12TH CBSE CHEMISTRY

SOLUTIONS/

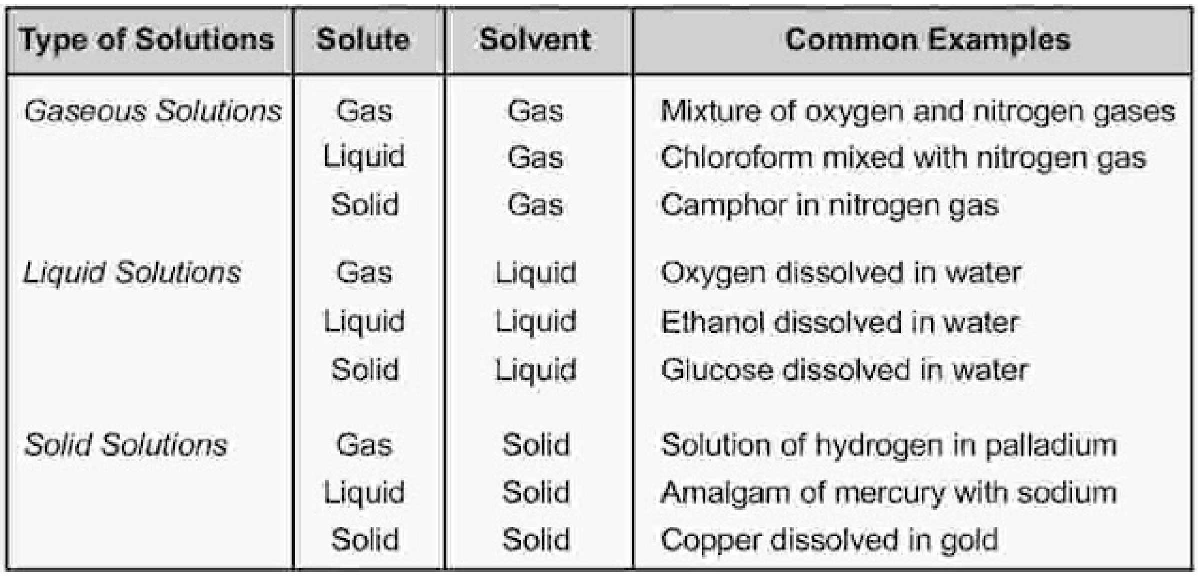
Introduction:

In this chapter, we will discuss about liquid solutions and their formation. This will be followed by studying the properties of solutions, like vapour pressure and colligative properties. We will begin with types of solutions and expressions for concentration of solutions in different units.

Thereafter, we will state and explain Henry's law and Raoult's law, distinguish between ideal and non-ideal solution and deviation of real solutions from Raoult's law. We will also discuss abnormal colligative properties alongwith association and dissociation of solute.

Types of Solutions

All the three states of matter (solid, liquid and gas) may behave either as solvent or solute. When a solution is composed of only two chemical substances, it is termed as binary solution. Depending upon the state of solute or solvent, binary solutions can be classified as



Type of Solutions	Solute	Solvent	Common Examples
Gaseous Solutions	Gas	Gas	Mixture of oxygen and nitrogen gases
	Liquid	Gas	Chloroform mixed with nitrogen gas
	Solid	Gas	Camphor in nitrogen gas
Liquid Solutions	Gas	Liquid	Oxygen dissolved in water
	Liquid	Liquid	Ethanol dissolved in water
	Solid	Liquid	Glucose dissolved in water
Solid Solutions	Gas	Solid	Solution of hydrogen in palladium
	Liquid	Solid	Amalgam of mercury with sodium
	Solid	Solid	Copper dissolved in gold

Some Important Definitions

- **ixture** - When two or more chemically non-reacting substances are mixed, they form mixture.
- **eterogeneous Mixture** - It consists of distinct phases, and the observed properties are just the sum of the properties of individual phases.
- **omogeneous Mixture** - It consists of a single phase which has properties that may differ from one of the individual components.
- **olution** - The homogeneous mixture of two or more components such that at least one component is a liquid is called solution.
- **olvent** - It is the constituent of solution which has same physical state as that of solution and generally present in greater amount than all the other components.

- **olute** - The component of a solution other than solvent is called solute, may or may not have same physical state as that of solution. Generally it is in smaller amount.

  Example- In a sugar syrup (liquid solution) containing 60% sugar (solid) and 40% water (liquid), water is termed as solvent, due to same physical state as that of solution.

Expressing the Strength of Solution

For a given solution the amount of solute dissolved per unit volume of solution is called concentration of solute. Strength of solution is the amount of solute in grams dissolved in one litre of solution. It is generally expressed in g/litre.

Other methods of expressing the strength of solution are:

1.

ass percentage –

$$\text{Mass \% of solute} = \frac{\text{Mass of solute}}{\text{Total mass of solution}} \times 100$$

$$\text{Mass \% of solvent} = \frac{\text{Mass of solvent}}{\text{Total mass of solution}} \times 100$$

2.

olume percentage –

$$\text{Volume \% of solute} = \frac{\text{Volume of solute}}{\text{Total volume of solution}} \times 100$$

$$\text{Volume \% of solvent} = \frac{\text{Volume of solvent}}{\text{Total volume of solution}} \times 100$$

3.

olality (m) -It is no. of moles of solute dissolved in 1 kg of the solvent.

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

4.

olarity (M) - It is no. of moles of solute dissolved in 1 litre of solution.

$$M = \frac{(\text{Number of moles of solute})}{(\text{Mass of soution \{in litre\}})}$$

5.

ormality (N) - It is no. of gram-equivalents of solute dissolved in 1 litre of solution

$$N = \frac{(\text{Number of gram equivalents of solute})}{(\text{Volume of solution \{in litre})}$$

6.

ormality - Ionic solutes do not exist in the form of molecules. These molecular mass is expressed as Gram-formula mass. Molarity for ionic compounds is actually called as **formality.**

7.

ole fraction –

$$\text{Mole fraction of solute} = \frac{(\text{Number of moles of solute})}{(\text{Total moles of solution})}$$

$$\text{Mole fraction of solvent} = \frac{(\text{Number of moles of solvent})}{(\text{Total moles of solution})}$$

For a binary solution,

mole fraction of solute + mole fraction of solvent = 1.

8.

arts per million (ppm)–

It is defined in two ways

$$\text{ppm} = \text{mass fraction} \times 10^6$$

$$\text{ppm} = \text{mole fraction} \times 10^6$$

Solubility

Solubility of a substance is its maximum amount that can be dissolved in a specified amount of solvent at a specified temperature. It depends upon the nature of solute and solvent as well as temperature and pressure. Let us consider the effect of these factors in solution of a **solid** or a **gas** in a **liquid**.

1.

olubility of Solid in Liquid

A solute dissolves in a solvent if the intermolecular interactions are similar in them, i.e., like dissolves like. Polar solute dissolves in polar solvent and non-polar solute in non-polar solvent. For e.g., sodium chloride and sugar dissolves readily

in water and napthalene and anthracene dissolves readily in benzene.
Solute + Solvent ⟶ Solution

i. **issolution:**When a solid solute is added to the solvent, some solute dissolves and its concentration increases in solution. This process is called **dissolution**.

ii. **rystallization**: Some solute particles collide with solute particles in solution and get separated out. This process is called **crystallization**.

iii. **aturated solution**: Such a solution in which no more solute can be dissolved at the same temperature and pressure is called a **saturated solution.**

iv. **nsaturated solution:** An unsaturated solution is one in which more solute can be dissolved at the same temperature.

v. **ffect of temperature**: In general, if in a nearly saturated solution, the dissolution process is endothermic, the solubility should increase with rise in temperature, if it is exothermic, the solubility should decrease with rise in temperature.

vi. **ffect of pressure**: Solids and liquids are highly incompressible, so pressure does not have any significant effect on solubility of solids and liquids.

vii. **upersaturated solution**: When more solute can be dissolved at higher temperature in a saturated solution, then the solution becomes supersaturated.

2.

olubility of Gas in Liquid

All gases are soluble in water as well as in other liquids to a greater or lesser extent. The solubility of a gas in liquid depends upon the following factors Nature of the gas, Nature of solvent, Temperature and Pressure.
Generally, the gases which can be easily liquified are more soluble in common solvents. For e.g., CO2 is more soluble than hydrogen or oxygen in water. The gases which are capable of forming ions in aqueous solutions are much more soluble in water than other solvents. For e.g., HCl and NH3 are highly soluble in water but not in organic solvents (like benzene) in which they do not ionize.

i.

ffect of temperature: Solubility of most of the gases in liquids decreases with rise in temperature. In dissolution of a gas in liquid, heat is evolved and thus this is an exothermic process. The dissolution process involves dynamic equilibrium and thus follows **Le Chatelier's** principle. As dissolution is exothermic the solubility of gas should decrease with rise in temperature.

ii.

ffect of pressure: Henry's law: At constant temperature, the solubility of a gas in a liquid is directly proportional to the pressure of the gas.

p = KH x,

KH = Henry's law constant.

Applications of Henry's law

1.

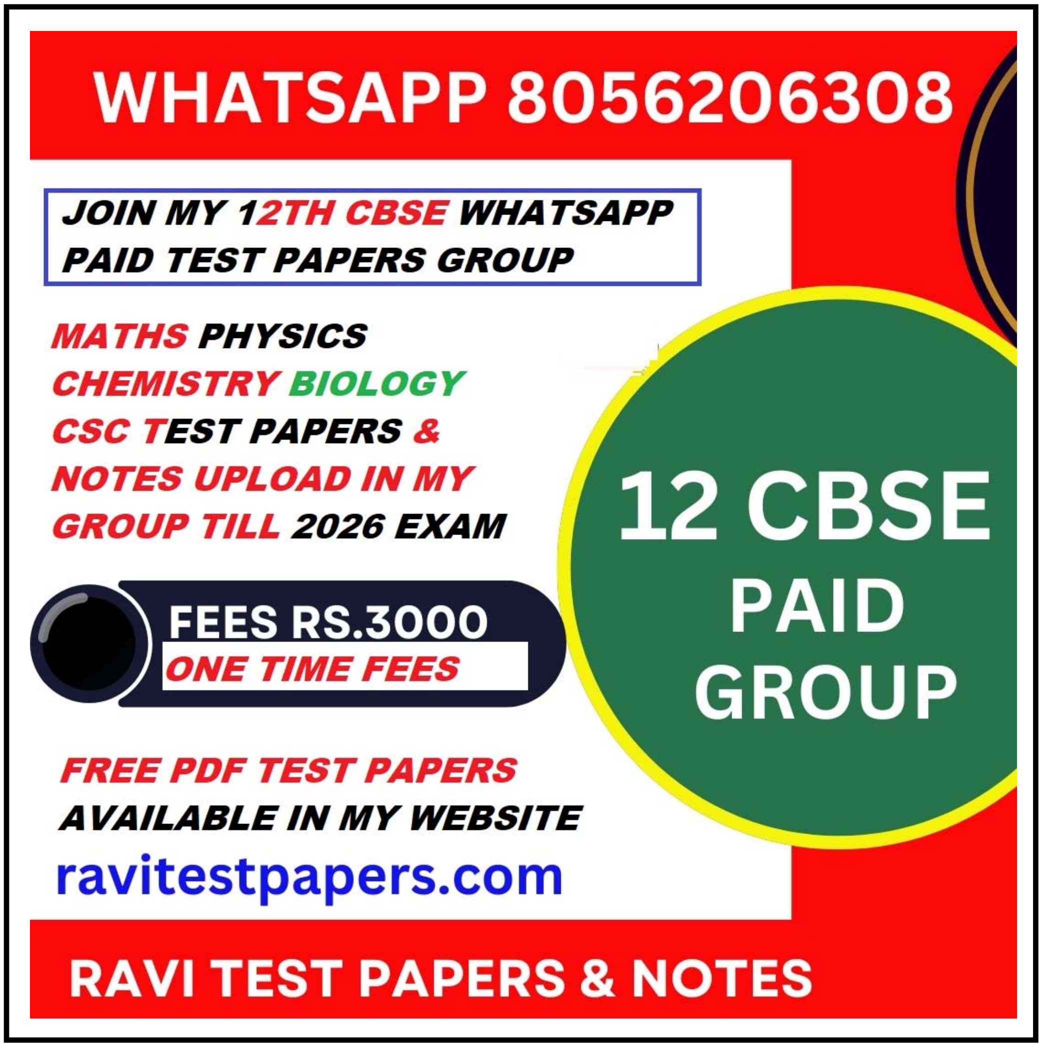
n manufacture of soft drinks and soda water, CO_2 is passed at high pressure to increase its solubility.

2.

o minimise the painful effects accompanying the decompression of deep sea divers. O_2 diluted with less soluble. He gas is used as breathing gas.

3.

t high altitudes, the partial pressure of O_2 is less then that at the ground level. This leads to low concentrations of O_2 in the blood of climbers which causes ‘**anoxia**’.



Vapour Pressure of Solution

It is the pressure exerted by vapour on the surface layer of liquid at equilibrium between vapour and liquid.

Factors affecting Vapour Pressure

i. **Nature of liquid -** Liquid with higher intermolecular attraction forces form less amount of vapour and hence lower vapour pressure and vice-versa.

ii. **Temperature -** Vapour pressure increases with temperature of liquid. This is because, as temperature increases, kinetic energy of the molecules increases, hence, more molecules leave the surface of the liquid and come into vapour phase.

Raoult's Law

According to Raoult's law, for a solution of volatile liquids, the relative lowering of vapour pressure of solution is directly proportional to its mole fraction of dissolved solvent in solute.

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

Ideal Solution

The solution which obeys **Raoult's law** at all compositions of solute and solvent and at all temperature is called an **ideal solution. Ex-** Benzene and Toluene, n-hexane and n-heptane.

Characteristics of an ideal Solution

1. Raoult's law is obeyed by it.
2. $\Delta H_{mixing} = 0$ i.e., no heat should be absorbed or evolved during mixing.
3. $\Delta V_{mixing} = 0$, i.e., no change in volume (expansion or contraction) on mixing.

Non-ideal Solution

Those solutions which deviate from ideal behaviour are called **non-ideal solutions** or real solutions. Acetone and CS_2, Acetone and C_2H_5OH

Characteristics of an non-ideal Solution

1. Raoult's law is not obeyed by it.
2. $\Delta H_{mixing} \neq 0$ i.e., solution may absorb or release heat.

3. $\Delta V_{mixing} \neq 0$ i.e., solution may expand or contract on mixing of solute and solvent.

Azeotropic Mixture

At the constant boiling temperature, liquid mixture vapouries without change in composition and the condensate contains same composition, i.e., mixture distills like a pure liquid, which has same composition. At this point, solution or mixture is called an **azeotropic mixture.**

Colligative Properties

The properties of dilute solution which depends only on number of particles of solute (molecules or ions) present in the solution and not on their nature, are called colligative properties. The important **colligative properties** are;

1. Relative lowering of vapour pressure
2. Elevation of boiling point
3. Depression in freezing point

4. Osmotic pressure

i. Relative Lowering of Vapour Pressure

When a non-volatile solute is added to a solvent, its vapour pressure gets **lowered**. If this pressure is divided by pressure of pure solvent, this is called **relativelowering of vapour pressure.**

According to **Raoult's law,**

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

where, P° = V.P. of pure solvent

P° - P = lwering in vapour pressure

$$\frac{P^\circ - P}{P^\circ} = \frac{n_A}{n_A + n_B} \ldots \left[X_A = \frac{n_A}{n_A + n_B}\right]$$

For dilute solution $n_A + n_A \approx n_A$

$$\frac{P^\circ - P}{P^\circ} = \frac{n_A}{n_B}$$

$$\frac{P^\circ - P}{P^0} = \frac{W_A}{M_A} \times \frac{M_B}{W_B}$$

where, W_A = weight of solute

W_B = weight of solvent

M_A = molecular weight of solute

M_B = molecular weight of solvent

ii. Relative Elevation of boiling point

A liquid boils at the temperature at which its vapour pressure is equal to the atmospheric pressure. The boiling point of a solution of non-volatile solute is always higher than that of the boiling point of pure solvent in which the solution is prepared. Similar to lowering of vapour pressure, the elevation of boiling point also depends on the number of solute particles rather than their nature.

Let T° be the boiling point of pure solvent and T be the boiling point of solution. The increase in boiling point ΔTb = T – T° is known as **elevation in boiling point**.

For dilute solutions, the ΔT_b is directly proportional to the molal concentration of the solute in a solution. Thus

$$\Delta T_b \propto m$$

$$\Delta T_b = Kbm$$

K_b is molal **elevation constant** (Ebullioscopic constant). The unit of K_b is K kg mol^{-1}.

Substituting the value of molality in above equation, we get

$$\Delta T_b = \frac{K_b \times 1000 \times w_2}{M_2 \times w_1}$$

$$M_2 = \frac{K_b \times 1000 \times w_2}{\Delta T_b \times w_1}$$

Where, w_1 = mass of solvent, w_2 = mass of solute and M_2 = molar mass of solute

iii. Depression in freezing point

Freezing point is the temperature at which vapour pressure of liquid phase becomes same as that of solid phase. The decrease in freezing point of a solvent on the addition of a non-volatile solute is known as **depression in freezing point**.

Let T° be the freezing point of pure solvent and T be the freezing point of solution. The decrease in freezing point ΔT_f = T° -T is known as depression in freezing point.

For dilute solutions, the ΔT_f is directly proportional to the molal concentration of the solute in a solution. Thus

$$\Delta T_f \propto m$$

$$\Delta T_f = Kf \cdot m$$

Here K_f is molal depression constant or cryoscopic constant

Substituting the value of molality in above equation, we get

$$\Delta T_f = \frac{K \times w_2 \times 1000}{M_2 \times w_1}$$

$$M_2 = \frac{K \times w_2 \times 1000}{\Delta T_f \times w_1}$$

Where, w_1 = mass of solvent, w_2 = mass of solute and M_2 = molar mass of solute

iv. Osmotic pressure

Osmosis is the spontaneous flow of the solvent molecules from a less concentrated solution (dilute) to a more concentrated solution through a **semi-permeable membrane**. The driving force of osmosis is called **osmotic pressure**. Osmotic pressure may be defined as "the minimum excess pressure that has to be applied on the solution to prevent the osmosis".

Osmotic pressure of a solution ∝ molar concentration of solute in that solution

$$\pi \propto c$$

$$\pi = cRT$$

where, R = Gas constant = 0.0821 lit atm K^{-1} $mole^{-1}$

T = Temperature

c = Molar concentration

$$\pi = \frac{n}{V} RT \text{.....}[c = \frac{n}{V}]$$

$$\pi = \frac{W_B}{M_B} \frac{RT}{V}$$

W_B = wt. of solute

M_B = Molar mass of solute

van’t Hoff Factor

To calculate the extent of association or dissociation, van’t Hoff in 1886 introduced a factor ‘i’ called van’t Hoff factor. **van’t Hoff factor** ‘i’ is defined as ratio of the experimental value of colligative property to the calculated value of colligative property.

$$\text{i. e. , i}\,\frac{\text{Experiment ecolligatival properties}}{\text{Calculated ecolligativ properties}}$$