans: Principle:

The basic principle of solubility is "Like dissolves like".

Nature of reactant and products:

PCl_5 is a non-polar species, while PCl_4^+ and PCl_6^- are polar ions.

Non polar solvent:

A non-polar solvent like CCl_4 will dissolve PCl_5 but it cannot dissolve ionic species, PCl_4^+ and PCl_6^-.

Polar solvent:

A polar solvent like water will dissolve ionic species, PCl_4^+ and PCl_6^-.
But it cannot dissolve PCl_5.

Conclusion:

Thus, in a non-polar solvent the equilibrium will be shifted towards reactant side. While in a polar solvent, the equilibrium will be shifted towards product side.

30. An aqueous solution of a compound boil at 102.4 °C. At what temperature will this solution freeze?

Solution:

According to relations of Elevation in Boiling point.

K_b

$$Molar\ mass = \frac{Mass\ of\ Solute}{Mass\ of\ Solvent} \times \frac{K_b}{\Delta T_b} \times 1000 \ldots\ldots\ldots\ldots (i)$$

According to relations of depression in Freezing point

$$Molar\ mass = \frac{Mass\ of\ Solute}{Mass\ of\ Solvent} \times \frac{K_f}{\Delta T_f} \times 1000 \ldots\ldots\ldots\ldots (ii)$$

On comparing equation (i) and (ii)

$$\frac{Mass\ of\ Solute}{Mass\ of\ Solvent} = \frac{K_b}{\Delta T_b} \times 1000 = \frac{Mass\ of\ Solute}{Mass\ of\ Solvent} = \frac{K_f}{\Delta T_f} \times 1000$$

$$\frac{K_b}{\Delta T_b} = \frac{K_f}{\Delta T_f}$$

$$\Delta T_f = K_f \times \frac{\Delta T_b}{K_b}$$

$$= 1.86 \times \frac{2.4}{0.52} = 8.5°C$$

$$Freezing\ point\ of\ solution = -8.5^\circ\ C$$

31. Water and carbon tetrachloride are not missible. When mixed they form two layers. If an aqueous solution of iodine I_2 is shaken with CCl_4, the iodine is extracted into the CCl_4 layer. Explain this behavior on the basis of your knowledge of intermolecular forces.

Ans: CCl_4 is a non-polar solvent while iodine (I_2) is a non-polar solute so they can dissolve in each other as we know that "**like dissolves like**". Water and carbon tetrachloride are not missible. When mixed they form two layers. If an aqueous solution of iodine I_2 is shaken with CCl_4, the iodine is extracted into the CCl_4 layer. It is because both of them are non-polar.

32. A Cucumber placed in concentrated brine (concentrated NaCl solution in water) shrivels into a pickle. Explain?

Ans: This happens due to the process of osmosis.

Reason:

The cucumber shrivels because the moisture (water) is extracted from it. This happens if the salt concentration in the brine solution is greater than the salt concentration in the cells of the cucumber. Osmosis will tend to equalize the concentrations, removing water from the cucumber and adding salt from the solution. Thus the brine solution is hypertonic (more concentrated) relative to the cucumber cells.

OR

Cucumber cells membranes are semi permeable membranes. On placing that in concentrated brine process of Osmosis takes place and water start to flow from higher concentration (in cucumber) to lower concentration (in brine). As a result, the cucumber shrivels.

33. Two beakers are placed in a sealed bell jar. One beaker contains water and the other contained concentrated glucose solution. With time volume of water decreases and solution volume increases. Explain why?

Ans: As water is a pure solvent, so its rate of evaporation is high than that of glucose solution. Water vapours go into the atmosphere of bell jar (A bell jar is a sealed laboratory equipment). Therefore, the water vapours are transferred from high pressure area to low pressure area. Thus, volume of water decreases and that of solution increases.

34. Consider two aqueous solutions one is sucrose and the other of glucose. Both of these solutions boil at 101.52 ^{0}C. List some common properties of these solutions.

Ans: Both these solutions have

i. Same number of particles.
ii. Same vapour pressure
iii. Same freezing point.
iv. Same boiling point.
v. Same molality.
vi. Same osmotic pressure.