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 50 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 quid 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 59.5°C. the composition of both the layers become identical Each layer contains 34% phenol and 86% 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 salvation. When water is the so vent, this process is known as hydration.

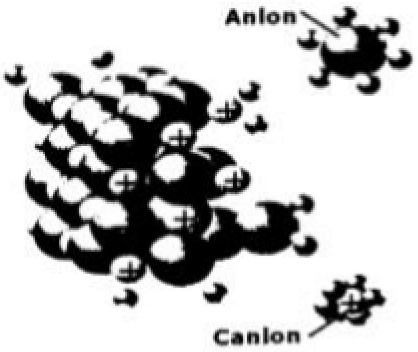
Ionic solids:

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 containsNH_4^+ 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^+icns 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:

Het 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 $CaCl_2$ and $MgSO_4$ for hot packs.

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

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

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.

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

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, ion. Thus Cu^{++} ion has greater ability to attract water molecules than SO ion. That is why, out of five water of hydration four are associated with Cu^{++} ion and only one with SOA ion. In hydrates, a though new bonds between ions and water molecules are formed but no hydrogen oxygen bond of water is broken.

Example 10.1:

250g of $CuSO_4$ $5H_2O$ on heating produced 159.82g CuSO4. Calculate the percent of water in the $CuSO_4xH_2O$ 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

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

X=no. of moles = $\frac{\text{Mass of water in the hydrate}}{\text{molar mass of water}}$

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

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 of solution or in terms of number of moles of solute per dm 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.

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

In a saturated solution dissolving of solute in fact 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."

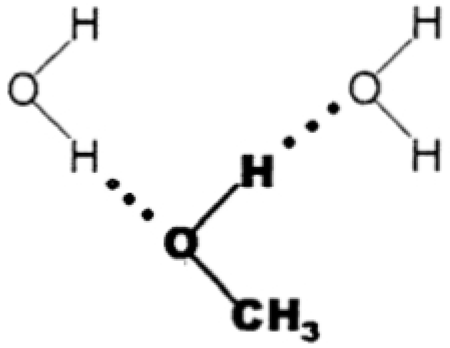
Substances which have similar structures and intermolecular forces tend to be soluble.

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 of 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*c and then surrounds the resulting Na' and C ions These ions are called hydrated ions.

ii. When NaCl crystals are placed in CCl, or CH Non-polar molecules of these liquids are unable to attractions 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, or benzene, C.Hg 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, 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

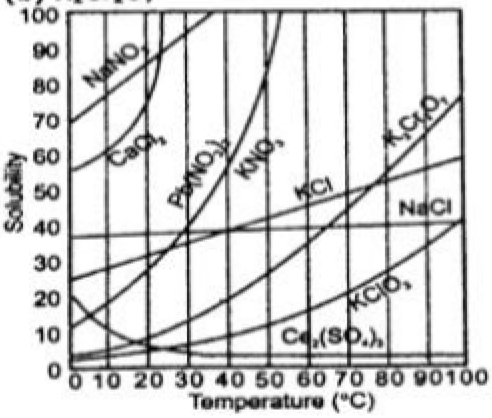
1. $\Delta H_{solution}$ of Ce:(S04), and that of $KClO_3$is positive or negative.

2. $\Delta H_{solution}$ of CaCl2 is greater or lesser than that of NaNO, at 20°C.

3. Indicate the temperature at which solubilities of Naci 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:

Ta exothermic reaction $\Delta H_{solution}$ = negative

In endothermic reaction $\Delta H_{solution}$ = positive

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 AH 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 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.

indicate the temperature at which solubilities of NaCl and KCl are same.

iii. $\Delta H_{solution}$ 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.

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

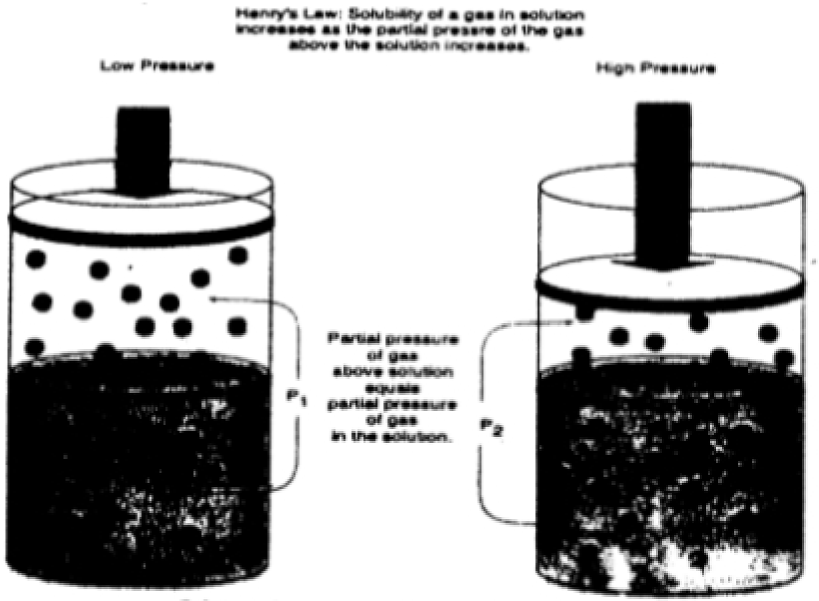
Ans: Effect of Pressure on Solubility:

Pressure has negligible effect on the solubilities 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 = $\frac{grams\ of\ sugar}{grams\ of\ solution} \times 100$

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

= 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.

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 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

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

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

volume of solute (cm)

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

Q19. Differentiate between molarity and molality.

Ans: Difference between molarity and molality:

Molarity (M):	**Molality (m)**
1. His defined as the number of moles of solute dissolved pe' dm' 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	iii. Molality is moles per Kilogram/s of
$M=\frac{moles\ of\ solute}{dm^3\ of\ solution}$ $M=\frac{grams\ of\ solute}{molar\ mass\ of\ solute \times dm^3\ of\ solution}$	$M=\frac{moles\ of\ solute}{kg\ of\ solution}$
Example: If 0.5 mole of NaOH (209) is dissolved in one kilogram of water, the dm of solution,0.5 molar or 0.5 M NaOH solution is obtained	Example: When 58 59 NaCl (1 mole) is dissolved in enough water to make one solution would be one molal or1m NaCl solution.

Example 10.3:

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

Solution:

Mass of $NaHCO_3$ = 0.25g

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

= 72g/mole

Volume of solution = $100cm^3$

=100/1000 = $0.1dm^3$

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

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

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 (C2H,OH) with 100g of water. Calculate molality of this solution.

Solution:

=1.00g

=12x2+1x6+16=46g/mole

$= \frac{1.00g}{46g / mole}$

=2.17x10^{-2} mole

=100g=100/1000 = 0.1kg

=moles of C_2H_5OH / Kg of water

=2.17 x 10^{-2} mole / 0.1 kg

=0.217m.

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 cm³ of solution and the other of 9.0-gram NaCl per 100 cm³ of solution. Calculate the molarity of these intravenous solutions.

(i) 5% dextrose means 5g glucose in 100 cm³ solution.

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

Molecular Mass of glucose $(C_6H_{12}O_6) = 12 \times 6 + 1 \times 12 + 16 \times 6$

$$= 180 \text{ g mole}^{-1}$$

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

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

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

(ii) 9% normal saline; 9g NaCl in 100 cm³ solution.

$$\text{Mass of NaCl} = 9 \text{ g}$$

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

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

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

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

Example 10.5:

An aqueous solution containing 100g ethanol per dm³ of solution has a density of 0.984 g cm⁻³. Calculate mole fraction of each component of solution.

Solution:

$$\text{Volume of solution} = 1\text{dm}^3 = 1000 \text{ cm}^3$$

$$\text{Mass of ethanol} = 100 \text{ g}$$

$$\text{Mass of solution} = \text{Density } \times \text{Vol of solution of solution}$$

$$= 0.984 \text{ of cm}^{-3} \times 1000 \text{ cm}^3$$

$$= 984 \text{ g}$$

$$\text{Mass of water (solvant)} = 984\text{g} - 100\text{g} = 884 \text{ g}$$

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

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

$$\text{Total No. of moles in solution} = 49.1 + 2.17 = 51.27\text{ mole}$$

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

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

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?

Solution:

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

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

$$= 0.35\text{ ppm}$$

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:

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

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

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

$$= 0.5 \times 10^3 \times 10^{-9}$$

$$= 5 \times 10^{-7}\text{ g}$$

Note: ppb; parts per billion

SELF-CHECK EXERCISE 10.4

1. **Calculate the ppm by mass of calcium in a 2.5g tablet contains 500mg calcium.**

(Ans: 0.2 ppm)

2. **Waste water from a cement factory contains 22mg Ca^{+2} ions and 0.006 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. $\mathbf{Mass\ of\ Calcium = 500mg = \frac{500}{1000} = 0.5}$

$\text{Mass of tablet} = 2.5\text{g}$

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

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

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

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

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

$$\text{Concentration of Ca}^{+2}\text{ ions in ppm} = \frac{\text{Mass of Ca}^{+2}}{\text{Mass of solution}} \times 10^6$$

$$= \frac{0.22\text{ g}}{1000\text{ g}} \times 10^6 = 22\text{ ppm}$$

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

$$= \frac{0.006\text{ g}}{1000\text{ g}} \times 10^6 = 6\text{ ppm}$$

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

Mass of gold recover = 1g

Mass of water = ?

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

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

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

$$=0.917 \times 10^2 \times 10^9\text{g}$$

$$=0.917 \times 10^{11}\text{g}$$

$$=9.17 \times 10^{10}\text{g}$$

$$=9.17 \times 10^{7}\text{kg}$$

Example 10.8:

Calculate the molarity of 15% (w/w) glucose ($C_6H_{12}0_6$) solution.

Solution:

Since solution of glucose is 15% w/w

Mass of glucose $= 15\ \text{g}$

Mass of solution $= 100\ \text{g}$

Mass of water $= 100\ \text{g} - 15\ \text{g} = 85\ \text{g}$

$$= \frac{85}{1000} = 0.085\ \text{kg}$$

Molar mass of glucose $= 12 \times 6 + 1 \times 12 + 16 \times 6 = 180 \frac{\text{g}}{\text{mole}}$

Molarity $= \dfrac{150\ \text{g}}{180 \frac{\text{g}}{\text{mole}} \times 0.085\ \text{kg}}$

Molarity $= 0.98$ moles per kg.

Molarity $= 0.98\ \text{m}$

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 fiber 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 $H_2SO_4 = 1 \times 2 + 32 \times 1 + 16 \times 4 = 98 \frac{\text{g}}{\text{mole}}$

Percentage of H_2SO_4 solution $= 98\%$

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

1 cm^{3}of H_2SO_4 solution contains 1.84g H_2SO_4 solution.

1000 cm^{3}of H_2SO_4 solution would have $= 1.84 \times 1000 = 1840\ \text{g}$ H_2SO_4 solution

of this total mass 98% is H_2SO_4

$$\text{Mass of } H_2SO_4 \text{ in solution} = \quad 1840 \times \frac{98}{100} = 1803.2 \text{ moles}$$

$$\text{Moles of } H_2SO_4 \text{ in solution} = \quad \frac{1803.2\text{g}}{98\frac{\text{g}}{\text{moles}}} = 18.4 \text{ moles}$$

Since solution has 18.4 moles of H_2SO_4 per 1000 cm^3 or 1 dm^3 of solution, the molarity of this solution in 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:

$$\text{Molarity of HCl} = 12 \text{ M}$$

$$\text{As} \quad d \quad = \frac{m}{v}$$

$$m = d \times v$$

Thus mass of HCl solution present it

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:

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

Total volume of solution $= 250\ \text{cm}^3 + 250\ \text{cm}^3$

$$= 500\text{cm}^3 = 0.5\ \text{dm}^3$$

Molrity of resulting solution $= \frac{0.5\ \text{moles}}{0.5\ \text{dm}^3} = 1\text{M}$

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 manufacturing of liquid soaps, in paints and varnish removers. An experimenter needs 25 cm^3 of 0.015M KOH. What mass of KOH, he will dissolve to make this solution? (Ans: 0.021 g)**

Solution:

1. **Molarity of Urea** $= \mathbf{0.25\ m}$

Number of moles of Urea $= n_2 = 0.25\ \text{m}$

Mass of water $= 1\ \text{kg} = 1000\ \text{g}$

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

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

Mole fraction of Urea $= X_2 = ?$

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

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

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

2. **Volume of solution** $= \mathbf{25\ cm^3 = 0.025\ dm^3}$

Molarity of solution $= M = 0.015\ M$

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

Mass of KOH $= ?$

As,

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

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

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

Q 20. State and derive Raoult's law.

Ans:

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

$$P \quad \propto \quad X_1$$

$P = P^{o}X_1$ .... (1)

Where P^{o} is the 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

$P = P^{o}(1 - X_2)$

$P = P^{o} - P^{o}X_2$

$P - P^{o} = -P^{o}X_2$

$P^{o} - P = P^{o}X_2$ .... (2)

$\Delta P = P^{o}X_2$

Where $P^{o} - P$ = 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 traction ot solute.

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

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

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

$\frac{P^{o} - P}{P^{o}}$ 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.

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

i. Raoult's Law also governs the vapor pressure of mixture of two or more volatile component.
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 pressure 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^{0}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:

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

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

Q 22. Derive equation for lowering of vapour pressure from Raoult's law and how molar mass of a non-volatile 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.

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

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,

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

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

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

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

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

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

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

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 86.7 torr. Calculate the molecular mass of M.

Solution:

P^o = 98.6 torr

P = 86.7 torr

$\Delta P = P^o - P$

$= 98.6 - 86.7 = 11.9 \text{ torr}$

$W_2 = 106.3\text{g}$

$W_1 = 863.5\text{g}$

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

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

$= 79.5 \text{ g/mole}$

Q 23. 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 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.

Q 24. Give quantitative and graphical explanation of boiling point elevation.

Ans: Quantitative Aspects of Boiling Point Elevation:

When one mole of non-electrolyte and non-volatile solute is dissolved in 1 kg of water boiling point is raised 0.52^0C. This value is known as the molal boiling point elevation constant (ebullioscopy constant) tor 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^o.
IV. The curve CD represents variation in vapour pressure of solution with temperature.
V. This curve must lie below that of pure solvent.
VI. This is because the vapour pressure of solution at all temperature 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^o.

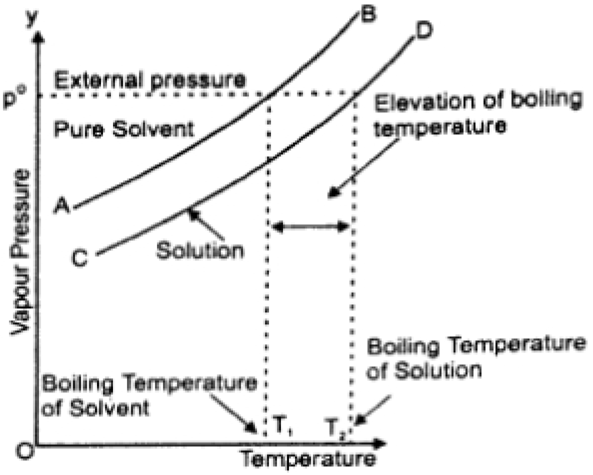


Figure 1: Elevation of boiling point-temperature curve

Mathematically:

The difference of two boiling points gives the elevation of boiling point (ΔT_0)

$$\Delta T_b \quad = \quad T_2 - T_1$$

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

$$\Delta T_b \propto \quad m$$

$$\Delta T_b \quad = \quad K_b m \text{ ....}\mathbf{(1)}$$

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

$$\Delta T_b \quad = \quad K_b$$

Thus, elevation of boiling point when 1 mole 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 molarity m of the solution containing W_2g of solute of molecular mass M_2 dissolved in W_1g of 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: 0.465^0C)

Solution:

$$\text{Molarity of solution} = 0.25\ \text{m}$$

$$K_f \text{ for water} = K_f = 1.86^0C$$

$$\Delta T_f = K_f \times m = 1.86 \times 0.25 = 0.465^0C$$

Let,

$$\text{Freezing point of pure water} = T_1 = 0^0C$$

$$\text{Freezing point of solution} = T_2 = ?$$

$$\Delta T_f = T_1 - T_2$$

$$T_2 = T_1 - \Delta T_f$$

$$= 0 - 0.465 = -0.465^oC$$

Q 25. Describe lands Berger's method to measure boiling point elevation.

Ans: Lands Berger's Method:

The apparatus used is shown in figure. 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 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 up to 0.01K

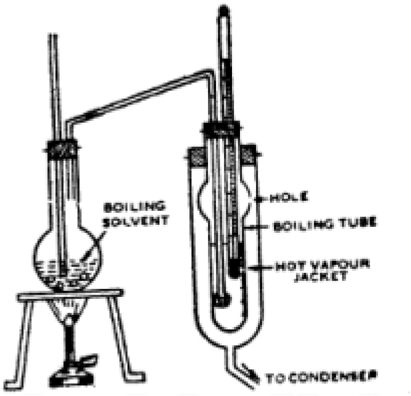


Figure 2: LandsBeger's apparatus for measurement of elevation of boiling points.

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 b 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 5.5^0C. 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 6.07^0C. Now stop the supply of benzene vapours. Remove thermometer and rose head from the inner tube and note the volume of solution. I will be 12.15 cm^3. Determine density of this solution. Its density will be 0.897 g 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 Cholestrol $= W_2 = 4.5\text{ g}$

Mass of Benzene $= W_1 = ?$

Boiling Point of Benzene $= T_1 = 5.5^\circ\text{C}$

Boiling Point of Solution $= T_2 = 6.07^\circ\text{C}$

K_o for Benzene $= 5.12$

Volume of solution $= V = 121.5\text{ cm}^3$

Density of solution $= d = 0.897\text{ g cm}^{-3}$

Molecular mass of cholestrol $= M_2 = ?$

As.

$$d = \frac{m}{v}$$

$$M = d \times V$$

Mass of Solution $= 0.897 \times 121.5 = 108.99\text{g}$

Mass of Benzene(W_1) $=$ Mass of solution $-$ Mass of Cholestrol

$$W_1 = 108.99 - 4.5 = 104.486\text{ g}$$

As,

$\Delta T_b = T_2 - T_1$

$\Delta T_b = 6.07 - 5.5 = 0.57^\circ\text{C}$

Now

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

SELF-CHECK EXERCISE 10.7

Calculate the boiling point of a solution containing 12.5g of benzoic acid, $C_7H_6O_2$ in 110g of benzene. Boiling point and K_b of benzene are 80.1^{0}C and 2.53 respectively.

(Ans: 82.637^{0}C)

Solution:

Mass of benzoic acid $= 12.5\ \text{g}$

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

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

Mass of Benzene $= 110\ \text{g}$

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

Boiling Point of Solution $= T_2 = ?$

K_o for Benzene $= 2.53$

As we know that,

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

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

Thus,

$\Delta T_b = T_2 - T_1$

$T_2 = T_1 + \Delta T_b$

$= 80.1 + 2.357 = 82.457^\circ\text{C}$

Q 26. Write quantitative and graphical explanation of freezing point depression for a solution.

Ans: Quantitative Aspects of Freezing Point Elevation:

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

Graphically:

Figure shows how the vapour pressure influences the freezing point

- **i.** Curve AB is for the pure solvent.
- **ii.** The solvent freezes at temperature T_1 corresponding to the point B when its vapour pressure is P^0
- **iii.** The portion of the cure BC is for the solid solvent.

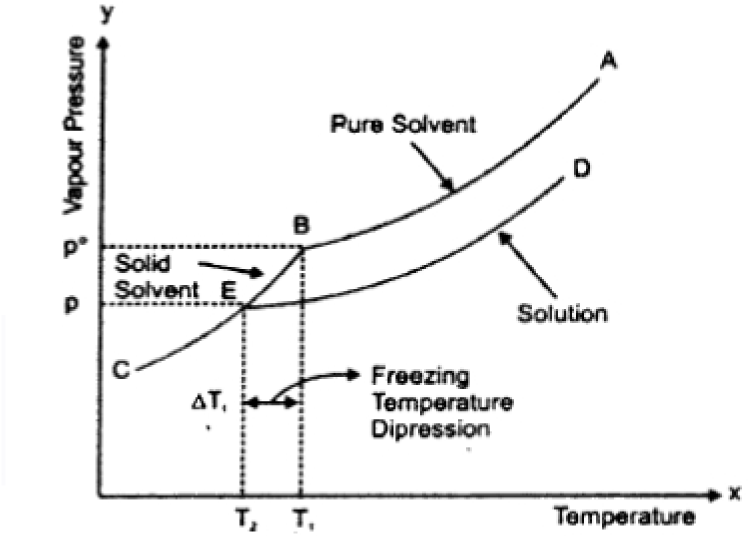


Figure 3: Depression of freezing point - temperature curve

- **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 cure BC at point E which is freezing point of solution T_2 when its vapour pressure is P which is lower than P^0.
- **vii.** This is because vapour pressure of solution is always less than that of the pure solvent.

Mathematically:

The difference between the two freezing points gives the depression of the freezing point ΔT_1.

$$\Delta T_1 = T_1 - T_2$$

Experimentally $\Delta T_1 \propto m$

$$\Delta T_1 = m K_1 \text{....................}\mathbf{(1)}$$

Where K_1 is constant called the molal freezing point constant or the cryoscopic constant for 1 molal solution.

$$\Delta T_1 = K_1$$

Thus, the depression of freezing point when 1 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_1 depends upon the nature of the solvents and 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 molarity 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 eq (1)

we have

$$\Delta T_1 = \frac{K_1 \times W_2 \times 1000}{M_2 \times W_1}$$

Rearranging

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

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

Q 27. Describe the method use for the measurement of freezing pint depression.

Ans: Measurement of Freezing Point Depression:

Beckmann's Method:

The apparatus used is shown in figure.

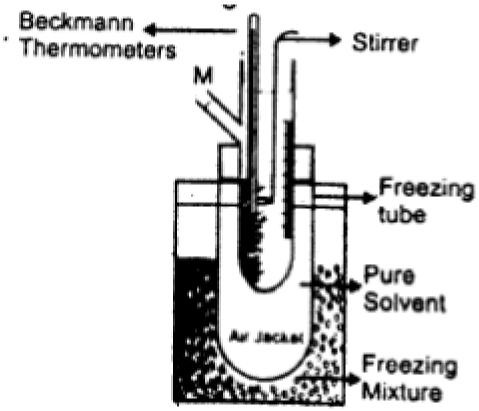


Figure 4: Beckmann's apparatus for measurement of depression of freezing point.

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.

Q 28. 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 non-volatile 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 overheating.

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 lowering 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 immerses 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 3.62 g 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^{0}C. Find the difference of the two freezing points. This will give the value of ΔT_1. From this data determine molecular mass of nicotine.

Solution:

Mass of Nicotine $= W_2 = 3.62\ \text{g}$

Mass of water $= W_1 = 73.4\text{g}$

K_f for water $= 1.86$

Feezing point of water $= T_1 = 0^{\circ}C$

Feezing point of Solution $= T_2 = -0.563^{\circ}C$

Molecular mass of nicotine $= M_2 = ?$

$$\Delta T_f = T_1 - T_2$$

$$= 0 - (-0.563) = 0.563^{\circ}C$$

Now,

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

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

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

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^{0}C) has a freezing point of 4.850^{0}C. K_f for benzene is 5.12. what is the molecular weight of Vitamin K_1?**

   (Ans:450.596 $gmole^{-1}$)

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^{0}C, B.P = 103.35^{0}C)**

Solution:

1. **Mass of Benzene $= \mathbf{1.048\ g}$**

K_f for benzene $= 5.12$

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

Feezing point of water $= T_1 = 5.455°C$

Feezing point of solution $= T_2 = 4.850°C$

Molecular mass of Vitamin $K_1 = M_2 = ?$

As we know that

$$\Delta T_f = T_1 - T_2$$

$$= 5.455 - (4.850) = 0.605°C$$

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

$$M_2 = \frac{5.12 \times 0.0558 \times 1000}{0.065 \times 1.048} = 450.6\ g\ mole^{-1}$$

2.

Mass of water = 5kg = 5000g

Mass of ethylene glycol = 2kg = 2000g

Boiling point of water = $T_1 = 100^oC$

K_b for benzene = 0.52

Boiling point of solution = T_2 =?

Molecular mass of ethylene glycol ($C_2H_6O_2$) = $M_2 = 12 \times 2 + 1 \times 6 + 16 \times 2$

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

As we know that

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

$$\Delta T_b = \frac{2.52 \times 2000 \times 1000}{62 \times 5000}$$

$$= 3.355^oC$$

Thus,

$$\Delta T_b = T_2 - T_1$$

$$T_2 = T_1 + \Delta T_b$$

$$T_2 = 100 + 3.355 = 103.355^oC$$

For Freezing point of solution:

Mass of water = 5kg = 5000g

Mass of ethylene glycol = 2kg = 2000g

Boiling point of water = $T_1 = 100^oC$

K_f for water = 1.86

Boiling point of solution = T_2 =?

Molecular mass of ethylene glycol $(C_2H_6O_2)$ $= \mathrm{M_2} = 12 \times 2 + 1 \times 6 + 16 \times 2$

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

As we know that,

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

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

$$\Delta\mathrm{T_f} = \frac{1.86 \times 2000 \times 1000}{62 \times 5000}$$

$= 12^oC$

Thus,

$$\Delta\mathrm{T_f} = \mathrm{T_1} - \mathrm{T_2}$$

$$\mathrm{T_2} = \mathrm{T_1} - \Delta\mathrm{T_f}$$

$$\mathrm{T_2} = 0 - 12 = -12^oC$$

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.

Q 29. 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).

$$\pi = MRT$$

where,

M is molarity

R= 0.08295746 L atm K^{-1} mol^{-1} is the gas constant

T is the thermodynamic (absolute) temperature

Examples:

i. Certain animal membrane 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.

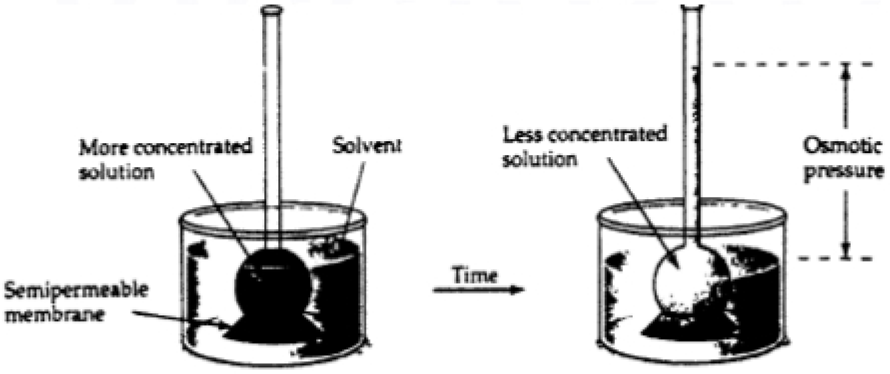


Figure 5: The phenomenon of osmosis

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 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 some osmotic pressure are called isotonic solution.
v. If concentration of NaCl solution is greater than 0.95% water flow 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.

Q 30. 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) Vapour pressure
(c) Freezing point
(d) Osmotic pressure

Solution:

Colligative properties depend upon the number of solute particles as the molarity of urea and glucose is same i.e 0.02 m thus they have same number of particles.

So, for both solutions boiling points, freezing points and osmotic pressures will remain same.

Q 31. 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 re bigger in size than those of true solutions.
- **iv.** These particles are visible in electron microscope but invisible in ordinary microscope.

Examples:

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

Q 32. Define lyophilic molecules.

Ans: Lyophilic molecules:

- **i.** When a protein crystal is dropped into water.
- **ii.** The polymer molecules dissolve to produce a colloidal dispersion.
- **iii.** Such colloidal dispersion in which dispersed phase shows an affinity or attraction for the dispersion medium are called lyophilic (means solvent loving).
- **iv.** The molecules of dispersed phase are called lyophilic molecules.

Examples:

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

Q 33. Define lyophobic.

Ans: lyophobic:

- **i.** When solid AgCl is brought in contact with water, it does nor 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 dispersion 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.

Q 34. Define suspensions and also give its examples.

Ans: Suspensions:

- **i.** A dispersion 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 reflected light.

Examples:

Mud or sand in water are suspensions.

Q 35. 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.
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.

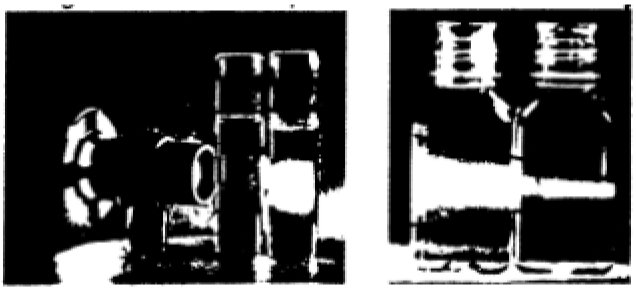


Figure 6: Light scattering by colloids suspensions

4. Colloidal particles do not coagulate into larger particles and settle out. This is due to the fact that particles have changed 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 paper.

8. Colloidal particles have high ration of surface area to volume as compared to the particles of true solutions.
9. Colloidal particles do not settle under the influence of gravity.

Q 36. 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.

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 another liquid	Milk, mayonnaise
Foam	Bubbles of a gas suspended in a liquid or a solid	Canned shaving cream, soap leather, whipped cream

Table 1: Types of colloids

Q 37. 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.

Table2: 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	0.1-1 nm
2.	Phase	Heterogenous	Heterogenous	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	Particles cannot pass through	Particles can pass through ordinary as

		paper, but cannot pass through ultra-filter paper.	ordinary as well as ultra-filter paper.	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 proportional to the mole fraction of the solvent. (X_1) $P=P^0X_1$
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 molality, which is defined as the number of moles of solute dissolved per dm^3 of solution.
5. Sometimes it is useful to know the ration 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 depends 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 depressions 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 heterogenous 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 dispersed in which particles of the dispersed substance are visible to the naked eye is called suspensions.