APPOLO STUDY CENTRE

(Acids, Bases and salts)

9th book	**Unit – 14**	**Acids, Bases And Salts**
10th book	**Unit – 9**	**Solutions**
	Unit – 10	**Types Of Chemical Reactions**

9th book
Unit – 14 – Acids, Bases and salts

Introduction

- We know that the physical world around us is made of large number of chemicals. Soil, air, water, all the life forms and the materials that they use are all consist of chemicals. Out of such chemicals, acids, bases and salts are mostly used in everyday life. Let it be a fruit juice or a detergent or a medicine, they play a key role in our day-to-day activities. Our body metabolism is carried out by means of hydrochloric acid secreted in our stomach. An acid is a the compound which is capable of forming hydrogen ions ($H+$) in aqueous solution whereas a base is a compoundthat forms hydroxyl ions ($OH–$) in solution. When an acid and a base react with each other, a neutral product is formed which is called salt. In this lesson let us discuss about them in detail.

Acids

- Look at the pictures of some of the materials used in our daily life, given below:

- All these edible items taste similar i.e. sour. What causes them to taste sour? A certain type of chemical compounds present in them gives sour taste. These are called acids. The word 'acid' is derived from the Latin name "acidus"

- which means sour taste. Substances with sour taste are called acids.

Source	Acid Present
Apple	Malic acid
Lemon	Citric acid
Grape	Tartaric acid
Tomato	Oxalic acid
Vinegar	Acetic acid

Curd	Lactic acid
Orange	Ascorbic acid
Tea	Tannic acid
Stomach juice	Hydrochloric acid
Ant, Bee	Formic acid

- In 1884, a Swedish chemist Svante Arrhenius proposed a theory on acids and bases. According to Arrhenius theory, an acid is a substance which furnishes H+ ions or H3O+ ions in aqueous solution. They contain one or more replaceable hydrogen atoms. For example, when hydrogen chloride is dissolved in water, it gives H1 and Cl2 ions in water.

$$HCl_{(aq)} \rightarrow H1_{(aq)} + Cl2_{(aq)}$$

- What happens to an acid or a base in water? Do acids produce ions only in aqueous solution? Hydrogen ions in HCl are produced in the presence of water. The separation of H+ ion from HCl molecules cannot occur in the absence of water.

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

- Hydrogen ions cannot exist alone, but they exist in combined state with water molecules. Thus, hydrogen ions must always be H+ (or) Hydronium (H3O+).

$$H^+ + H_2O \rightarrow H_3O^+$$

All acids essentially contain one or more hydrogens. But all the hydrogen containing substances are not acids. For example, methane (CH4) and ammonia (NH3) also contain hydrogen. But they do not produce H+ ions in aqueous solution.

- The following table enlists various acids and the ions formed by them in water.

Acid	Molecular Formula	Ions formed		No. of replaceable hydrogen
Acetic Acid	$CH-_3COOH$	H+	$CH-_3COO^-$	1
Formic Acid	HCOOH	H+	$HCOO^-$	1
Nitric Acid	HNO_3	H+	NO_3^-	1
Sulphuric Acid	H_2SO_4	2H+	SO_4^{2-}	2
Phosphoric Acid	H_3PO_4	3H+	PO_4^{3-}	3

Classification of Acids

- Acids are classified in different ways as given below:

(a) Based on their sources:

- **Organic Acids**: Acids present in plants and animals (living things) are organic acids. Example: HCOOH, CH_3COOH

- **Inorganic Acids**: Acids prepared from rocks and minerals are inorganic acids or mineral acids. Example: HCl, HNO_3, H2SO$_4$

(b) Based on their Basicity

- **Monobasic Acid**: Acid that contain only one replaceable hydrogen atom per molecule is called monobasic acid. It gives one hydrogen ion per molecule of the acid in solution. Example: HCl, HNO_3

For acids, we use the term basicity that refers to the number of replaceable hydrogen atoms present in one molecule of an acid. For example, acetic acid (CH3COOH) has four hydrogen atoms but only one can be replaced. Hence it is monobasic.

- **Dibasic Acid:** An acid which gives two hydrogen ions per molecule of the acid in solution. Example: H_2SO_4, H_2CO_3

- **Tribasic Acid**: An acid which gives three hydrogen ions per molecule of the acid in solution. Example: H3PO4

(c) Based on Ionisation

- Acids get ionised in water (produce H+ ions) completely or partially. Based on the extent of ionisation acids are classified as below.

- **Strong Acids**: These are acids that ionise completely in water. Example: HCl

- **Weak Acids**: These are acids that ionise partially in water. Example: CH_3COOH.

Ionisation is the condition of being dissociated into ions by heat or radiation or chemical reactions or electrical discharge.

(d) Based on Concentration

- **Concentrated Acid:** It has relatively large amount of acid dissolved in a solvent.

- **Dilute Acid:** It has relatively smaller amount of acid dissolved in solvent.

Properties of Acids

a) They have sour taste.
b) Their aqueous solutions conduct electricity since they contain ions.
c) Acids turns blue litmus red.
d) Acids react with active metals to give hydrogen gas.

$Mg + H_2SO_4 \rightarrow MgSO_4 + H_2 \uparrow$
$Zn + 2HCl \rightarrow ZnCl_2 + H_2 \uparrow$

Few metals do not react with acid and liberate hydrogen gas. For example: Ag, Cu.

e) Acids react with metal carbonate and metal hydrogen carbonate to give carbon dioxide.

$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2 \uparrow$
$NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2 \uparrow$

f) Acids react with metallic oxides to give salt and water.
CaO + H2SO4 → CaSO4 + H2O
g) Acids react with bases to give salt and water.
HCl + NaOH → NaCl + H2O

Uses of Acids

- Sulphuric acid is called King of Chemicals because it is used in the preparation of many other compounds. It is used in car batteries also.
- Hydrochloric acid is used as a cleansing agent in toilets.
- Citric acid is used in the preparation of effervescent salts and as a food preservative.
- Nitric acid is used in the manufacture of fertilizers, dyes, paints and drugs.
- Oxalic acid is used to clean iron and manganese deposits from quartz crystals. It is also used as bleach for wood and removing black stains.
- Carbonic acid is used in aerated drinks.
- Tartaric acid is a constituent of baking powder.

Role of water in acid solution
Acids show their properties only when dissolved in water. In water, they ionise to form H+ ions which determine the properties of acids. They do not ionise in organic solvents. For example, when HCl is dissolved in water it produces H+ ions and Cl_ ions whereas in organic solvents like ethanol they do not ionise and remain as molecule.

Aquaregia

- **We** know that metals like gold and silver are not reactive with either HCl or HNO3. But the mixture of these two acids can dissolve gold. This mixture is called Aquaregia. It is a mixture of hydrochloric acid and nitric acid prepared optimally in a molar ratio of 3:1. It is a yellow-orange fuming liquid. It is a highly corrosive liquid, able to attack gold and other substances.

Chemical formula : $3 HCL + HNO_3$

Solubility in water : Miscible in water
Melting point : - 42°C (- 44°F, 231K)
Boiling point : 108°C (226°F, 381K)

- The **term**aquaregia is a Latin phrase meaning 'King's Water'. The name reflects the ability of aquaregia to dissolve the noble metals such as gold, platinum and palladium.

Uses of Aquaregia

- **It** is used chiefly to dissolve metals such as gold and platinum.

It is used for cleaning and refining gold.

Bases

- According to Arrhenius theory, bases are substances that ionise in water to form hydroxyl ions (OH–). There are some metal oxides which give salt and water on reaction with acids. These are also called bases. Bases that are soluble in water are called alkalis. A base reacts with an acid to give salt and water only.

Base + Acid → Salt + Water

For example, zinc oxide (ZnO) reacts with HCl to give the salt zinc chloride and water.

$ZnO_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl2_{(aq)} + H2O_{(l)}$

Similarly, sodium hydroxide ionises in water to give hydroxyl ions and thus get dissolved in water. So it is an alkali.

$NaOH_{(aq)} \rightarrow Na^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$

- Bases contain one or more replaceable oxide or hydroxyl ions in solution. Table 14.3 enlists various bases and ions formed by them in water.

All alkalis are bases but not all bases are alkalis. For example: NaOH and KOH are alkalis whereas Al(OH)3 and Zn(OH)2 are bases.

Base	**Molecular Formula**	**Ions formed**		**No. of replaceable hydroxyl ion**
Calcium oxide	CaO	Ca^{2+}	O^{2-}	1
Sodium oxide	Na_2O	$2Na^{+}$	O^{2-}	1
Potassium hydroxide	KOH	K^{+}	OH^{-}	1
Calcium hydroxide	$Ca(OH)_2$	$Ca2^{+}$	OH^{-}	2

Aluminium hydroxide	$Al(OH)_3$	Al^{3+}	OH^-	3

Classification of Bases

(a) Based on their Acidity

- Monoacidic Base: It is a base that ionises in water to give one hydroxide ion per molecule. Example: NaOH, KOH
- Diacidic **Base**: It is a base that ionises in water to give two hydroxide ions per molecule. Example: $Ca(OH)_2$. $Mg(OH)_2$
- Triacidic **Base:** It is a base that ionises in water to give three hydroxide ions per molecule. Example: $Al(OH)_3$, $Fe(OH)_3$

(b) Based on concentration

- Concentrated **Alkali**: It is an alkali having a relatively high percentage of alkali in its aqueous solution.
- Dilute **Alkali**: It is an alkali having a relatively low percentage of alkali in its aqueous solution.

(c) Based on Ionisation

- Strong **Bases**: These are bases which ionise completely in aqueous solution.

Example: NaOH, KOH

- Weak **Bases**: These are bases that ionise partially in aqueous solution.

Example: NH_4OH, $Ca(OH)_2$

The term acidity is used for base, which means the number of replaceable hydroxyl groups present in one molecule of a base.

Properties of Bases

a) They have bitter taste.

b) Their aqueous solutions have soapy touch.

c) They turn red litmus blue.

d) Their aqueous solutions conduct electricity.

e) Bases react with metals to form salt with the liberation of hydrogen gas.

$$Zn + 2\,NaOH \rightarrow Na_2ZnO_2 + H_2\uparrow$$

- f) Bases react with non-metallic oxides to produce salt and water. Since this is similar to the reaction between a baseand an acid, we can conclude that non-metallic oxides are acidic in nature.

$$Ca(OH)_2 + CO_2 \rightarrow CaCO_3 + H_2O$$

g) Bases react with acids to form salt and water.

$$KOH + HCl \rightarrow KCl + H2O$$

The above reaction between a base and an acid is known as Neutralisation reaction.
h) On heating with ammonium salts, bases give ammonia gas.

$$NaOH + NH_4Cl \rightarrow NaCl + H_2O + NH_3\uparrow$$

Few metals do not react with sodium hydroxide. Example: Cu, Ag, Cr

- In the above activity you can observe that the bulb will start glowing only in the case of acids. But, you will observe that glucose and alcohol solution do not conduct electricity. Glowing of the bulb indicates that there is a flow of electric current through the solution. The electric current is carried through the solution by ions. Repeat the same activity using alkalis such as sodium hydroxide and calcium hydroxide.

Uses of Bases

(i) Sodium hydroxide is used in the manufacture of soap.
(ii) Calcium hydroxide is used in white washing of building.
(iii) Magnesium hydroxide is used as a medicine for stomach disorder.
(iv) Ammonium hydroxide is used to remove grease stains from cloths.

Tests for Acids and Bases

a) Test with a litmus paper:

- An acid turns blue litmus paper into red. A base turns red litmus paper into blue.

b) Test with an indicator Phenolphthalein:

- In acid medium, phenolphthalein is colourless. In basic medium, phenolphthalein is pink in colour.

c) Test with an indicator Methyl orange:

- In acid medium, methyl orange is pink in colour. In basic medium, methyl orange is yellow in colour.

Indicator	Colour in acid	Colour in acid
Litmus	Blue to Red	Red to Blue
Phenolphthalein	Colourless	Pink
Methyl orange	Pink	Yellow

Strenght of Acidic or Basic solutions

pH Scale

- A scale for measuring hydrogen ion concentration in a solution is called pH scale. The 'p' in pH stands for 'potenz' in German meaning power. pH scale is a set of numbers from 0 to 14 which is used to indicate whether a solution is acidic, basic or neutral.

  - ❖ Acids have pH less than 7
  - ❖ Bases have pH greater than 7
  - ❖ A neutral solution has pH equal to 7

Salts

- When you say salt, you may think of the common salt. Sea water contains many salts dissolved in it. Sodium chloride is separated from these salts. There are many other salts used in other fields. Salts are the products of the reaction between acids and bases. Salts produce positive ions and negative ions when dissolved in water.

Types of Salts

- Normal Salts: A normal salt is obtained by complete neutralization of an acid by a base.

$$\mathbf{NaOH + HCl \rightarrow NaCl + H2O}$$

- Acid Salts: It is derived from the partial replacement of hydrogen ions of an acid by a metal. When a calculated amount of a base is added to a polybasic acid, acid salt is obtained.

$$\mathbf{NaOH + H2SO_4 \rightarrow NaHSO_4 + H_2O}$$

- **Basic Salts**: Basic salts are formed by the partial replacement of hydroxide ions of a diacidic or triacidic base with an acid radical.

$Pb(OH)2 + HCl \rightarrow Pb(OH)Cl + H2O$

- Double **Salts**: Double salts are formed by the combination of the saturated solution of two simple salts in equimolar ratio followed by crystallization. For example, potash alum is a mixture of potassium sulphate and aluminium sulphate. $KAl(SO_4)_2 \cdot 12H2O$

Properties of Salts

- Salts are mostly solids which melt as well as boil at high temperature.
- Most of the salts are soluble in water. For example, chloride salts of potassium and sodium are soluble in water. But, silver chloride is insoluble in water
- They are odourless, mostly white, cubic crystals or crystalline powder with salty taste.
- Salt is hygroscopic in nature.

Water of Crystallisation

- Many salts are found as crystals with water molecules. These water molecules are known as water of crystallisation. Salts that contain water of crystallisation are called hydrated salts. The number of molecules of water hydrated to a salt is indicated after a dot in its chemical formula. For example, copper sulphate crystal have five molecules of water for each molecule of copper sulphate. It is written as $CuSO4.5H2O$ and named as copper sulphate pentahydrate. This water of crystallisation makes the copper sulphate blue. When it is heated, it loses its water molecules and becomes white.

- Salts that do not contain water of crystallisation are called anhydrous salt. They are generally found as powders. Fill in the blanks in the following table based on the concept of water of crystallisation.

Identification of Salts

(i) Physical examination of the salt.

- The physical examination of the unknown salt involves the study of colour, smell and density. This test is not much reliable.

(ii) Dry heating Test.

- This test is performed by heating a small amount of salt in a dry test tube. After all the water get evaporated, the dissolved salts are sedimented in the container.

(iii) Flame Test.

- Certain salts on reacting with concentrated hydrochloric acid (HCl) form their chlorides. The paste of the mixture with con. HCl is introduced into the flame with the help of platinum wire.

Colour of the flame	Inference
Brick red	Ca^{2+}
Golden Yellow	Na^{2+}
Pink Violet	K^{+}
Green Fleshes	Zn^{2+}

(iv) WhenHCl is added with a carbonate salt, it gives off CO2 gas with brisk effervescence.

Uses of Salts

- Common Salt (Sodium Chloride - NaCl)

It is used in our daily food and used as a preservative.

Washing Soda (Sodium Carbonate-Na_2CO_3)

i. It is used in softening hard water.
ii. It is used in glass, soap and paper industries.

Baking Soda (Sodium bicarbonate -$NaHCO_3$)

i. It is used in making of baking powder which is a mixture of baking soda and tartaric acid.
ii. It is used in soda-acid fire extinguishers.
iii. Baking powder is used to make cakes and bread, soft and spongy.
iv. It neutralizes excess acid in the stomach and provides relief.

Bleaching powder (Calcium Oxychloride - $CaOCl_2$)

i. It is used as disinfectant.
ii. It is used in textile industry for bleaching cotton and linen.

Plaster of Paris (Calcium Sulphate Hemihydrate - $CaSO_4 .½ H_2O$)

i. It is used for plastering bones.
ii. It is used for making casts for statues.

Points to Remember

- Acid is a substance which furnishes H^+ ions or H_3O^+ ions when dissolved in water.
- Base is a substance which releases OH^- ions when dissolved in water.
- Salt is the product of reaction between acids and bases.
- Acids and bases neutralize each other to form corresponding salts and water.
- Salts have various uses in everyday life.
- Acidic and basic solutions in water conduct electricity because they produce hydrogen and hydroxide ions respectively.
- When an acid reacts with a metal, hydrogen gas is evolved and a corresponding salt is formed.
- Phenolphthalein and Methyl orange are used as indicators to find out whether the given solution is acid or base.
- Litmus paper is also used to find out whether the given solution is acid or base.
- pH paper is find out the given solution whether acidic or basic in nature.
- Aquaregia is a mixture of hydrochloric acid and nitric acid optimally in a molar ratio of 3:1
- pH Scale is used to find out the power of hydrogen ion concentration in a solution.

10th STD
Unit – 9 SOLUTIONS

INTRODUCTION

- You have learnt about mixtures in your lower classes. Most of the substances that we encounter in our daily life are mixtures of two or more substances. The substances present in a mixture may exist in one or more physical state. For example, when we burn wood, the smoke released is a mixture of solid carbon and gases like CO2, CO, etc.

- In some cases of mixtures, their components can be separated easily whereas in some other cases they cannot be. Consider the two mixtures, one which contains salt and water, and the another which contains sand and water. Water is the one of the components in both the mixtures. In the first case salt disolves in water. In the second case the sand does not disolve in water. Sand in water can be separated by filtration but salt cannot be separated as it dissolves in water to form a homogeneous

mixture. This kind of homogenous mixtures are termed as "**solutions**".

SOLUTIONS IN DAY-TO-DAY LIFE

- One of the naturally existing solutions is sea water. We cannot imagine life on earth without sea water. It is a mixture of many dissolved salts. The another one is air. It is a mixture of gases like nitrogen, oxygen, carbon dioxide and other gases.

- All the life forms on the earth are associated with solutions. Plants take solutions of nutrients for their growth from the soil. Most of the liquids found in human body including blood, lymph and urine are solutions. Day to day human activities like washing, cooking, cleaning and few other activities involve the formation of solutions with water. Similarly, the drinks what we take, like fruit juice, aerated drinks, tea, coffee etc. are also solutions. Therefore, the ability of water to form solutions is responsible for sustenance of life. On the other hand, the same characteristic forms the basic cause of the addition of pollutants to water. However, the ability of water to form solutions influences the survival of man on the earth. In this lesson, let us learn the science of solutions.

COMPONENTS OF SOLUTIONS

- We know that, a **solution is a homogeneous mixture of two or more substances**. In a solution, the component which is present in lesser amount (by weight), is called **solute** and the component, which is present in a larger amount (by weight) is called **solvent**. The solute gets distributed uniformly throughout the solvent and thus forming the mixture homogeneous. So, the solvent acts as a dissolving medium in a solution. The process of uniform distribution of solute into solvent is called **dissolution**. Figure 9.2 shows the schematic representation of solution.

- A solution must at least be consisting of two components (a solute and a solvent). Such solutions which are made of one solute and one solvent (two components) are called **binary solutions**. e.g. On adding copper sulphate crystals to water, it dissolves in water forming a solution of copper sulphate as shown in Figure 9.3. It contains two components i.e. one solute- copper sulphate and one solvent- water. So it is a binary solution. Similarly, a solution may contain more than two components. For example if salt and sugar are added to water, both dissolve in water forming a solution. Here two solutes are dissolved in one solvent. Such kind of solutions which contain three components are called **ternary solutions**.

Types of Solutions

Based on the physical state of the solute and the solvent

- We know that substances normally exist in three physical states (phases) i.e., solid, liquid and gas. In binary solutions, both the solvent and solute may exist in any of these physical states. But the solvent constitutes the major part of the solution. Its physical state is the primary factor which determine the characteristics of the solution. Therefore, there are different types of binary solutions as listed.

Solute	**Solvent**	**Example**
Solid solution		
Solid	Solid	Copper dissolved in gold (Alloys)
Liquid	Solid	Mercury with sodium (amalgam)
Liquid solution		
Solid	Liquid	Sodium chloride dissolved in water
Liquid	Liquid	Ethyl alcohol dissolved in water
Gas	Liquid	carbon-di-oxide dissolved in water (Soda water)
Gaseous solution		
Liquid	Gas	Water vapour in air (cloud)
Gas	Gas	Mixture of Helium-Oxygen gases,

Based on the type of solvent

- Most of the substances are soluble in water. That is why, water is called as ‘Universal solvent”. However some substances do not dissolve in water. Therefore, other solvents such as ethers, benzene, alcohols etc., are used to prepare a solution. On the basis of type of solvent, solutions are classified into two types. They are aqueous solutions and non-aqueous solutions.

a) Aqueous solution:

- The solution in which water acts as a solvent is called aqueous solution. In general, ionic compounds are soluble in water and form aqueous solutions more readily than covalent compounds. E.g. Common salt in water, Sugar in water, Copper sulphate in water etc.

b) Non – Aqueous solution:

- The solution in which any liquid, other than water, acts as a solvent is called non-aqueous solution. Solvent other than water is referred to as non-aqueous solvent. Generally, alcohols, benzene, ethers, carbon disulphide, acetone, etc., are used as non-aqueous solvents. Examples for non-aqueous solutions: Sulphur dissolved in carbon disulphide, Iodine dissolved in carbon tetrachloride.

Based on the amount of solute

- The amount of the solute that can be dissolved in the given amount of solvent is limited under any given conditions. Based on the amount of solute, in the given amount of solvent, solutions are classified into the following types:

(i) Saturated solution
(ii) Unsaturated solution
(iii) Super saturated solution

(i) Saturated solution: A solution in which no more solute can be dissolved in a definite amount of the solvent at a given temperature is called saturated solution. e.g. 36 g of sodium chloride in 100 g of water at 25°C forms saturated solution.Further addition of sodium chloride, leave it undissolved.

(ii) Unsaturated solution: Unsaturated solution is one that contains less solute than that of the saturated solution at a given temperature. e.g. 10 g or 20 g or 30 g of Sodium chloride in 100 g of water at 25°C forms an unsaturated solution.

(iii) Super saturated solution: Supersaturated solution is one that contains more solute than the saturated solution at a given temperature. e.g. 40 g of sodium chloride in 100 g of water at 25°C forms super saturated solution. This state can be achieved by altering any other conditions liken temperature, pressure. Super saturated solutions are unstable, and the solute is reappearing as crystals when the solution is disturbed.

Concentrated and dilute solutions

- It is another kind of classification of unsaturated solutions. It expresses the relative concentration of two solutions with respect to their solutes present in the given amount of the solvent. For example, you are given two cups of tea. When you taste them, you feel that one is sweeter than the other. What do you infer from it? The tea which sweet more contains higher amount of sugar than the other. How can you express your observation? You can say that the tea is stronger. But a chemist would say that it is ‘concentrated’.

- When we compare two having same solute and solvent in a solutions, the one which contains higher amount of solute per the given amount of solvent is said to be ‘**concentratedsolution’ and the another is said to be ‘dilute solution’. They are schematically represented by Figure 9.5.**

- Differentiating solutions as dilute and concentrated is a qualitative representation. It does not imply the quantity of the solute. This difference is observed by means of some physical characteristics such as colour, density, etc.

Solubility

- Usually, there is a limit to the amount of solute that can be dissolved in a given amount of solvent at a given temperature. When this limit is reached, we have a saturated solution and any excess solute that is added, simply resides at the bottom of the solution. The extent of dissolution of a solute in a solvent can be better explained by its solubility. Solubility is measure of how much of a solute can be dissolved in a specified amount of a solvent.

$$\textbf{Solubility}=\frac{\textit{mass of the solve}}{\textit{mass of the solvent}}\times 100$$

Solubility’s of some common substances in water at 25°C

Name of the solute	**Formula of the solute**	**Solubility g/100 g water**
Calcium carbonate	$CaCO_3(s)$	0.0013
Sodium chloride	$NaCl\ (s)$	36
Ammonia	$NH_3(g)$	48
Sodium hydroxide	$NaOH(s)$	80
Glucose	$C_6H_{12}O_6(s)$	91
Sodium bromide	$NaBr(s)$	95
Sodium iodide	$NaI(s)$	184

Factors affecting solubility

- There are three main factors which govern the solubility of a solute. They are:

(i) Nature of the solute and solvent
(ii) Temperature
(iii) Pressure

(i) Nature of the solute and solvent

- The nature of the solute and solvent plays an important role in solubility. Although water dissolves an enormous variety of substances, both ionic and covalent, it does not dissolve everything. The phrase that scientists often use when predicting solubility is “like dissolves

- like.” This expression means that dissolving occurs when similarities exist between the solvent and the solute. For example: Common salt is a polar compound and dissolves readily in polar solvent like water.
- Non-polar compounds are soluble in non-polar solvents. For example, Fat dissolved in ether. But non-polar compounds, do not dissolve in polar solvents; polar compounds do not dissolve in non-polar solvents.

(ii) Effect of Temperature

Solubility of Solids in Liquid:

- Generally, solubility of a solid solute in a liquid solvent increases with increase in temperature. For example, a greater amount of sugar will dissolve in warm water than in cold water.
- In endothermic process, solubility increases with increase in temperature.
- In exothermic process, solubility decreases with increase in temperature.

Solubility of Gases in liquid

- Do you know why is it bubbling when water is boiled? Solubility of gases in liquid decrease with increase in temperature. Generally, water contains dissolved oxygen. When water is boiled, the solubility of oxygen in water decreases, so oxygen escapes in the form of bubbles.
- Aquatic animals live more in cold regions because, more amount of dissolved oxygen is present in the water of cold regions. This shows that the solubility of oxygen in water is more at low temperatures.

(iii) Effect of Pressure

- Effectof pressure is observed only in the case of solubility of a gas in a liquid. When the pressure is increased, the solubility of a gas in liquid increases.

- The common examples for solubility of gases in liquids are carbonated beverages, i.e. soft drinks, household cleaners containingaqueous solution of ammonia, formalin-aqueous solution of formaldehyde, etc.

Concentration of a Solution

- So far, we discussed what is a solution? what does it consist of and its types. Most of the chemical reactions take place in solutions form. So it is essential to quantify the solute in solvent to study the reactions. To quantify the solute in a solution, we can use the term “**concentration**”.

Concentration of a solution may be defined as the amount of solute present in a given amount of solution or solvent.

- Quantitatively, concentration of a solution may be expressed in different methods. But here, we shall discuss percentage by mass (% mass) and percentage by volume (% volume).

Mass percentage

- Mass percentage of a solution is defined as the percentage by mass of the solute present in the solution. It is mostly used when solute is solid and solvent is liquid.

Mass percentage =$\frac{\textit{mass of the solute}}{\textit{mass of the solution}} \times \mathbf{100}$

Mass percentage =$\frac{\textit{mass of the solute}}{\substack{\textit{mass of the solution+} \\ \textit{mass of the solvent}}} \times \mathbf{100}$

For example: 5% sugar solution (by mass) means 5 g of sugar in 95 g of water. Hence it is made 100g of solution.

- Usually, mass percentage is expressed as w/w (weight / weight); mass percentage is independent of temperature.

Volume percentage

- Volume percentage is defined as the percentage by volume of solute (in ml) present in the given volume of the solution.

Volume percentage =$\frac{\textit{volume of the solute}}{\textit{volume of the solution}} \times \mathbf{100}$

Volume percentage $=\frac{\boldsymbol{volume\ of\ the\ solute}}{\boldsymbol{volume\ of\ the\ solution+}\atop \boldsymbol{volume\ of\ the\ solvent}}\times \mathbf{100}$

- For example, 10% by volume of the solution of ethanol in water, means 10 ml of ethanol in 100 ml of solution (or 90 ml of water)

- Usually volume percentage is expressed as v/v (volume / volume). It is used when both the solute and solvent are liquids. Volume percentage decreases with increases in temperature, because of expansion of liquid.

- You can notice that in the commercial products that we come across in our daily life such as a solution of syrups, mouth wash, antiseptic solution, household disinfectants etc., the concentration of the ingredients is expressed as v/v. Similarly, in ointments, antacid, soaps, etc., the concentration of solutions are expressed as w/w.

Hydrated salts and Water of Crystallization

- When ionic substances are dissolved in water to make their saturated aqueous solution, their ions attract water molecules which then attached chemically in certain ratio. This process is called hydration. These ionic substances crystallize out from their saturated aqueous solution with a definite number of molecules of water. The number of water molecules found in the crystalline substance is called **water of crystallization**. Such salts are called hydrated salts.

- On heating these hydrated crystalline salts, they lose their water of crystallization and become amorphous or lose their colour (if they are coloured). Table 9.3 shows some common hydrated salts:

Common Name	IUPAC Name	Molecular Formula
Blue Vitriol	Copper (II) sulphate pentahydrate	$CuSO_4.5H_2O$
Epsom Salt	Magnesium sulphate heptahydrate	$MgSO_4.7H_2O$
Gypsum	Calcium sulphate dihydrate	$CaSO_4.2H_2O$
Green Vitriol	Iron (II) sulphate heptahydrate	$FeSO_4.7H_2O$
White Vitriol	Zinc sulphate heptahydrate	$ZnSO_4.7H_2O$

Copper sulphatepentahydrateCuSO4.5H2O (Blue vitriol)

- The number of water molecules in blue vitriol is five. So its water of crystallization is 5. When blue coloured copper sulphate crystals are gently heated, it loses its five watermolecules and becomes colourless anhydrous copper sulphate.

$$CuSO_4 \cdot 5H_2O \underset{\text{Cooling}}{\overset{\text{Heating}}{\rightleftharpoons}} CuSO_4 + 5H_2O$$

(Copper sulphate pentahydrate) (Anhydrous copper sulphate)

- If you add few drops of water or allow it to cool, the colourless anhydrous salt again turns back into blue coloured hydrated salt.

Magnesium sulphateheptahydrateMgSO4.7H2O (Epsom salt)

- Its water of crystallization is 7. When magnesium sulphateheptahydrate crystals are gently heated, it loses seven water molecules, and becomes anhydrous magnesium sulphate.

$$MgSO_4 \cdot 7H_2O \underset{\text{Cooling}}{\overset{\text{Heating}}{\rightleftharpoons}} MgSO_4 + 7H_2O$$

(Magnesium sulphate heptahydrate) (Anhydrous Magnesium sulphate)

- If you add few drops of water or allow it to cool, the colourless anhydrous salt again turns back into hydrated salt.

Hygroscopy

- Certain substances, when exposed to the atmospheric air at ordinary temperature, absorb moisture without changing their physical state. Such substances are called **hygroscopic substances** and this property is called hygroscopy.

Hygroscopic substances are used as drying agents.

Example:

1. Conc.Sulphuric acid (H_2SO_4).
2. Phosphorus Pentoxide (P_2O_5).
3. Quick lime (CaO).
4. Silica gel (SiO_2).
5. Anhydrous calcium chloride ($CaCl_2$).

Deliquescence

- Certain substances which are so hygroscopic, when exposed to the atmosphericair at ordinary temperatures, absorb enough water and get completely dissolved. Such substances are called **deliquescent substances** and this property is called **deliquescence**.
- Deliquescent substances lose their crystalline shape and ultimately dissolve in the absorbed water forming a saturated solution.

Deliquescence is maximum when:

1)The temperature is low
2)The atmosphere is humid

Examples: Calcium chloride ($CaCl_2$), Caustic soda (NaOH), Caustic potash (KOH) and Ferric chloride ($FeCl_3$).

Problems Based on Solubility and Percentage by Mass and Volume

I. Problems based on solubility

- 1) 1.5 g of solute is dissolved in 15 g of water to form a saturated solution at 298K. Find out the solubility of the solute at the temperature.

Hygroscopic substances	Deliquescence substances
When exposed to the atmosphere at ordinary temperature, they absorb moisture and do not dissolve.	When exposed to the atmospheric air at ordinary temperature, they absorb moisture and dissolve.
Hygroscopic substances do not change	Deliquescent substances change its

its physical state on exposure to air.	physical state on exposure to air.
Hygroscopic substances may be amorphous solids or liquids.	Deliquescent substances are crystalline solids.

Mass of the solute = 1.5 g
Mass of the solvent = 15 g

Solubility of the solute = $\frac{mass\ of\ the\ solute}{mass\ of\ the\ solvent}\times 100$

Solubility of the solute =$\frac{1.5}{15}\times 100$

= 10 g

- 2) Find the mass of potassium chloride would be needed to form a saturated solution in 60 g of water at 303 K? Given that solubility of the KCl is 37/100 g at this temperature.

Mass of potassium chloride in 100 g of water in saturated solution = 37 g

Mass of potassium chloride in =$\frac{37}{100}\times 60$ =22.2g
60 g of water in saturated solution

- 3) What is the mass of sodium chloride that would be needed to form a saturated solution in 50 g of water at 30°C. Solubility of sodium chloride is 36 g at 30°C?

At 30°C, 36 g of sodium chloride is dissolved in 100 g of water.

∴Mass of sodium chloride that would be need for 100 g of water = 36 g

∴ Mass of sodium chloride=$\frac{37\times 60}{100}$ =18g
dissolved in 50 g of water

- 4) The Solubility of sodium nitrate at 50°C and 30°C is 114 g and 96 g respectively. Find the amount of salt that will be thrown out when a saturated solution of sodium nitrate containing 50 g of water is cooled from 50°C to 30°C?

Amount of sodium nitrate dissolved in 100 g of water at 50°C is 114 g

∴ Amount of sodium nitratedissolving in 50 g of water at 50°C is

$= \frac{114\times50}{100}$ =57g

Similarly amount of sodium nitratedissolving in 50g of water at 30°C is

$= \frac{96\times50}{100}$ =48g

Amount of sodium nitrate thrown when 50g of water is cooled from 50°C to 30°C is

57-48 = 9 g

II. Problem based on mass percentage

- 1) A solution was prepared by dissolving 25 g of sugar in 100 g of water. Calculate the mass percentage of solute.

Mass of the solute = 25 g
Mass of the solvent = 100 g

Mass Percentage= $\frac{\textit{mass of the solute}}{\textit{mass of the solvent}}\times\mathbf{100}$

Mass Percentage= $\frac{\textit{mass of the solute}}{\textit{mass of the solute} + \textit{mass of the solvent}}\times\mathbf{100}$

$= \frac{25}{25+100}\times\mathbf{100}$

$= \frac{25}{125}\times\mathbf{100 = 20g}$

- 2) 16 grams of NaOH is dissolved in 100 grams of water at 25°C to form a saturated solution. Find the mass percentage of solute and solvent.

Mass of the solute (NaOH) = 16 g
Mass of the solvent H2O = 100 g

(i) Mass percentage of the solute

Mass percentage of solute $= \frac{\text{mass of the solute}}{\text{mass of the solute} + \text{mass of the solvent}} \times 100$

$$= \frac{16 \times 100}{16+100}$$

$$= \frac{1600}{116}$$

Mass percentage of the solute = 13.79 %

(ii) Mass percentage of solvent = 100 – (Mass percentage of the solute)

= 100 – 13.79
= 86.21%

- 3) Find the amount of urea which is to be dissolved in water to get 500 g of 10% w/w aqueous solution?

Mass Percentage(w/w) $= \frac{\text{mass of the solute}}{\text{mass of the solvent}} \times 100$

$$10 = \frac{\text{mass of the urea}}{500} \times 100$$

Mass of urea = 50g

(iii) **Problem based on Volume – volume percentage.**

- 1) A solution is made from 35 ml of Methanol and 65 ml of water. Calculate the volume percentage.

Volume of the ethanol = 35 ml
Volume of the water = 65 ml

Volume Percentage $= \frac{\text{volume of the solute}}{\text{volume of the solution}} \times 100$

Volume Percentage $= \frac{\text{volume of the solute}}{\text{volume of the solution} + \text{volume of the solvent}} \times 100$

Volume percentage = $\frac{35}{35+65} \times 100$

Volume percentage = $\frac{35}{100}\times\mathbf{100}$

$= \mathbf{35\%}$

- 2) Calculate the volume of ethanol in 200 ml solution of 20% v/v aqueous solution of ethanol.

Volume of aqueous solution = 200 ml
Volume percentage = 20%

VolumePercentage $= \frac{\textbf{volume of the solute}}{\textbf{volume of the solution}}\times\mathbf{100}$

$\mathbf{20} = \frac{\textbf{volume of the ethanol}}{\mathbf{200}}\times\mathbf{100}$

Volume of ethanol $=\frac{\mathbf{20\times200}}{\mathbf{100}} = \mathbf{40\ ml}$

10th book
Unit 10 - TYPES OF CHEMICAL REACTIONS

INTRODUCTION

- As you know from your earlier studies, a chemical reaction involves breaking of old chemical bonds and formation of new chemical bonds. This change may happen spontaneously or it may be facilitated by external forces or energy. Chemistry is all about chemical reactions. In your day to day life, you could observe many chemical reactions. A clear understanding of these reactions is essential in order to manipulate them for the sake of human life and environment. So, chemistry mainly focuses on chemical reactions. Let us try to find the answer for the following questions:

  - ❖ You need energy to play, walk, run or to perform various physical activities. Where do you get the energy from?
  - ❖ How do plants grow and get their food?
  - ❖ How does a car move using fuel?
  - ❖ Why does iron rust on its exposure to water or air?

- You get energy from the digestion of the food you eat. Plants grow by absorbing nutrients from the Earth and get their food by photosynthesis. The combustion of a fuel makes the car to move. Oxidation of iron causes rusting. So, all these processes are chemical changes i.e. the materials, which undergo changes are converted into some other new materials. For example, by burning petrol, the hydrocarbons present in it are converted into carbon dioxide and water. In this chapter, let us discuss the nature and types of chemical reactions.

What happens during a chemical reaction?

- In a chemical reaction, the atoms of the reacting molecules or elements are rearranged to form new molecules.
- Old chemical bonds between atoms are broken and new chemical bonds are formed.
- Bond breaking absorbs energy whereas bond formation releases energy

How are chemical reactions represented?

- When methane reacts with oxygen, it forms carbon dioxide and water. How can you represent this reaction? It can be written as a word equation as shown below:

Methane + Oxygen → Carbon dioxide + Water

- But, this equation does not give the chemical composition of the reactants and products. So, to learn the characteristics of a chemical reaction, it is represented by a chemical equation. In the chemical equation, the chemicals of the reaction are represented by their chemical formulas.The compounds or elements, which undergo reactions (reactants) are shown to the left of an arrow and the compounds formed (products) are shown to the right of the arrow. The arrow indicates the direction of the reaction. Thus, the aforesaid reaction can be written as follows:

$$CH_4 + O_2 \rightarrow CO_2 + H_2O$$

- But, this is also an incomplete chemical equation. Because, the law of conservation of matter states that matter cannot be created or destroyed. You cannot create new atoms by a chemical reaction. In contrast, they are rearranged in different ways by a chemical reaction to form a new compound. So, in a chemical equation, the number of atoms of the reactants and that of the products must be equal. The number of hydrogen and oxygen atoms in the reactants and the products are notequal in the given equation. On balancing the number of atoms, the following equation can be obtained:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

- Further, the chemical equation provides information on the physical state of the substances and the conditions under which the reaction takes place.

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$$

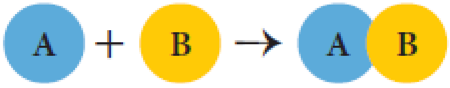
TYPES OF CHEMICAL REACTIONS

Classification based on the nature of rearrangements of atoms

- So far you studied about a chemical reaction and how it can be described as a chemical equation. A large number of chemical reactions are taking place around us every day.Are they taking place in a similar way? No. Each reaction involves different kinds of atoms and hence the way they react also differs. Thus, based on the manner by which the atoms of the reactants are rearranged, chemical reactions are classified as follows.

Combination reactions

- A combination reaction is a reaction in which two or more reactants combine to form a compound. It is otherwise called 'synthesis reaction' or 'composition reaction'. When a reactant ‘A’ combines with ‘B’, it forms the product ‘AB’. The generalised scheme of a combination reaction is given below:



Example:

- Hydrogen gas combines with chlorine gas to form hydrogen chloride gas.

$$H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$$

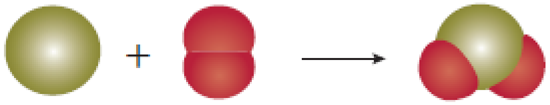
- Depending on the chemical nature of the reactants, there are threeclassesof combination reactions:

Element + Element → Compound

- In this type of combination reaction, two elements react with one other to form a compound. The reaction may take place between a metal and a non-metal or two non-metals.

Example 1:

- When solid sulphur reacts with oxygen, it produces sulphur dioxide. Here both the reactants are non-metals.



$$S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$$

Example 2:

- Sodium, a silvery-white metal, combines with chlorine, a pale yellow green gas, to form sodium chloride, an edible compound. Here one of the reactants is a metal (sodium) and the other (chlorine) is a non-metal.

$$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$$

Test Yourself:

- Identify the possible combination reactions between the metals and non-metals given in the following table and write their balanced chemical equations:

Metals	Non-metals

Na, K, Cs, Ca, Mg F, Cl, Br, I

Compound + Element → Compound

- In this case, a compound reacts with an element to form a new compound. For instance, phosphorous trichloride reacts with chlorine gas and forms phosphorous pentachloride.

$$PCl_{3(l)} + Cl_{2(g)} \rightarrow PCl_{5(s)}$$

Compound + Compound → Compound

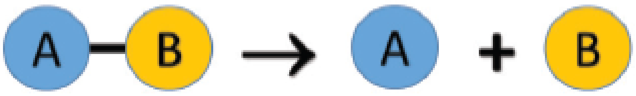
- It is a reaction between two compounds to form a new compound. In the following reaction, silicon dioxide reacts with calcium oxide to form calcium silicate.

$$SiO_{2(s)} + CaO_{(s)} \rightarrow CaSiO_{3(s)}$$

- Most of the combination reactions are exothermic in nature. Because, they involve the formation of new bonds, which releases a huge amount of energy in the form of heat.

Decomposition reactions

- In a decomposition reaction, a single compound splits into two or more simpler substances under suitable conditions. It is the opposite of the combination reaction. The generalised schemeof a decomposition reaction is given below:



- Breaking of bonds is the major phenomenon in a decomposition reaction and hence it requires energy to break the bonds, depending on the nature of the energy used in the decomposition reaction.

There are three main classes of decomposition reactions. They are

- Thermal Decomposition Reactions
- Electrolytic Decomposition Reactions
- Photo Decomposition Reactions

A solution of slaked lime is used for white washing walls. Calcium hydroxide reacts slowly with the carbon dioxide in air to form a thin layer of calcium carbonate on the walls. Calcium carbonate is formed aft er two to three days of white washing and gives a shiny finish to the walls. It is interesting to note that the chemical formula for marble is also$CaCO_3$

$$Ca(OH)_{2(aq)} + CO_{2(g)} \rightarrow CaCO_{3(s)} + H_2O_{(I)}$$

Thermal Decomposition Reactions

- In this type of reaction, the reactant is decomposed by applying heat. For example, on heating mercury (II) oxide is decomposed into mercury metal and oxygen gas. As the molecule is dissociated by the absorption of heat, it is otherwise called 'Thermolysis'.It is a class of compound to element/element decomposition. i.e. a compound (HgO) is decomposed into two elements (Hg and Oxygen).

$$2HgO_{(s)} \xrightarrow{Heat} 2Hg_{(l)} + O_{2(g)}$$

- Similarly, when calcium carbonate is heated, it breaks down in to calcium oxide and carbon dioxide. It is a type of compound to compound/compound decomposition.

$$CaCO_{3(s)} \xrightarrow{Heat} CaO_{(s)} + CO_{2(g)}$$

- In thermal decomposition reaction, heat is supplied to break the bonds. Such reactions, in which heat is absorbed, are called 'Endothermic reactions'.

Electrolytic Decomposition Reactions

- In some of the decomposition reactions, electrical energy is used to bring about the reaction. For example, decomposition of sodium chloride occurs on passing electric current through its aqueous solution. Sodium chloride decomposes in to metallic sodium and chlorine gas. This process is termed as 'Electrolysis'.

$$2NaCl_{(aq)} \xrightarrow{Electricity} 2Na_{(s)} + Cl_{2(g)}$$

- Here, a compound (NaCl) is converted into elements (Na and chlorine). So it is a type of compound to element/element decomposition.

Photo Decomposition Reactions

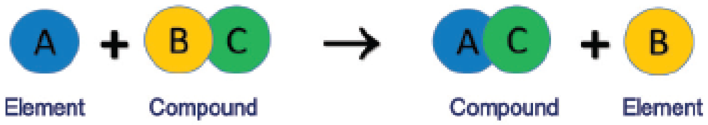
- Light is an another form of energy, which facilitates some of the decomposition reactions. For example, when silver bromide is exposed to light, it breaks down into silver metal and bromine gas. As the decomposition is caused by light, this kind of reaction is also called 'Photolysis'.

$$2AgBr_{(s)} \xrightarrow{Ligh} 2Ag_{(s)} + Br_{2(g)}$$

- The yellow coloured silver bromide turns into grey coloured silver metal. It is also a compound to element/element decomposition.

Single Displacement Reactions

- It is a reaction between an element and a compound. When they react, one of the elements of the compound-reactant is replaced by the element-reactant to form a new compound and an element. The general schematic representation of a single displacement reaction is given as:



- 'A' displaces element 'B' from the compound 'BC' and hence a single displacement reaction occurs. If zinc metal is placed in hydrochloric acid, hydrogen gas is evolved. Here, hydrogen is displaced by zinc metal and zinc chloride is formed.

$$Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

$$Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$$

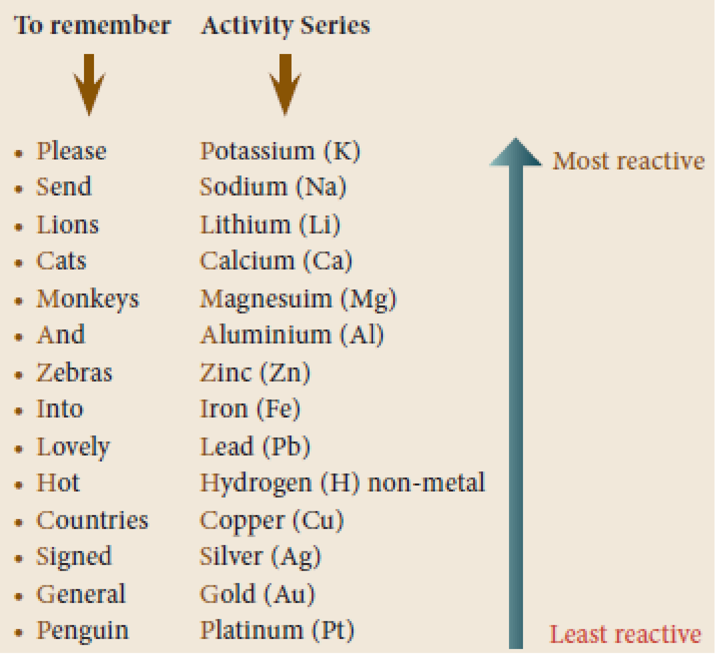
- If an iron nail is placed in an aqueous solution of copper (II) sulphate as shown in Fig. 10.2, the iron displaces copper from its aqueous solution and the so formed copper deposits over the iron nail.

- It is easy to propose so many reactions of this kind with different combinations ofreactants. Will they all occur in practice? No. This is most easily demonstrated with halogens. Let us consider the following two reactions:

$$2NaCl_{(aq)} + F_{2(g)} \rightarrow 2NaF_{(aq)} + Cl_{2(g)}$$
$$2NaF_{(aq)} + Cl_{2(g)} \rightarrow 2NaCl_{(aq)} + F_{2(g)}$$

- The first reaction involves the displacement of chlorine from NaCl, by fluorine. In the second reaction, chlorine displaces fluorine from NaF. Out of these two, the second reaction will not occur. Because, fluorine is more active than chlorine and occupies the upper position in the periodic table. So, in displacement reactions, the activity of the elements and their relative position in the periodic table are the key factors to determine the feasibility of the reactions. More active elements readily displace less active elements from their aqueous solution.

The activity series of some elements is given below:



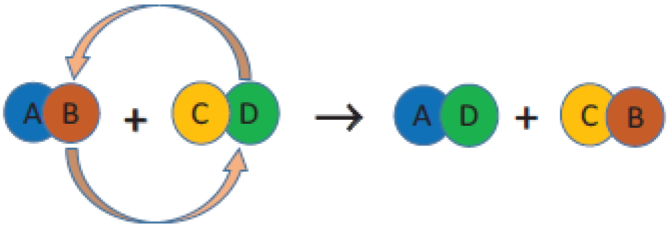
To remember	Activity Series
Please	Potassium (K)
Send	Sodium (Na)
Lions	Lithium (Li)
Cats	Calcium (Ca)
Monkeys	Magnesuim (Mg)
And	Aluminium (Al)
Zebras	Zinc (Zn)
Into	Iron (Fe)
Lovely	Lead (Pb)
Hot	Hydrogen (H) non-metal
Countries	Copper (Cu)
Signed	Silver (Ag)
General	Gold (Au)
Penguin	Platinum (Pt)

By referring the activity series, try to answer the following questions:

- Which of the metals displaces hydrogen gas from hydrochloric acid? Silver or Zinc. Give the chemical equation of the reaction and Justify your answer

Double Displacement Reactions

- When two compounds react, if their ions are interchanged, then the reaction is called double displacement reaction. The ion of one compound is replaced by the ion of the another compound. Ions of identical charges are only interchanged, i.e., a cation can be replaced by other cations. This reaction is also called 'Metathesis Reaction'. The schematic representation of a double displacement reaction is given below:



- For a double displacement reaction to take place, one of the products must be a precipitate or water. By this way, there are major classes of double displacement reactions. They are:

1. Precipitation Reactions
2. Neutralization Reactions

Precipitation Reactions

- When aqueous solutions of two compounds are mixed, if they react to form an insoluble compound and a soluble compound, then it is called precipitation reaction. Because the insoluble compound, formed as one of the products, is a precipitate and hence the reaction is so called.

Differences between combination and decomposition reactions

COMBINATION REACTIONS	DECOMPOSITION REACTIONS
One or more reactants combine to form a single product	A single reactant is decomposed to form one or more products
Energy is released	Energy is absorbed
Elements or compounds may be the reactants	Single compound is the reactant

- When the clear aqueous solutions of potassium iodide and lead (II) nitrate are mixed, a double displacement reaction takes place between them.

$$Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \rightarrow PbI_{2(s)}\downarrow + 2KNO_{3(aq)}$$

Potassium and lead displace or replace one other and form a yellow precipitate of lead.

Neutralization Reactions

- In your lower classes, you have learned the reaction between an acid and a base. It isanother type of displacement reaction in which the acid reacts with the base to form a salt and water. It is called 'neutralization reaction' as both acid and base neutralize each other.

Acid + Base → Salt + Water

- Reaction of sodium hydroxide with hydrochloric acid is a typical neutralization reaction. Here, sodium replaces hydrogen from hydrochloric acid forming sodium chloride, a neutral soluble salt.

$$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

- Similarly, when ammonium hydroxide reacts with nitric acid, it forms ammonium nitrate and water.

$$HNO_{3(aq)} + NH_4OH_{(aq)} \rightarrow NH_4NO_{3(aq)} + H_2O_{(l)}$$

Combustion Reactions

- A combustion reaction is one in which the reactant rapidly combines with oxygen to form one or more oxides and energy (heat). So in combustion reactions, one of the reactants must be oxygen. Combustion reactions are majorly used as heat energy sources in many of our day to day activities. For instance, we use LPG gas for domestic cooking purposes. We get heat and flame from LPG gas by its combustion reaction of its constituent gases. LPG is a mixture of hydrocarbon gases like propane, butane, propylene, etc. All these hydrocarbons burn with oxygen to form carbon dioxide and water.

$$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)} + \text{Heat}$$

- Since heat is evolved, it is an exothermic reaction. As oxygen is added, it is also an oxidation. So, combustion may be called as an exothermic oxidation. If a flame is formed (as shown in Fig. 10.4), then it is called burning.

  1. Digestion of Food
  2. Rusting of iron

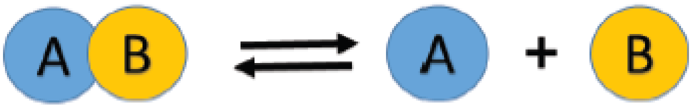
- Many thousands of reactions fall under these five categories and further you will learn in detail about these reactions in your higher classes.

Classification based on the direction of the reaction

- You know that innumerable changes occur every day around us. Are all they permanent? For example, liquid water freezes into ice, but then ice melts into liquid water. In other words, freezing is reversed. So, it is not a permanent change. Moreover, it is a physical change. Physical changes can be reversed easily. Can chemical changes be reversed? Can the products be converted into reactants? Let us consider the burning of a wood. The carbon compounds present in the wood are burnt into carbon dioxide gas and water. Can we get back the wood immediately from carbon dioxide and water? We cannot. So, it is a permanent change. In most of the cases, we cannot. But, some chemical reactions can be reversed. Our mobile phone gets energy from its lithium ion battery by chemical reactions. It is called discharging. On recharging the mobile, these chemical reactions are reversed. Thus, chemical reactions may be reversed under suitable conditions. Hence, they aregrouped into two categories such as reversible and irreversible reactions.

Reversible Reactions

- A reversible reaction is a reaction that can be reversed, i.e., the products can be converted back to the reactants. A reversible reaction is represented by a double arrow with their heads in the direction opposite to each other. Thus, a reversible reaction can be represented by the following equation:



Explanation:

- Here, the compound 'AB' undergoes decomposition to form the products 'A' and 'B'. It is the forward reaction. As soon as the products are formed, they combine together to form 'AB'. It is the backward reaction. So, the reaction takes place in both the directions. Do you think then that no products are formed in the aforesaid reaction? If you think so, you are wrong. Because, even though the reaction takes place in both the directions, at the initial stage the rates (speed) of these reactions are not equal. Consider the following decomposition reaction of phosphorous pentachloride into phosphorous trichloride and chlorine.

$$PCl_{5(g)} \rightleftarrows PCl_{3(g)} + Cl_{2(g)}$$

- The forward reaction is the decomposition of PCl_5 and the backward reaction is the combination of PCl_3 and Cl_2. Initially, the forward reaction proceeds faster than the backward reaction. After sometimes, the speed of both the reactions become equal. So, PCl_5 cannot be completely converted into the products as the reaction is reversed. It is a reversible reaction. The actual measurements of the given reaction show that the reaction is at equilibrium, but the amount of PCl_5 is more than that of PCl_3 and Cl_2.

- Thus, more amount of products can be obtained in a reversible reaction by the periodical removal of one of the products or the periodical addition of the reactants.

$$2H2O_{2(aq)} \rightleftarrows 2H_2O_{(l)} + O_{2(g)}$$

Irreversible Reactions

- The reaction that cannot be reversed is called irreversible reaction. The irreversible reactions are unidirectional, i.e., they take place only in the forward direction. Consider the combustion of coal into carbon dioxide and water.

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)} + \text{Heat}$$

- In this reaction, solid coal burns with oxygen and gets converted into carbon dioxide gas and water. As the product is a gas, as soon as it is formed it escapes out of the reaction container. It is extremely hard to decompose a gas into a solid. Thus, the backward reaction is not possible in this case. So, it is an irreversible reaction. Table 10.2 provides themain differences between a reversible and an irreversible reaction:

Differences between reversible and irreversible reactions

REVERSIBLE REACTION	IRREVERSIBLE REACTION
It can be reversed under suitable conditions.	It cannot be reversed.
Both forward and backward reactions take place simultaneously.	It is unidirectional. It proceeds only in forward direction.
It attains equilibrium	Equilibrium is not attained.
The reactants cannot be converted completely into products.	The reactants can be completely converted into products.
It is relatively slow.	It is fast.

RATE OF A CHEMICAL REACTION

- So far we discussed various types of chemical reactions and the nature of the reactants and products. Let us consider the following reactions:

  1. Rusting of iron
  2. Digestion of food
  3. Burning of petrol
  4. Weathering of rock

- How fast is each reaction? Rank them from the slowest to fastest. How will you determine, which is the fastest and which is the slowest? One of the ways to find out how fast a reaction is as follows: Measure the amount of reactants or products before and after a specific period of time. For example, let us assume that 100 gof a substance ‘A’ undergoes a reaction and after an hour 50 g of ‘A’ is left.

$$A \rightarrow \text{Product}$$

- In another instance, 100 g of substance 'C' undergoes a reaction and after an hour, 20 g of 'C' is left.

C → Product

- Can you say which is the faster reaction? In the first reaction, 50 g of the reactant (A) is converted into products whereas in the second reaction 80 g of the reactant is converted into products in one hour. So, the second reaction is faster. This measurement is called 'the reaction rate'.

- "Rate of a reaction is the change in the amount or concentration of any one of the reactants or products per unit time".

Consider the following reaction

A → B

The rate of this reaction is given by

$$\text{Rate} = -\frac{d[A]}{dt} = +\frac{d[B]}{dt}$$

Where,

[A] – Concentration of A
[B] – Concentration of B

- The negative sign indicates the decrease in the concentration of A with time.

- The postive sign indicates the increase in the concentration of B with time.

Note: '[ ]' represents the concentration, 'd' represents the infinitesimal change in the concentration.

Why is reaction rate important?

- Faster the reaction, more will be the amount of the product in a specified time. So, the rate of a reaction is important for a chemist for designing a process to get a good yield of a product. Rate of reaction is also important for a food processor who hopes to slow down the reactions that cause food to spoil.

Factors influencing the rate of a reaction

- Can the rate of a reaction be changed? The rate of a reaction can be changed. Forexample, iron gets rusted faster in an acid than in water. Important factors that affect rate of a reaction are

1. Nature of the reactants
2. Concentration of the reactants
3. Temperature
4. Catalyst
5. Pressure
6. Surface area of the reactants

Nature of the reactants

- The reaction of sodium with hydrochloric acid is faster than that with acetic acid. Do you know why? Hydrochloric acid is a stronger acid than acetic acid and thus more reactive. So, the nature of the reactants influence the reaction rate.

$$2Na_{(s)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + H_{2\,(g)} \text{ (fast)}$$

$$2Na_{(s)} + 2CH_3COOH_{(aq)} \rightarrow 2CH_3COONa_{(aq)} + H_{2(g)}\text{(slow)}$$

Concentration of the reactants

- Changing the amount of the reactants also increases the reaction rate. The amount of the substance present in a certain volume of the solution is called ‘concentration’. More the concentration, more particles per volume exist in it and hence faster the reaction. Granulated zinc reacts faster with 2M hydrochloric acid than 1M hydrochloric acid.

Temperature

- Most of the reactions go faster at higher temperature. Because adding heat to the reactants provides energy to break more bonds and thus speed up the reaction. Calcium carbonate reacts slowly with hydrochloric acid at room temperature. When the reaction mixture is heated the reaction rate increases.

Food kept at room temperature spoils faster than that kept in the refrigerator. In the refrigerator, the temperature is lower than the room temperature and hence the reaction rate is less.

Pressure

- If the reactants are gases, increasing their pressure increases the reaction rate. This is because, on increasing the pressure the reacting particles come closer and collide frequently.

Catalyst

- A catalyst is a substance which increases the reaction rate without being consumed in the reaction. In certain reactions, adding a substance as catalyst speeds up the reaction. For example, on heating potassium chlorate, it decomposes into potassium chloride and oxygen gas, but at a slower rate. If manganese dioxide is added, it increases the reaction rate.

Surface area of the reactants

- When solid reactants are involve in a reaction, their powdered form reacts more readily. For example, powdered calcium carbonate reacts more readily with hydrochloric acid than marble chips. Because, powdering of the reactants increases the surface area and more energy is available on collision of the reactant particles. Thus, the reaction rate is increased.

STATE OF EQUILIBRIUM

- In a reversible reaction, both forward and backward reactions take place simultaneously. When the rate of the forward reaction becomes equal to the rate of backward reaction, then no more product is formed. This stage of the reaction is called ‘equilibrium state’. After this stage, no net change in the reaction can occur and hence in the amount of the reactants and products. Since this equilibrium is attained in a chemical reaction, it is called ‘Chemical Equilibrium’. Chemical Equilibrium: It is state of a reversible chemical reaction in which no change in the amount of the reactants and products takes place. At equilibrium,

Rate of forward reaction = Rate of backward reaction

Explanation:

- Initially the rate of the forward reaction is greater than the rate of the backward reaction. However, during the course of reaction, the concentration of the reactants decreases and the concentration of the products increases. Since the rate of a reaction is directly proportional to the concentration, the rate of the forward reaction decreases with time, whereas the rate of the backward reaction increases.

- At a certain stage, both the rates become equal. From this point onwards, there will be no change in the concentrations of both the reactants and the products with time. This state is called as equilibrium state.

- Let us consider the decomposition of calcium carbonate into lime and carbon dioxide. It is a reversible reaction. The speed of each reaction can be determined by how quickly the reactant disappears. If the reaction is carried out in a closed vessel, it reaches a chemical equilibrium. At this stage,

$$CaCO_{3(s)} \rightleftarrows CaO_{(s)} + CO_{2(g)}$$

- The rate of decomposition of $CaCO_3$= The rate of combination of CaO and CO_2.

- Not only chemical changes, physical changes also may attain equilibrium. When water kept in a closed vessel evaporates, it forms water vapour. No water vapour escapes out of the container as the process takes place in a closed vessel. So, it builds up the vapour pressure in the container. At one time, the water vapour condenses back into liquid water and when the rate of this condensation becomes equal to that of vapourisation, the process attains equilibrium.

- At this stage, the volume of the liquid and gaseous phases remain constant. Since it is a physical change, the equilibrium attained is called 'Physical Equilibrium'. Physical equilibrium is a state of a physical change at which the volume of all the phases remain unchanged.

$$H2O_{(l)} \underset{\text{Condensation}}{\overset{\text{Evaporation}}{\rightleftarrows}} H_2O\ C_{\ (g)}$$

Characteristics of equilibrium .

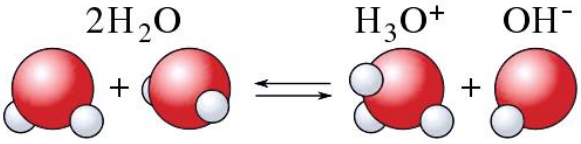
- In a chemical equilibrium, the rates of the forward and backward reactions are equal.
- The observable properties such as pressure, concentration, colour, density, viscosity etc., of the system remain unchanged with time.
- The chemical equilibrium is a dynamic equilibrium, because both the forward and backward reactions continue to occur even though it appears static externally.
- In physical equilibrium, the volume of all the phases remain constant.

Aerated soft drinks contain dissolved carbon dioxide in a pop bottle (Soda). When the bottle is sealed, the dissolved carbon dioxide (in the form of carbonic acid) and gaseous CO2 are in equilibrium with each other. When you open the bottle, the gaseous CO2 can escape. So, the dissolved CO2 begins to undissolve back to the gas phase trying to replace the gas that was lost, when you opened the bottle. That's why if you leave it open long enough, it will goes 'flat'. All the CO2 will be gone, blown away in the

IONIC PRODUCT OF WATER

- Although pure water is often considered as a non-conductor of electricity, precise measurements show that it conducts electricity to a little extent. This conductivity of water has resulted from the self-ionisation of water. Self-ionisation or auto ionisation is a reaction in which two like molecules react to give ions. In the process of ionisation of water, a proton from one water molecule is transferred to another water molecule leaving behind an OH— ion. The proton gets dissolved in water forming the hydronium ion as shown in the following equation:

$$H2O_{(l)} + H_2O_{(l)} \rightleftarrows H_3O^+_{(aq)} + OH^-_{(aq)}$$



- The hydronium ion formed is a strong acid and the hydroxyl ion is a strong base. So as fast as they are formed, they react again to produce water. Thus, it is a reversible reaction and attains equilibrium very quickly. So, the extent of ionisation is very little and the concentration of the ions produced is also very less. The product of the concentration of the hydronium ion and the hydroxyl ion is called 'ionic product of water'. It is denoted as 'Kw'. It is mathematically expressed as follows:

$$K_w = [H_3O^+] [OH^-]$$

$[H3O^+]$ may be simply written as $[H^+]$. Thus the ionic product of water may also be expressed as

$$K_w = [H^+] [OH^-]$$

Its unit is $mol^2\ dm^{-6}$. At 25° C, its value is 1.00×10^{-14}.

pHSCALE

- All the aqueous solutions may contain hydrogen and hydroxyl ions due to self-ionisation of water. In addition to this ionisation, substances dissolved in water also may produce hydrogen ions or hydroxyl ions. The concentration of these ions decides whether the solution is acidic or basic. pH scale is a scale for measuring the hydrogen ion concentration in a solution. The 'p' in pH stands for 'Potenz' in German meaning 'power'. pH notation was devised by the Danish biochemist Sorensen in 1909. pH scale is a set of numbers from 0 to 14 which is used to indicate whether a solution is acidic, basic or neutral.

  - Acids have pH less than 7
  - Bases have pH greater than 7
  - A neutral solution has pH equal to 7

The pH is the negative logarithm of the hydrogen ion concentration

$$pH = -\log_{10}[H^+]$$

Common Acids	**p^H**	**Common bases**	**p^H**
HCL(4%)	0	Blood plasma	7.4
Stomach acid	1	Egg white	8
Lemon juice	2	Sea water	9
Vinegar	3	Baking soda	10
Oranges	3.5	Antacids	10
Soda, Graps	4	Ammonia water	11
Sour milk	4.5	Lime water	12
Fresh Milk	5	Drain cleaner	13
Human saliva	6-8	Caustic soda 4% (NaOH)	14
Pure water	7	Milk of magnesia	10
Tomato juice	4.2	Coffee	5.6

How can we measure the pH of a given solution using pH Paper

- The pH of a solution can be determined by using a universal indicator. It contains a mixture of dyes. It comes in the form of a solution or a pH paper.

- A more common method of measuring pH in a school laboratory is by using the pH paper. A pH paper contains a mixture of indicators. It shows a specific colour at a given pH. A colour guide is provided with the bottle of the indicator or the strips of paper impregnated with it, which are called pH paper strips. The test solution is tested with a drop of the universal indicator, or a drop of the test solution is put on the pH paper. The colour of the solution on the pH paper is compared with the colour chart and the pH value is read from it. The pH values thus obtained are only approximate values.

ROLE OF pH IN EVERYDAY LIFE
Are plants and animals pH sensitive?

- Our body works within the pH range of 7.0 to 7.8. Living organisms can survive only in a narrow range of pH change. Different body fluids have different pH values. For example, pH of blood is ranging from 7.35 to 7.45. Any increase or decrease in this value leads to diseases. The ideal pH for blood is 7.4.

pH in our digestive system

- It is very interesting to note that our stomach produces hydrochloric acid. It helps in the digestion of food without harming the stomach. During indigestion the stomach produces too much acid and this causes pain and irritation. pH of the stomach fluid is approximately 2.0.

pH changes as the cause of tooth decay

- pH of the saliva normally ranges between 6.5 to 7.5. White enamel coating of our teeth is calcium phosphate, the hardest substance in our body. When the pH of the mouth saliva falls below 5.5, the enamel gets weathered. Toothpastes, which are generally basic are used for cleaning the teeth that can neutralise the excess acid and prevent tooth decay.

pH of soil

- In agriculture, the pH of the soil is very important. Citrus fruits require slightly alkaline soil, while rice requires acidic soil and sugarcane requires neutral soil.

pH of rain water

- The pH of rain water is approximately 7, which means that it is neutral and also represents its high purity. If the atmospheric air is polluted with oxide gases of sulphur and nitrogen, they get dissolved in the rain water and make its pH less than 7. Thus, if the pH of rain water is less than 7, then it is called acid rain. When acid rain flows into the rivers it lowers the pH of the river water also.

pH CALCULATION

The pH is the negative logarithm of the hydrogen ion concentration

$$pH = -\log_{10} [H^+]$$

Example:

Calculate the pH of 0.01 M HNO_3?

Solution:

$[H^+] = 0.01$
$pH = -\log_{10} [H^+]$
$pH = -\log_{10} [0.01]$
$pH = -\log_{10} [1 \times 10^{-2}]$
$pH = -(\log_{10}1 - 2 \log_{10}10)$
$pH = 0 + 2 \times \log_{10}10$
$pH = 0 + 2 \times 1 = 2$
$pH = 2$

---

pOH:

- The pOH of an aqueous solution is realted to the pH.
  The pOH is the negative logarithm of the hydroxyl ion concentration

$$pOH = -\log_{10}[OH^-]$$

Example:

- The hydroxyl ion concentration of a solution is 1×10^{-9}M. What is the pOH of the solution?

Solution

$pOH = -\log_{10} [OH^-]$
$pOH = -\log_{10} [1 \times 10^{-9}]$
$pOH = -(\log_{10} 1.0 + \log10\ 10^{-9})$
$pOH = -(0-9 \log_{10}10)$
$pOH = -(0 - 9)$
$pOH = 9$

Relationship between pH and pOH

- The pH and pOH of a water solution at 25oC are related by the following equation.

pH + pOH = 14

- If either the pH or the pOH of a solution is known, the other value can be calculated.

Example: A solution has a pOH of 11.76. What is the pH of this solution?

$$pH = 14 - pOH$$
$$pH = 14 - 11.76 = 2.24$$

PROBLEMS

Example 1:

Calculate the pH of 0.001 molar solution of HCl.

- **Solution:**

HCl is a strong acid and is completely dissociated in its solutions according to the process:

$$HCl_{(aq)} \rightarrow H^{+}_{(aq)} + Cl^{-}_{(aq)}$$

From this process it is clear that one mole of HCl would give one mole of H^+ ions. Therefore, the concentration of H+ ions would be equal to that of HCl, i.e., 0.001 molar or 1.0×10^{-3}mol litre^{-1}.

$$\text{Thus, } [H^+] = 1 \times 10–3 \text{ mol litre}^{-1}$$

$pH = –log10[H^+] = –log1010^{-3}= –(–3 \times log_{10}) = –(3 \times 1) = 3$

Thus, pH = 3

Example 2:

What would be the pH of an aqueous solution of sulphuric acid which is 5 × 10^{-5}mol litre^{-1} in concentration.

- **Solution:**

Sulphuric acid dissociates in water as:

$$H2SO_{4(aq)} \rightarrow 2\ H^{+}_{(aq)} + SO^{2-}{}_{4(aq)}$$

Each mole of sulphuric acid gives two mole of H^{+} ions in the solution. One litre of H_2SO_4 solution contains 5 × 10^{-5} moles of H_2SO_4 which would give 2 × 5 × 10–5 = 10 × 10–5 or 1.0 × 10^{-4} moles of H^{+} ion in one litre of the solution.

Therefore,

$[H+] = 1.0 \times 10^{-4}$mol litre^{-1}

$pH = -\log_{10}[H^{+}] = -\log_{10}10^{-4} = -(-4 \times \log_{10}10) = -(-4 \times 1) = 4$

Example 3:

Calculate the pH of 1 × 10–4 molar solution of NaOH.

- **Solution:**

NaOH is a strong base and dissociates in its solution as:

$$NaOH_{(aq)} \rightarrow Na^{+}_{(aq)} + OH^{-}_{(aq)}$$

One mole of NaOH would give one mole of OH– ions. Therefore,

$$[OH^{-}] = 1 \times 10^{-4}\text{mol litre}^{-1}$$

$$pOH = -\log_{10}[OH^{-}] = -\log_{10} \times [10^{-4}]$$

$$= -(-4 \times \log_{10}10) = -(-4) = 4$$

$$\text{Since, } pH + pOH = 14$$

$$pH = 14 - pOH = 14 - 4 = 10$$

Example 4:

Calculate the pH of a solution in which the concentration of the hydrogen ions is 1.0×10^{-8}mol litre^{-1}.

- **Solution:**

Here, although the solution is extremely dilute, the concentration given is not of an acid or a base but that of H^+ ions. Hence, the pH can be calculated from the relation:

$pH = -\log10[H^+]$
given $[H^+] = 1.0 \times 10^{-8}$mol litre^{-1}
$pH = -\log_{10}10^{-8} = -(-8 \times \log_{10}10)$
$= -(-8 \times 1) = 8$

Example 5:

If the pH of a solution is 4.5, what is its pOH?

Solution:

pH + pOH = 14
pOH = 14 – 4.5 = 9.5
pOH = 9.5

Periodic table with properties

9th STD

Unit- 12- Periodic Classification of Elements

9th STD	**Unit- 12**	**Periodic Classification of Elements**
10th STD	**Unit- 7**	**Atoms And Molecules**
	Unit- 8	**Periodic Classification Of Elements**
11 th vol - 1	**Unit - 3**	**Periodic Classification of Elements**

Introduction

- We live in the world of substances with great diversity. Substances are formed by the combination of various elements. All the elements are unique in their nature and property. To categorize these elements according to their properties, scientists started to look for a way. In 1800, there were only 31 known elements. By 1865, their number became 63. Now 118 elements have been discovered. As different elements were being discovered, scientists gathered more and more information about the properties of these elements. They found it difficult to organize all that was known about the elements. They started looking for some pattern in their properties, on the basis of which they could study such a large number of elements with ease. Let us discuss the concepts of classification of elements proposed by various scientists from early to modern period.

Early Concepts of Classification of Elements

Dobereiner's Triads

- In 1817, Johann Wolfgang Dobereiner, a German chemist, suggested a method of grouping elements based on their relative atomic masses. He arranged the elements into groups containing three elements each. He called these groups as 'triads' (tri – three).

- Dobereiner showed that when the three elements in a triad are arranged in the ascending order of their atomic masses, the atomic mass of the middle element is nearly the same as average of atomic masses of other two elements. This statement is called the Dobereiner's law of triads. Table 12.1 shows the law of triads proposed by Dobereiner.

- Example: In the triad group (1), arithmetic mean of atomic masses of 1st and 3rd elements, (6.9 + 39.1)/2 = 23. So the atomic mass of Na (middle element) is 23.

Triad Group (1)		Triad Group (2)		Triad Group (3)	
Element	Atomic Mass	Element	Atomic Mass	Element	Atomic Mass
Li	6.9	Cl	35.5	Ca	40.1
Na	23	Br	79.9	Sr	87.6
K	39.1	I	126.9	Ba	137.3

Limitations:

- Dobereiner could identify only three triads from the elements known at that time and all elements could not be classified in the form of triads.

- The law was not applicable to elements having very low and very high atomic mass.

Newlands' Law of Octaves

- In 1866, John Newlands arranged 56 known elements in the increasing order of their atomic mass. He observed that every eighth element had properties similar to those of the first element like the eighth note in an octave of music is similar to the first. This arrangement was known as 'law of octaves'.

- The octave of Indian music system is sa, re, ga, ma, pa, da, ni, sa. The first and last notes of this octave are same i.e. sa. Likewise, in the Newlands' table of octaves, the element 'F' is eighth from the element 'H', thus they have similar properties.

Limitations:

- There are instances of two elements being fitted into the same slot, e.g. cobalt and nickel.

- Some elements, totally dissimilar in their properties, were fitted into the same group. (Arrangement of Co, Ni, Pd, Pt and Ir in the row of halogens)

- The law of octaves was not valid for elements that had atomic masses higher than that of calcium.

- Newlands' table was restricted to only 56 elements and did not leave any room for new elements.

- Discovery of inert gases (Neon. Argon….) at later stage made the 9th element similar to the first one. Eg: Neon between Fluorine and Sodium.

Mendeleev's Periodic Table

- In 1869, Russian chemist, Dmitri Mendeleev observed that the elements of similar properties repeat at regular intervals when the elements are arranged in the order of their atomic masses. Based on this, he proposed the law of periodicity which states that "the physical and chemical properties of elements are the periodic functions of their atomic masses". He arranged 56 elements known at that time according to his law of periodicity. This was best known as the short form of periodic table.

Features of Mendeleev's Periodic Table:

- It has eight vertical columns called 'groups' and seven horizontal rows called 'period'.
- Each group has two subgroups 'A' and 'B'. All the elements appearing in a group were found to have similar properties.
- For the first time, elements were comprehensively classified in such a way that elements of similar properties were placed in the same group

- It was noticed that certain elements could not be placed in their proper groups in this manner. The reason for this was wrongly determined atomic masses. Consequently those wrong atomic masses were corrected. Eg: The atomic mass of beryllium was known to be 14. Mendeleev reassessed it as 9 and assigned beryllium a proper place.

- Columns were left vacant for elements which were not known at that time and their properties also were predicted. This gave motivation to experiment in Chemistry. Eg: Mendeleev gave names EkaAluminium andEka Silicon to those elements which were to be placed below Aluminium and Silicon respectively in the periodic table and predicted their properties. The discovery of Germanium later on, during his life time, proved him correct.

No.		No.		No.		No.		No.		No.		No.		No.	
H	1	F	8	Cl	15	Co & Ni	22	Br	29	Pd	36	I	42	Pt & Ir	50
Li	2	Na	9	K	16	Cu	23	Rb	30	Ag	37	Cs	44	Os	51
G	3	Mg	10	Ca	17	Zn	24	Sr	31	Cd	38	Ba & V	45	Hg	52
Bo	4	Al	11	Cr	19	Y	25	Ce & La	33	U	40	Ta	46	Ti	53
C	5	Si	12	Ti	18	In	26	Zr	32	Sn	39	W	47	Pb	54
N	6	P	13	Mn	20	As	27	Di & Mo	34	Sb	41	Nb	48	Bi	55
O	7	S	14	Fe	21	Se	28	Ro & Ru	35	To	43	Au	49	Th	56

Limitations:

- ❖ Elements with large difference in properties were included in the same group. Eg: Hard metals like copper (Cu) and silver (Ag) were included along with soft metals like sodium (Na) and potassium (K).
- ❖ No proper position could be given to the element hydrogen. Non-metallic hydrogen was placed along with metals like lithium (Li), sodium (Na) and potassium (K).
- ❖ The increasing order of atomic mass was not strictly followed throughout. Eg. Co & Ni, Te& I.
- ❖ No place for isotopes in the periodic table.

Properties of Germanium

Property	**Mendeleev's prediction (1871)**	**Actual property (1886)**
Atomic Mass	About 72	72.59
Specific Gravity	5.5	5.47
Colour	Dark grey	Dark grey
Formula of Oxide	EsO_2	GeO_2
Nature of Chloride	$EsCl_4$	$GeCl_4$

Modern Periodic Table

- ❖ In 1913, the English Physicist Henry Moseley, through his X-ray diffraction experiments, proved that the properties of elements depend on the atomic number and not on the atomic mass. Consequently, the modern periodic table was prepared by arranging elements in the increasing order of their atomic number.

- ❖ This modern periodic table is the extension of the original Mendeleev's periodic table and known as the long form of periodic table.

Modern Periodic Law

- Atomic number of an element (Z) indicates the number of protons (positive charge) or the number of electrons (negative charge). The physical and chemical properties of elements depend not only on the number of protons but also on the number of electrons and their arrangements (electronicconfiguration) in atoms. Hence, the modern periodic law can be stated as follows: "The chemical and physical properties of elements are periodic functions of their atomic numbers". Based on the modern periodic law, the modern periodic table is derived.

Features of Modern Periodic Table

- All the elements are arranged in the increasing order of their atomic number.
- The horizontal rows are called periods. There are seven periods in the periodic table.
- The elements are placed in periods based on the number of shells in their atoms.
- Vertical columns in the periodic table starting from top to bottom are called groups. There are 18 groups in the periodic table.
- Based on the physical and chemical properties of elements, they are grouped into various families.

Groups in modern periodic table

Group	Families
1	Alkali metals
2	Alkaline earth metals
3 to 12	Transition metals
13	Boron Family
14	Carbon Family
15	Nitrogen Family
16	Oxygen Family (or) Chalcogen Family
17	Halogens
18	Noble gases

Classification of elements into blocks

- We know that the electrons in an atom are accommodated in shells around the nucleus. Each shell consists of one or more subshells in which the electrons are distributed in certainmanner. These subshells are designated as s, p, d, and f. Based on the arrangement of electrons in subshells, the elements of periodic table are classified into four blocks namely s, p, d and f blocks.

- **s-Block Elements**: It includes group 1 (alkali metals) and group 2 (alkaline earth metals) elements. They are also called as representative elements. The elements of group 1 (except hydrogen) are metals. They react with water to form solutions that change the colour of a vegetable dye from red to blue. These solutions are said to be highly alkaline or basic. Hence they are called alkali metals.

- The elements of group 2 are also metals. They combine with oxygen to form oxides, formerly called 'earths', and these oxides produce alkaline solutions when they are dissolved in water. Hence, these elements are called alkaline earth metals.

- **(2) p-Block Elements:** These elements are in group 13 to 18 in the periodic table. They include boron, carbon, nitrogen, oxygen, fluorine families in addition to noble gases (Except helium). They are also called as representative elements. The p-block is home to the biggest variety of elements and is the only block that contains all three types of elements: metals, nonmetals, and metalloids.

- **(3) d-Block Elements:** It includes group 3 to group 12 elements. They are found in the centre of the periodic table. Their properties are intermediate to that of s block and p block elements and so they are called transition elements.

- **(4) f – Block Elements:** It includes 14 elements after (Lanthanum) La (57), called Lanthanoides and 14 elements after (Actinium) Ac (89), called Actinoides. They are placed at the bottom of the periodic table. They are also called as inner Transition elements.

Advantages of the Modern Periodic Table

- The table is based on a more fundamental property i.e., atomic number.
- It correlates the position of the element with its electronic configuration more clearly.
- The completion of each period is more logical. In a period, as the atomic number increases, the energy shells are gradually filled up until an inert gas configuration is reached.
- It is easy to remember and reproduce.
- Each group is an independent group and the idea of subgroups has been discarded.
- One position for all isotopes of an element is justified, since the isotopes have the same atomic number.
- The position of the eighth group (in Mendeleev's table) is also justified in this table. All transition elements have been brought in the middle as the properties of transition elements are intermediate between left portion and right portion elements of the periodic table.
- The table completely separates metals from nonmetals. The nonmetals are present in upper right corners of the periodic table.
- The positions of certain elements which were earlier misfit (interchanged) in the Mendeleev's periodic table are now justified because it is based on atomic number of the elements.
- Justification has been offered for placing lanthanides and actinides at the bottom of the periodic table.

Position of hydrogen in the periodic table

- Hydrogen is the lightest, smallest and first element of the periodic table. Its electronic configuration (1s1) is the simplest of all the elements. It occupies a unique position in the periodic table. It behaves like alkali metals as well as halogens in its properties.

- In the periodic table, it is placed at the top of the alkali metals.

(1) Hydrogen can lose its only electron to form a hydrogen ion (H+) like alkali metals.
(2) It can also gain one electron to form the hydride ion (H-) like halogens.
(3) Alkali metals are solids while hydrogen is a gas.

- Hence the position of hydrogen in the modern periodic table is still under debate as the properties of hydrogen are unique.

Position of Rare Gases

- The elements Helium, Neon, Argon, Krypton, Xenon and Radon of group 18 in the periodic table are called as Noble gases or Rare gases. They are monoatomic gases and do not react with other substances easily, due to completely filled subshells. Hence they are called as inert gases. They are found in very small quantities and hence they are called as rare gases.

Metals, Non-Metals and Metalloids
Metals

- Metals are typically hard, shiny, malleable (can be made as sheet), fusible and ductile (can be drawn into wire) with good electrical and thermal conductivity. Except mercury, most of the metals are solids at room temperature. Metals occupy larger area in the periodic table and are categorized as:

  (1) Alkali metals. e.g. Lithium to Francium (top to bottom)
  (2) Alkaline earth metals. e.g: Beryllium to Radium (top to bottom)
  (3) Transition Metals. Group III B to II A
  (4) P-Block metals. e.g: Al, Ga, In, Tl, Sn, Pb and Bi.

Non-metals

- A non-metal is an element that does not have the characters like hardness, shiny, malleable, suitable and ductile. In other words, a non-metal is an element that does not have the properties of metal. e.g. All non metals are arranged in P-Block only. P-Block non metals: C, N O, P, S, Se, Halogen (F, Cl, Br and I) and inert gases (HeyoRn).

Metalloids

- Elements which have the properties of both metals and non-metals are called as metalloids. (eg) Boron, Arsenic.

Alloys

- During 3500 BC(BCE), people used an alloy named 'bronze'. The idea of making an alloy was quite old. The majority of the metallic substances used today are alloys. Alloys are mixtures of two or more metals and are formed by mixing molten metals thoroughly. Rarely nonmetals are also mixed with metals to produce alloys.

- It is generally found that alloying produces a metallic substance that has more useful properties than the original pure metals from which it is made. For example, the alloy brass is made from copper and zinc.

Advantages of alloys

- Alloys do not get corroded or get corroded to very less extent.
- They are harder and stronger than pure metals (Example: Gold is mixed with copper and it is harder than pure gold).
- They have less conductance than pure metals (Example: Copper is good conductor of heat and electricity whereas brass and bronze are not good conductors).
- Some alloys have lower melting point than pure metals (Example: Solder is an alloy of lead and tin which has lower melting point than each of the metals).
- When metal is alloyed with mercury, it is called amalgam.

10th STD
7.ATOMS AND MOLECULES

INTRODUCTION

- You have learnt, in your lower classes that matter is around us everywhere. Matter is made of atoms. Curiously the idea of atom was first proposed by the Greek philosophers in the fifth century BC (BCE). But, their theory was more philosophical than scientific.

- The first scientific theory of the atom was proposed by John Dalton. Few of the postulates of Dalton's theory about an atom were found incorrect by the later on studies made by J.J. Thomson, Rutherford, Neils Bohr and Schrodinger. In the light of the result of the researches most of the limitations of the Dalton's theory wereremoved and a new theory known as the modern atomic theory was put forward. **'The main postulates of modern atomic theory'** are as follows:

1. An atom is no longer indivisible (after the discovery of the electron, proton, and neutron).
2. Atoms of the same element may have different atomic mass. (discovery of isotopes $_{17}Cl^{35}$, $_{17}Cl^{37}$).
3. Atoms of different elements may have same atomic masses (discovery of Isobars $_{18}Ar^{40}$, $_{20}Ca^{40}$).
4. Atoms of one element can be transmuted into atoms of other elements. In other words, atom is no longer indestructible (discovery of artificial transmutation).
5. Atoms may not always combine in a simple whole number ratio (E.g. Glucose $C_6H_{12}O_6$ C:H:O = 6:12:6 or 1:2:1 and Sucrose $C_{12}H_{22}O_{11}$ C:H:O = 12:22:11).
6. Atom is the **smallest particle** t**hat takes part in a chemical reaction.**
7. The mass of an atom can be converted into energy ($E = mc^2$).

- The modern atomic theory is the basis for all the studies of chemical and physical processes that involve atoms. You have studied the most fundamental ideas about an atom in your lower classes. Let us discuss some more concepts about atoms in this lesson.

ATOM AND ATOMIC MASS

- As you know, anything that has mass and occupies space is called matter. Atoms are the building blocks of matter. Since matter has mass, it must be due to its atoms. According to the modern atomic theory, an atom contains subatomic particles such as protons, neutrons and electrons. Protons and neutrons have considerable mass, but electrons don't have such a considerable mass. Thus, the mass of an atom is mainly contributed by its protons and neutrons and hence the sum of the number of protons and neutrons of an atom is called its mass number.

- Individual atoms are very small and it is difficult to measure their masses. You can measure the mass of macroscopic materials in gram or kilogram. The mass of an atom is measured in atomic mass unit (amu).

- Atomic mass unit is one-twelfth of the mass of a carbon-12 atom; an isotope of carbon, which contains 6 protons and 6 neutrons.

- (Note: The symbol ‘amu’ is no longer used in the modern system and instead, it uses the symbol ‘u’ to denote unified atomic mass. The mass of a proton or neutron is approximately 1 amu).

Relative Atomic Mass (RAM)

- As an atom is very small, its absolute mass cannot be determined directly. The early pioneers of chemistry used to measure the atomic mass of an atom relative to an atom of another element. They measured the masses of equal number of atoms of two or more elements at a time, to determine their relative masses. They established one element as a standard, gave it an arbitrary value of atomic mass and using this value they measured the relative mass of other elements. The mass obtained by this way is called relative atomic mass. In the beginning, the mass of hydrogen atom was chosen as a standard and masses of other atoms were compared with it, because of the existence of isotopic character of hydrogen ($_1H^1$, $_1H^2$, $_1H^3$). Later hydrogen atom was replaced by oxygen atom as the standard. Now, the stable isotope of carbon (C-12) with atomic mass 12 is used as the standard for measuring the relative atomic mass of an element.

- Relative atomic mass of an element is the ratio between the average mass of its isotopes to $\frac{1}{12^{th}}$ part of the mass of a carbon-12 atom. It is denoted as Ar. It is otherwise called “Standard Atomic Weight”.

Relative Atomic Mass

$$(\text{Ar}) = \frac{\text{Average mass of the isotopes of the element}}{\frac{1}{12^{th}}\text{ of the mass of one Carbon-12 atom}}$$

- Modern methods of determination of atomic mass by Mass Spectrometry uses C-12 as standard. For most of the elements, the relative atomic mass is very closer to a whole number and it is rounded off to a whole number, to make calculations easier.

Element	Symbol	A_r
Hydrogen	H	1
Carbon	C	12
Nitrogen	N	14
Oxygen	O	16
Sodium	Na	23
Magnesium	Mg	24
Sulphur	S	32

Relative Atomic Mass is only a ratio, so it has no unit. If the atomic mass of an element is expressed in grams, it is called as **Gram Atomic Mass**
Gram Atomic Mass of hydrogen = 1 g
Gram Atomic Mass of carbon = 12 g
Gram Atomic Mass of nitrogen = 14 g
Gram Atomic Mass of oxygen = 16 g

Average Atomic Mass (AAM)

- How can one measure the atomic mass of an element? It is somewhat more complicated because most of the naturally occuring elements exist as a mixture of isotopes, each of which has its own mass. Thus, it is essential to consider this isotopic mixture while calculating the atomic mass of an element.

- The average atomic mass of an element is the weighted average of the masses of its naturally occurring isotopes.

- But, the abundance of isotopes of each element may differ. So, the abundancy of all these isotopes are taken into consideration while calculating the atomic mass. Then, what do we mean by a weighted average? Let usconsider an element which exists as a mixture of 50% of an isotope having a mass of 9 amu, and 50% of another isotope having a mass of 10 amu. Then, its average atomic mass is calculated by the following equation:

  Average atomic mass

= (Mass of 1st isotope × % abundance of 1st isotope) + (Mass of 2nd isotope × % abundance of 2nd isotope)

Thus, for the given element the average

- atomic mass $= \left(9 \times \frac{50}{100}\right) + \left(10 \times \frac{50}{100}\right)$

  $= 4.5 + 5 = 9.5 \text{ amu}$

- (Note: In the calculations involving percentages, you need to convert percentage abundance into fractional abundance. For example, 50 percent is converted into 50/100 or 0.50 as shown in the a foresaid calculation.)

- The atomic masses of elements, given in the periodic table, are average atomic masses. Sometimes, the term atomic weight is used to mean average atomic mass. It is observed, from the periodic table that atomic masses of most of the elements are not whole numbers. For instance, the atomic mass of carbon given in the periodic table is 12.01 amu, not 12.00 amu. The reason is that while calculating the atomic mass of carbon, both of its natural isotopes such as carbon-12. and carbon-13 are considered. The natural abundance of C-12 and C-13 are 98.90 % and 1.10 % respectively. The average of the atomic mass of carbon is calculated as follows:

  Average atomic mass of carbon

$$= \left(12 \times \frac{98.9}{100}\right) + \left(13 \times \frac{1.1}{100}\right)$$

$$= (12 \times 0.989) + (13 \times 0.011)$$

$$= 11.868 + 0.143 = 12.011 \text{ amu}$$

- So it is important to understand that if it is mentioned that the atomic mass of carbon is 12 amu, it refers to the average atomic mass of the carbon isotopes, not the mass of the individual atoms of carbon.

Atomic Number	Name	Symbol	Atomic Mas (amu)
1	Hydrogen	H	1.008
2	Helium	He	4.003
3	Lithium	Li	6.941
4	Beryllium	Be	9.012

5	Boron	B	10.811

Calculation of average atomic mass – Solved Examples

- Example 1: Oxygen is the most abundant element in both the Earth's crust and the human body. It exists as a mixture of three stable isotopes in nature.

Isotope	**Mass (amu)**	**% abundance**
${}_{8}O^{16}$	15.9949	99.757
${}_{8}O^{17}$	16.9991	0.038
${}_{8}O^{18}$	17.9992	0.205

The atomic mass of

oxygen = (15.9949 × 0.99757) + (16.9991 ×0.00038) + (17.9992 × 0.00205)
= 15.999 amu.

- Example 2: Boron naturally occurs as a mixture of boron-10 (5 protons + 5 neutrons) and boron-11 (5 protons + 6 neutrons) isotopes. The percentage abundance of B-10 is 20 and that of B-11 is 80. Then, the atomic mass of boron is calculated as follows:

Atomic mass of

boron $= \left(10 \times \frac{20}{100}\right) + \left(11 \times \frac{80}{100}\right)$

$= (10 \times 0.20) + (11 \times 0.80)$

$= 2 + 8.8$

$= 10.8$ amu

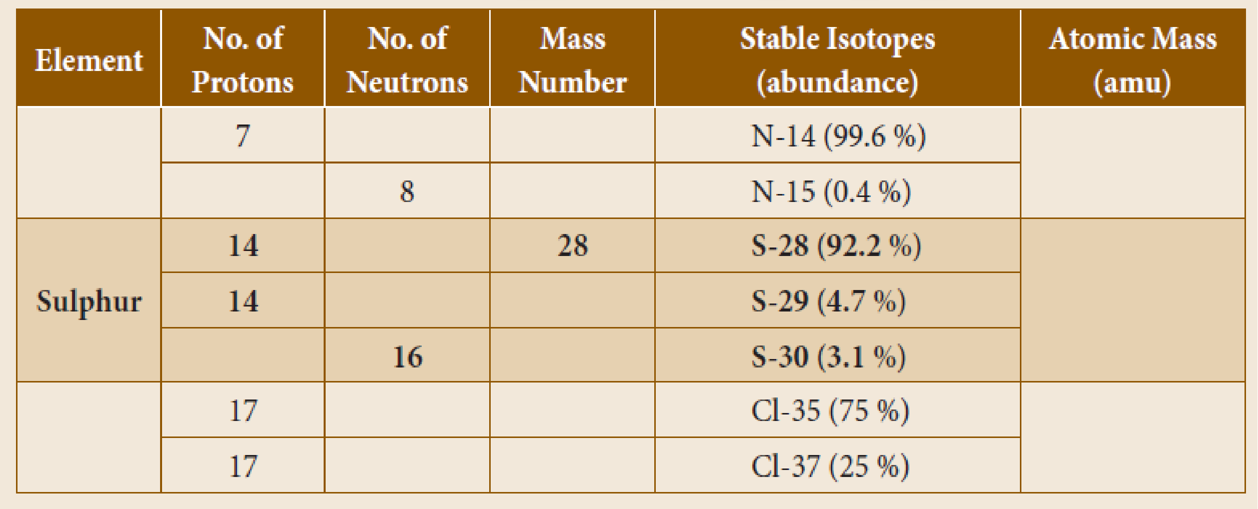
MOLECULE AND MOLECULAR MASS

- Except noble gases, atoms of most of the elements are found in the combined form with itself or atoms of other elements. It is called as a molecule. A molecule is a combination of two or more atoms held together by strong chemical forces of attraction, i.e. chemical bonds.

Classification of molecules

- A molecule may contain atoms of the same element or may contain atoms of two or more elements joined in a fixed ratio, in accordance with the law of definite proportions. Thus, a molecule may be an element or a compound. If the molecule is made of similar kind of atoms, then it is called homoatomic molecule.
- The molecule that consist of atoms of different elements is called heteroatomic molecule. A compound is a heteroatomic molecule. The number of atoms present in the molecule is called its 'atomicity'.

- Let us consider oxygen. Oxygen gas exists in two allotropic forms: Oxygen (O_2) and Ozone (O_3). In oxygen molecule, there are two oxygen atoms. So its atomicity is two. Since both the atoms are similar, oxygen (O_2) is a homodiatomic molecule. Other elementsthat exist as diatomic molecules are hydrogen (H_2), nitrogen (N_2) and halogens: fluorine (F_2), chlorine (C_{l2}), bromine (Br_2) and iodine (I_2).



Element	No. of Protons	No. of Neutrons	Mass Number	Stable Isotopes (abundance)	Atomic Mass (amu)
	7			N-14 (99.6 %)	
		8		N-15 (0.4 %)	
Sulphur	**14**		**28**	**S-28 (92.2 %)**	
	14			**S-29 (4.7 %)**	
		16		**S-30 (3.1 %)**	
	17			Cl-35 (75 %)	
	17			Cl-37 (25 %)	

- Ozone (O3) contains three oxygen atoms and hence it is called homotriatomic molecule. If a molecule contains more than three atoms, then it is called polyatomic molecule.

- Conisder hydrogen chloride. It consists of two atoms, but of different elements, i.e. hydrogen and chlorine. So, its atomicity is two. It is a heterodiatomic molecule. Similarly, the water molecule contains two hydrogen atoms and one oxygen atom. So its atomicity is three. It is a heterotriatomic molecule.

Classify the following molecules based on their atomicity and fill in the table:

- Fluorine (F_2), Carbon dioxide (CO_2), Phosphorous (P_4), Sulphur (S_8), Ammonia (NH_3), Hydrogen iodide (HI), Sulphuric Acid (H_2SO_4), Methane (CH_4), Glucose ($C_6H_{12}O_6$), Carbon monoxide (CO)

Relative Molecular Mass (RMM)

- As the molecules are made of atoms, they also have their own mass. The mass of the molecule of an element or compound is measured in the C-12 scale and hence called relative molecular mass.

- The Relative Molecular Mass of a molecule is the ratio between the mass of one molecule of the substance to $\frac{1}{12^{th}}$ mass of an atom of Carbon -12.

Relative Molecular Mass is only a ratio. So, it has no unit. If the molecular mass of a compound is expressed in grams, it is called Gram Molecular Mass.
Gram Molecular Mass of water = 18 g
Gram Molecular Mass of carbon dioxide = 44 g
Gram Molecular Mass of ammonia = 17 g
Gram Molecular Mass of HCl = 36.5 g

- The relative molecular mass is obtained by adding together the relative atomic masses of all the atoms present in a molecule.

Calculation of relative molecular mass – Solved examples:

- **Example 1:** Relative molecular mass of sulphuric acid (H2SO4) is calculated as follows: Sulphuric acid conatins 2 atoms of hydrogen, 1 atom of sulphur and 4 atoms of oxygen.

  Therefore, Relative molecular mass of sulphuric acid
  = (2 × mass of hydrogen) + (1 × mass of sulphur) + (4 × mass of oxygen)
  = (2 × 1) + (1 × 32) + (4 × 16)
  = 98
- i.e., one molecule of H_2SO^4 is 98 times as heavy as $\frac{1}{12^{th}}$ of the mass of a carbon –12.

- **Example 2:** Relative molecular mass of water (H2O) is calculated as follows: A water molecule is made of 2 atoms of hydrogen and one atom of oxygen.

  So, the relative molecular mass of water
  = (2 × mass of hydrogen) + (1 × mass of oxygen)
  = (2 × 1) + (1 × 16)
  = 18

- i.e., one molecule of H_2O is 18 times as heavy as $\frac{1}{12^{th}}$ of the mass of a carbon – 12.

DIFFERENCE BETWEEN ATOMS AND MOLECULES

- Even though atoms are the basic components of molecules, they differ in many aspects when compared to the molecules. Table 7.5 consolidates the major difference between atoms and molecules.

Atom	Molecule
An atom is the smallest particle of an element	A molecule is the smallest particle of an element or compound.
Atom does not exist in free state except in a noble gas	Molecule exists in free a state
Except some of noble gas, other atoms are highly reactive	Molecules are less reactive
Atom does not have a chemical bond	Atoms in a molecule are held by chemical bonds

MOLE CONCEPT

- So far we discussed about matters in terms of individual atoms and molecules. Atomic mass units provide a relative scale for the masses of the elements. Since the atoms have such small masses, no usable scale can be devised to weigh them in the calibrated units of atomic mass units. In any real situation, we deal with macroscopic samples containing enormous number of atoms. Therefore, it is convenient to have a special unit to describe a very large number of atoms. The idea of a 'unit' to denote a particular number of objects is not new. For example, the pair (2 items) and the dozen (12 items), are all familiar units. Chemists measure atoms and molecules in 'moles'. So, you can now understand that 'mole' denotes a number of particles.

- In the SI system, the mole (mol) is the amount of a substance that contains as many elementary entities (atoms, molecules, or other particles) as there are atoms in exactly 12 g (or 0.012 kg) of the carbon-12 isotope. The actual number of atoms in 12 g of carbon-12 is determined experimentally. This is called Avogadro's Number (N_A), named after an Italian scientist Amedeo Avogadro who proposed its significance. Its value is 6.023×10^{23}. So one mole of a substance contains 6.023×10^{23} entities. Thus, 5 moles of oxygen molecules contain $5 \times 6.023 \times 10^{23}$ molecules.

- Mole Concept: The study of the collection of particles by using mole as the counting unit, in order to express the mass and volume of such unit particles in a bulk of matter is known as mole concept.

- The number of moles of a substance can be calculated by various means depending on the data available, as follows:

- Number of moles of molecules.
- Number of moles of atoms.
- Number of moles of a gas (Standard molar volume at STP = 22.4 litre).
- Number of moles of ions

Mole of atoms:

- One mole of an element contains 6.023×10^{23} atoms and it is equal to its gram atomic mass.
- i.e., one mole of oxygen contains 6.023×10^{23} atoms of oxygen and its gram atomic mass is 16 g.

Mole of molecules:

- One mole of matter contains 6.023×10^{23} molecules and it is equal to its gram molecular mass.

- i.e., one mole of oxygen contains 6.023×10^{23} molecules of oxygen and its gram molecular mass is 32 g.

Molar volume:

- One mole of any gas occupies 22.4 litre or 22400 ml at S.T.P. This volume is called as molar volume.

Calculation of number of moles by Different modes

Number of moles = Mass / Atomic Mass
= Mass / Molecular mass
= Number of Atoms / 6.023 × 1023
= Number of Molecules / 6.023 × 1023

PERCENT COMPOSITION

- So for, we were dealing with the number of entities present in a given substance. But many times, the information regarding the percentage of a particular element present in a compound is required.

- The percentage composition of a compound represents the mass of each element present in 100 g of the compound.

- Let us understand the percentage composition of oxygen and hydrogen by takingthe example of H_2O. It can be calculated using the formula

  Mass % of an element

  $$= \frac{\text{mass of that element in the compound}}{\text{molar mass of the compound}} \times 100$$

  molar mass of H2O = 2(1) + 16 =18 g

  Mass % of hydrogen = $\frac{2}{18} \times 100 = 11.11$ %

  Mass % of oxygen = $\frac{16}{18} \times 100 = 88.89$ %

- This percentage composition is useful to determine the empirical formula and molecular formula.

- **Example 1:** Find the mass percentage composition of methane (CH_4).

  Molar mass of CH4 = 12 + 4 = 16 g

  Mass % of carbon = $\frac{12}{16} \times 100$ = 75 %

  Mass % of hydrogen = $\frac{4}{16} \times 100$ = 25 %

AVOGADRO HYPOTHESIS

- In 1811 Avogadro framed a hypothesis based on the relationship between the number of molecules present in equal volumes of gases in different conditions.

- The Avogadro's law states that "equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules"

- It follows that the volume of any given gas must be proportional to the number of molecules in it. If 'V' is the volume and 'n' is thenumber of molecules of a gas, then Avogadro law is represented, mathematically, as follows:

  $$V \propto n$$
  $$V = \text{constant} \times n$$

- Thus, one litre (1 dm3) of hydrogen contains the same number of molecules as in one litre of oxygen, i.e. the volume of the gas is directly proportional to the number of molecules of the gas.

Explanation

- Let us consider the reaction between hydrogen and chlorine to form hydrogen chloride gas

$H_{2(g)} + Cl_{2(g)} \rightarrow 2\ HCl_{(g)}$
1 vol + 1 vol → 2 volumes

- According to Avogadro's law 1 volume of any gas is occupied by "n" number of molecules. n molecules + n molecules → 2n molecules

if n = 1 then

1molecule + 1 molecule → 2 molecules.
½ molecule + ½ molecule → 1 molecule

- 1 molecule of hydrogen chloride gas is made up of ½ molecule of hydrogen and ½ molecule of chlorine. Hence, the molecules can be subdivided. This law is in agreement with Dalton's atomic theory.

APPLICATIONS OF AVOGADRO'S LAW

1. It explains Gay-Lussac's law.
2. It helps in the determination of atomicity of gases.
3. Molecular formula of gases can be derived using Avogadro's law
4. It determines the relation between molecular mass and vapour density.
5. It helps to determine gram molar volume of all gases (i.e, 22.4 litre at S.T.P)

RELATIONSHIP BETWEEN VAPOUR DENSITY AND RELATIVE MOLECULAR MASS

Relative molecular mass: (Hydrogen scale)

- The Relative Molecular Mass of a gas or vapour is the ratio between the mass of one molecule of the gas or vapour to mass of one atom of Hydrogen.

Vapour Density:

- Vapour density is the ratio of the mass of a certain volume of a gas or vapour, to the mass of an equal volume of hydrogen, measured under the same conditions of temperature and pressure.

Vapour Density (V.D.)

$$= \frac{\text{Mass of a given volume of gas or vapour at S.T.P.}}{\text{Mass of the same volume of hydrogen}}$$

- According to Avogadro's law, equal volumes of all gases contain equal number of molecules.

- Thus, let the number of molecules in one volume = n, then

$$= \frac{\text{Mass of 'n' molecules of a gas or vapour at S.T.P}}{\text{Mass of 'n' molecules of hydrogen}}$$

- Cancelling 'n' which is common, you get
V.D.

$$= \frac{\text{Mass of 1 molecule of a gas or vapour at S.T.P.}}{\text{Mass of 1 molecules of hydrogen However, since hydrogen is diatomic}}$$

V.D.

$$= \frac{\text{Mass of 1 molecule of a gas or vapour at S.T.P.}}{\text{Mass of 2 atoms of hydrogen}}$$

- When you compare the formula of vapour density with relative molecular mass, they can be represented as

V.D.

$$= \frac{\text{Mass of 1 molecule of a gas or vapour at S.T.P.}}{2 \times \text{Mass of 1 atom of hydrogen}}$$

Relative molecular mass (hydrogen scale)

$$= \frac{\text{Mass of 1 molecule of a gas or vapour at STP}}{\text{Mass of 1 atom of hydrogen}}$$

- You can therefore substitute the above equation to an Eqn 7.1 and arrive at the following formula

$$\text{V.D.} = \frac{\text{Relative molecular mass}}{2}$$

Now on cross multiplication, you have

2 × vapour density = Relative molecular mass of a gas (Or) Relative molecular mass = 2 × Vapour density

SOLVED PROBLEMS
I. Calculation of molar mass

- Calculate the gram molar mass of the following.
- H_2O
- CO_2
- $Ca_3(PO_4)_2$

Solution:
1) H_2O
Atomic masses of H = 1, O = 16

Gram molar mass of $H_2O = (1 \times 2) + (16 \times 1) = 2 + 16$
Gram molar mass of H_2O = 18 g

2) CO_2
Atomic masses of C = 12, O = 16

Gram molar mass of $CO_2 = (12 \times 1) + (16 \times 2) = 12 + 32$
Gram molar mass of CO_2 = 44 g

3) $Ca_3(PO_4)_2$
Atomic masses of Ca = 40, P = 30, O = 16.

Gram molar mass of $Ca_3(PO_4)_2 = (40 \times 3) + [30 + (16 \times 4)] \times 2$
$= 120 + (94 \times 2) = 120 + 188$
Gram molar mass of $Ca_3(PO_4)_2$ = 308 g

Calculation based on number of moles from mass and volume

Calculate the number of moles in 46 g of sodium?

$$\text{Number of moles} = \frac{\text{Mass of the element}}{\text{Atomic mass of the element}}$$

= 46 / 23
= 2 moles of sodium

5.6 litre of oxygen at S.T.P

$$\text{Number of moles} = \frac{\text{Given volume of } O_2 \text{ at S.T.P}}{\text{Molar volume at S.T.P}}$$

$$\text{Number of moles of oxygen} = \frac{5.6}{22.4} = 0.25 \text{ mole of oxygen}$$

Calculate the number of moles of a sample that contains 12.046 × 1023 atoms of iron ?

$$\text{Number of moles} = \frac{\text{Number of atoms of iron}}{\text{Avogadro's number}}$$

= $12.046 \times 10^{23} / 6.023 \times 10^{23}$
= 2 moles of iron

Calculation of mass from mole
Calculate the mass of the following

0.3 mole of aluminium (Atomic mass of Al = 27)

$$\text{Number of moles} = \frac{\text{Mass of Al}}{\text{Atomic mass of Al}}$$

Mass = No. of moles × atomic mass
So, mass of Al = 0.3 × 27
= 8.1 g

2.24 litre of SO_2 gas at S.T.P

Molecular mass of SO_2= 32 + (16 × 2) = 32 + 32 = 64

$$\text{Number of moles of } SO_2 = \frac{\text{Given volume of } SO_2 \text{ at S.T.P}}{\text{Molar volume } SO_2 \text{at S.T.P}}$$

$$\text{Number of moles of } SO_2 = \frac{2.24}{22.4} = 0.1 \text{ mole}$$

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molecular mass}}$$

Mass = No. of moles × molecular mass
Mass = 0.1 × 64
Mass of SO_2 = 6.4 g

1.51×10^{23}molecules of water Molecular mass of H_2O = 18

$$\text{Number of moles} = \frac{\text{Number of molecules of water}}{\text{Avogadro's number}}$$

$= 1.51 \times 10^{23} / 6.023 \times 10^{23}$
= 1 / 4
= 0.25 mole

$$\text{Number of moles} = \frac{\text{Mass}}{\text{Molecular mass}}$$

0.25 = mass / 18
Mass = 0.25 × 18
Mass = 4.5 g

5×10^{23} molecules of glucose ?

Molecular mass of glucose = 180

$$\text{Mass of glucose} = \frac{\text{Molecular mass} \times \text{number of particles}}{\text{Avogadro's number}}$$

$= (180 \times 5 \times 10^{23}) / 6.023 \times 10^{23}$
= 149.43 g

Calculation based on number of atoms/ molecules.

Calculate the number of molecules in 11.2 litre of CO_2 at S.T.P

$$\text{Number of moles of } CO_2 = \frac{\text{Volume at S.T.P}}{\text{Molar volume}}$$

= 11.2 / 22.4
= 0.5 mole

- Number of molecules of CO_2 = number of moles of CO_2 × Avogadro's number

$= 0.5 \times 6.023 \times 10^{23}$
$= 3.011 \times 10^{23}$ molecules of CO_2

Calculate the number of atoms present in 1 gram of gold (Atomic mass of Au = 198)

$$\text{Number of atoms of Au} = \frac{\text{Mass of Au} \times \text{Avogadro's number}}{\text{Atomic mass of Au}}$$

$$\text{Number of atoms of Au} = \frac{1}{198} \times 6.023 \times 10^{23}$$

Number of atoms of Au = 3.042×10^{21} g

Calculate the number of molecules in 54 gm of H_2O?

$$\text{Number of molecules} = \frac{(\text{Avogadro number} \times \text{Given mass})}{\text{Gram molecular mass}}$$

Number of molecules of water = $6.023 \times 10^{23} \times 54 / 18 = 18.069 \times 10^{23}$ molecules

Calculate the number of atoms of oxygen and carbon in 5 moles of CO_2.

1 mole of CO_2 contains 2 moles of oxygen
5 moles of CO_2 contain 10 moles of oxygen

Number of atoms of oxygen = Number of moles of oxygen × Avogadro's number

$= 10 \times 6.023 \times 10^{23}$
$= 6.023 \times 10^{24}$ atoms of Oxygen

- 1 mole of CO_2 contains 1 mole of carbon
- 5 moles of CO_2 contains 5 moles of carbon

No. of atoms of carbon = No. of moles of carbon × Avogadro's number

$= 5 \times 6.023 \times 10^{23}$
$= 3.011 \times 10^{24}$ atoms of Carbon

Calculation based on molar volume
Calculate the volume occupied by:

2.5 mole of CO_2 at S.T.P

$$\text{Number of moles of } CO_2 = \frac{\text{Given volume at S.T.P}}{\text{Molar volume at S.T.P}}$$

$$2.5 \text{ mole of } CO_2 = \frac{\text{Volume of CO2 at S.T.P}}{22.4}$$

Volume of CO_2 at S.T.P = 22.4 × 2.5 = 56 litres.
3.011×10^{23} of ammonia gas molecules

$$\text{Number of moles} = \frac{\text{Number of molecules}}{\text{Avogadro's number}}$$

$= 3.011 \times 10^{23} / 6.023 \times 10^{23}$
= 2 moles

Volume occupied by NH_3

= number of moles × molar volume
= 2 × 22.4
= 44.8 litres at S.T.P

14 g nitrogen gas
Number of moles = 14 / 28 = 0.5 mole

Volume occupied by N_2 at S.T.P

= no. of moles × molar volume
= 0.5 × 22.4
= 11.2 litres

Calculation based on % composition

Calculate % of S in H_2SO_4
Molar mass of H_2SO_4

= (1 × 2) + (32 × 1) + (16 × 4)
= 2 + 32 + 64
= 98 g

% of S in $H_2SO_4 = \frac{\text{Mass of sulphur}}{\text{Molar mass of } H_2SO_4} \times 100$

% of S in $H_2SO_4 = \frac{32}{98} \times 100$ = 32.65 %

8.PERIODIC CLASSIFICATION OF ELEMENTS

Introduction

- The eighteenth and nineteenth centuries witnessed a rapid developmentin chemistry in all spheres of scientific activities. By 1860, scientists had already discovered 60 elements and determined their atomic masses. They noticed that some elements had similar properties and hence arranged them into groups. During this period, several new elements were discovered. These elements were found to have different properties. It was realized that instead of studying the properties of all these elements individually, it would be more convenient to divide them into groups and periods in such a way that each group contained a certain number of elements (like an array of fruits and vegetables showing orderliness) with similar properties and periods showing a regular gradation. So, scientists made several attempts to arrange elements in a logical way.

MODERN PERIODIC LAW

- Mendeleev's periodic table had some discrepancies, which were difficultto overcome. For example, the atomic mass of argon (39.95 amu) is greater than that of potassium (39.10 amu), but argon comes before potassium in the periodic table. If elements were arranged solely according to increasing atomic mass, argon would appear in the position occupied by potassium in our modern periodic. No chemist would place argon, a gas with no tendency to react, in the same group as lithium and sodium, which are two highly reactive metals. This kind of discrepancies suggested that some fundamental property other than atomic mass must be the basis of periodicity. The fundamental property turned out to be the number of protons in an atom's nucleus, something that could not have been known by Mendeleev and his contemporaries.

- Henry Moseley, a British scientist in 1912, discovered a new property ofelements called atomic number, which provided a better basis for the periodic arrangement of the elements. It is a well-known fact that atomic number of an element is equal to the number of protons or the number of electrons present in the neutral atom of an element. The periodic law was, therefore, modified to frame a modern periodic law, which states that "The physical and chemical properties of the elements are the periodic functions of their atomic numbers".

MODERN PERIODIC TABLE

- With reference to the modern periodic law, the elements were arrangedin the increasing order of their atomic numbers to form the modern periodic table. The modern periodic table is a tabular arrangement of elements in rows and columns, highlighting the regular repetition of properties of the elements.

Features of Periods

- The **horizontal rows are called periods**. There are seven periods in the periodic table.
- **First period** (Atomic number 1 and 2): This is the shortest period. It contains only two elements (Hydrogen and Helium).
- **Second period** (Atomic number 3 to 10): This is a short period. It contains eight elements (Lithium to Neon).
- **Third period** (Atomic number 11 to 18): This is also a short period. It contains eight elements (Sodium to Argon).
- **Fourth period** (Atomic number 19 to 36): This is a long period. It contains eighteen elements (Potassium to Krypton). This includes 8 normal elements and 10 transition elements.
- **Fifth period** (Atomic number 37 to 54): This is also a long period. It contains 18 elements (Rubidium to Xenon). This includes 8 normal elements and 10 transition elements.
- **Sixth period** (Atomic number 55 to 86): This is the longest period. It contains 32 elements (Caesium to Radon). This includes 8 normal elements, 10 transition elements and 14 inner transition elements (Lanthanides).
- **Seventh period** (Atomic number 87 to 118): Like the sixth period, this period also accommodates 32 elements. Recently 4 elements have been included by IUPAC

Features of Groups

- The vertical columns in the periodic table starting from top to bottom are called groups. There are 18 groups in the periodic table.
- Based on the common characteristics of elements in each group, they can be grouped as various families.

Group Number	Family
1	

	Alkali Metals
2	Alkaline earth metals
3 to 12	Transition metals
13	Boron Family
14	Carbon Family
15	Nitrogen Family
16	Oxygen Family (or) Chalcogen family
17	Halogens
18	Noble gases

- The Lanthanides and Actinides, which form part of Group 3 are called inner transition elements.
- Except 'group 0', all the elements present in each group have the same number of electrons in their valence shell and thus have the same valency. For example, all the elements of group 1 have one electron in their valence shells (1s1). So, the valency of all the alkali metals is ‘1’.
- As the elements present in a group have identical valence shell electronic configurations, they possess similar chemical properties.
- The physical properties of the elements in a group such as melting point, boiling point and density vary gradually.
- The atoms of the 'group 0' elements have stable electronic configuration in their valence shells and hence they are unreactive.

PERIODIC TRENDS IN PROPERTIES

- The electronic configurations of elements help us to explain the periodicrecurrence of physical and chemical properties. Anything which repeats itself after a regular interval is called periodic and this behaviour is called periodicity. Some of the atomic properties of the elements are periodic.

- Properties such as atomic radius, ionic radius, ionisation energy,electronegativity, electron affinity, show a regular periodicity and hence they are called periodic properties. The main significance of the modern periodic table is that it gives a clear understanding of the general properties and trends within a group or a period to predict with considerable accuracy, the properties of any element, even though that element may be unfamiliar to us. Let us discuss the periodic trend of some of the properties.

Atomic Radius

- Atomic radius of an atom is defined as the distance between the centreofits nucleus and the outermost shell containing the valence electron. Direct measurement of the radius of an isolated atom is not possible. Except for noble gases, usually the atomic radius is referred to as covalent radius or metallic radius depending on the nature of the bonding between the concerned atoms. Atomic radius in metal atoms is known as metallic radius. It is defined as half the distance between the nuclei of adjacent metal atoms.

- In non-metallic elements, their atomic radius is known as Covalentradius. It is defined as half the distance between the nuclei of two covalently bonded atoms of the same element in a molecule. For example, let us consider H2 molecule. The distance between the two hydrogen nuclei of the molecule is 0.74 Å. So its covalent radius is 0.74/2 = 0.37 Å.

- When you look at the variation of the atomic radii in the periodic table,there are two distinct trends. Along the period, from left to right, the atomic radius of the elements decreases whereas along the groups, from the top to bottom, the atomic radius increases. The increase, down a group, is due to the increase in the valence shell number down the group. As the shell number increases, the distance between the valence shell and the nucleus increases. In contrast, when you observe along the period, the shell number remains the same but the number of protons (i.e. atomic number) increases. More and more positive charges impose a strong attraction over the electrons and thus the electron cloud shrinks towards the nucleus, which results in the decrease in the atomic size.

Ionic Radii

- It is defined as the distance from the centre of the nucleus of the ionupto the point where it exerts its influence on the electron cloud of the ion. You know that ions are formed when an atom lose or gain electrons. When a neutral atom loses an electron, it becomes a positively charged ion called cation, whereas the gain of an electron by a neutral atom forms a negatively charged ion called anion. The size of the ions is important to determine their behaviours in solutions and the structure of ionic solids. The size of a cation is always smaller than its corresponding neutral atom. But, the anion is larger than its neutral atom.

- For instance, lithium and sodium lose the single electron from theiroutermost energy level to form cations. The ions so formed are smaller because the remaining electrons are at a inner cells and attracted more strongly by the nucleus. Fluorine and chlorine become negative ions by gaining an electron. When electrons are added, the charge on the nucleus is not great enough to hold the increased number of electrons as closely as it holds the electrons in the neutral atom. So, as seen in atomic radius, ionic radii also decrease along the period from left to right and increase down the group.

Ionisation Energy

- Ionisation energy is the minimum energy required to remove anelectron from a gaseous atom in its ground state to form a cation. It is otherwise called ionisation enthalpy. It is measured in kJ/mol. Higher the ionisation energy, it is more difficult to remove the electron.

- As the atomic size decreases from left to right in a period, more energyis required to remove the electrons. So, the ionisation energy increases along the period. But, down the group, the atomic size increases and hence the valence electrons are loosely bound. They require relatively less energy for the removal. Thus, ionisation energy decreases down the group in the periodic table.

As the positive charge increases the size of the cation decreases As thenegative charge increases the size of the anion increases

Electron Affinity

- Electron affinity is the amount of energy released when a gaseous atomgains an electron to form its anion. It is also measured in kJ/mol and represented by the following equation:

  $A(g) + e^- \rightarrow A^{-(g)}$ + Energy
  $Cl(g) + e- \rightarrow Cl^-_{(g)}$+ energy
- Like ionisation energy, electron affinity also increases from left to rightin a period and decreases from top to bottom in a group.

Electronegativity

- Electronegativity of an element is the measure of the tendency of itsatom to attract the shared pair of electrons towards itself in a covalent bond. Let us consider HCl molecule. Both the hydrogen and chlorine atoms share one electron each to form the covalent bond between them. chlorine atom has a higher electronegativity and hence it pulls the shared electrons towards itself more strongly than hydrogen. Thus, when the bond breaks, the bonding electrons are left with chlorine forming H+ and Cl– ions.

- Electronegativity is based on various experimental data such as bondenergy, ionization potential, electron affinity, etc.

- Pauling scale is the widely used scale to determine the electronegativity,which in turn predicts the nature of bonding (ionic or covalent) between the atoms in a molecule.

  Electronegativity of some of the elements.
  F = 4.0, Cl = 3.0, Br = 2.8, I = 2.5, H = 2.1, Na = 1

- If the difference in electronegativity between two elements is 1.7, thebond has 50% ionic character and 50% covalent character.

- If the difference is less than 1.7, the bond is considered to be covalent.

- If the difference is greater than 1.7, the bond is considered to be ionic.

- Along the period, from left to right in the periodic table, theelectronegativity increases because of the increase in the nuclear charge which in turn attracts the electrons more strongly. On moving down a group, the electronegativity of the elements decreases because of the increased number of energy levels.

Periodic Property	**In periods**	**In Groups**
Atomic radius	Decreases	Increases
Ionic radius	Decreases	Increases
Ionisation energy	Increases	Decreases
Electron affinity	Increases	Decreases
Electronegativity	Increases	Decreases

METALLURGY

- Human life is associated with various metals. We use metals in our dayto day activities. It is the utmost need to have some metals like sodium, potassium, calcium, iron, etc. in the human body. Deficiency of these metals affects the metabolic activities thereby causing diseases. So, metals play a vital role in our life. In this section, let us discuss how metals are obtained from various sources by the process of metallurgy.

- Metallurgy is a science of extracting metals from their ores andmodifying the metals into alloys for various uses, based on their physical and chemical properties and their structural arrangement of atoms. A metallurgical process involve three main steps as follows:

- **Concentration or Separation of the ore:** It is the process of removal of impuries from the ore.

- **Production of the metal:** It is the convertion of the ore into metal.

- **Refining of the metal:** It is the process of purification of the metal

Terminology in metallurgy

- **Minerals:** A mineral may be a single compound or a complex mixture ofvarious compounds of metals found in the Earth.

- **Ore:** The mineral from which a metal can be readily and economicallyextracted on a large scale is said to be an ore.

- **For example:** Clay ($Al_2O_3 . 2 SiO_2 . 2 H_2O$) and bauxite ($Al_2O_3 . 2 H_2O$) are the two minerals of aluminium, but aluminium can be profitably extracted only from bauxite. Hence, bauxite is an ore of aluminium and clay is its mineral.

- **Mining:** The process of extracting the ores from the Earth's crust iscalledmining.

- **Gangue or Matrix:** The rocky impurity associated with an ore is calledgangue or matrix.

- **Flux:** It is the substance added to the ore to reduce the fusiontemperature and to remove the impurities. E.g. Calcium oxide (basic), Silica (acidic). If the gangue is acidic, then basic flux is added and vice versa.
- **Slag:** It is the fusible product formed when a flux reacts with a gangueduring the extraction of metals.

  Flux + Gangue → Slag

- **Smelting:** Smelting is the process of reducing the roasted metallic oxidefrom the metal in its molten condition. In this process, impurities are removed as slag by the addition of flux.

Types of separation or concentration of an ore

- There are four major types of separation of ores based on the nature ofthe ore. The different kinds of ores of metals.

Hydraulic (Gravity Separation) method

- **Principle:** The difference in the densities or specific gravities of the oreand the gangue is the main principle behind this method. Oxide ores are purified by this method. e.g., Haematite$_{Fe_2O_3}$ the ore of iron.

- **Method:** The ore is poured over a sloping, vibrating corrugated tablewith grooves and a jet of water is allowed to flow over it. The denser ore particles settle down in the grooves and lighter gangue particles are washed down by water.

Magnetic separation method

- **Principle:** The magnetic properties of the ores form the basis ofseparation. When either the ore or the gangue is magnetic, this method is employed. e.g., Tinstone SnO_2, the ore of tin.

- **Method:** The crushed ore is placed over a conveyer belt which rotatesaround two metal wheels, one of which is magnetic. The magnetic particles are attracted to the magnetic wheel and fall separately apart from the non-magnetic particles.

Froth floatation

- **Principle:** This process depends on the preferential wettability of the orewith oil (pine oil) and the gangue particles by water. Lighter ores, such as sulphide ores, are concentrated by this method. e.g., Zinc blende (ZnS).

- **Method:** The crushed ore is taken in a large tank containing oil andwater and agitated with a current of compressed air. The ore is wetted by the oil and gets separated from the gangue in the form of froth. Since the ore is lighter, it comes on the surface with the froth and the impurities are left behind. e.g., Zinc blende (ZnS).

Chemical method or Leaching

This method is employed when the ore is in a very pure form.

Oxide Ores	**Carbonate Ores**	**Halide Ores**	**Sulphide ores**
Bauxite ($Al_2O_{3.}2H_2O$)	Marble ($CaCO_3$)	Cryolite (Na_3AIF_6)	Galena (pbs)
Cuprite (Cu_2O)	Magnesite ($MgCO_3$)	Fluorspar (CaF_2)	Iron Pyrite (FeS_2)
Haematite (Fe_2O_3)	Siderite ($FeCO_3$)	Rock Salt (NaCl)	Zinc Blende (Zns)

- The ore is treated with a suitable reagent such that the ore is soluble in itbut the impurities are not. The impurities are removed by filtration. The solution of the ore, ie., the filtrate is treated with a suitable reagent which precipitates the ore. E.g. Bauxite $Al_2O_3.2H_2O$, the ore of aluminium.

OCCURRENCE OF ORES IN TAMIL NADU

- **Lime stone:** Coimbatore, Cuddalore, Dindugul
- **Gypsum:**Tiruchi and Coimbatore Distiricts
- **Titanium minerals:**Kanyakumari, Tirunelveli and Tuticorin.

- **Chromite:** Coimbatore and Salem district.
- **Magnetite:**.Dharmapuri, Erode, Salem, Thiruvannamalai.
- **Tungsten:** Madurai and Dindugal.

PROPERTIES OF METALS

Physical properties

- **Physical state:** All metals are solids at room temperature except mercury and gallium.

- **Lustre:** Metals possess a high lustre (called metallic lustre).

- **Hardness:** Most of the metals are hard and strong (exceptions: sodium and potassium can be cut with a knife)

- **Melting point and Boiling point:** Usually, metals possess high melting and boiling points and vaporize only at high temperatures (exceptions: gallium, mercury, sodium and potassium).

- **Density:** Metals have a high density (exceptions: sodium and potassium are less dense than water).

- **Ductility:** Metals are usually ductile. In other words, they can be drawn into thin wires without breaking.

- **Malleability:** Metals are usually malleable, i.e, they can be beaten into thin sheets without cracking (except zinc and mercury).

- **Conduction of heat and electricity:** Metals are good conductors of heat and electricity; silver and copper excel in this property (exception: tungsten)

- **Solubility:** Usually, metals do not dissolve in liquid solvents.

Chemical Properties

- **Valence electrons:** Atoms of metals usually have 1,2 or 3 electrons in their outermost shell.

- **Formation of ions:** Metals form Positive ions by the loss of electrons and hence they are electro positive.

- **Discharge of ions:** Metals are discharged at the cathode during the electrolysis of their compounds.

- **Atomicity:** Molecules of metals in their vapour state are usually monoatomic.

- **Nature of oxides:** Oxides of metals are usually basic.

EXTRACTIVE METALLURGY OF ALUMINIUM

- ❖ Aluminium is the metal found most abundantly in the Earth's crust.Since it is a reactive metal, it occurs in the combined state. The important ores of aluminium are as follows.

Ores of Aluminium	**Formula**
Bauxite	$Al_2O_3.2H_2O$
Cryolite	Na_3AlF_6
Corundum	Al_2O_3

- Bauxite is the chief ore of aluminium. The extraction of aluminium from bauxite involves two steps:

Conversion of bauxite into alumina – Baeyer's Process

- The conversion of Bauxite into Alumina involves the following steps:

- Bauxite ore is finely ground and heated under pressure with a solution of concentrated caustic soda solution at 150° C to obtain sodium meta aluminate.
- On diluting sodium meta aluminate with water, a precipitate of aluminium hydroxide is formed.
- The precipitate is filtered, washed, dried and ignited at 1000°C to get alumina.

$$2Al(OH)_3 \xrightarrow{1000^\circ c} Al_2O_3 + 3H_2O$$

Electrolytic reduction of alumina – Hall's Process

- Aluminium is produced by the electrolytic reduction of fused alumina (Al_2O_3) in the electrolytic cell

- **Cathode**: Iron tank linked with graphite

- **Anode**: A bunch of graphite rods suspended in molten electrolyte.

- **Electrolyte**: Pure alumina+ molten cryolite + fluorspar (fluorspar lowers the fusion temperature of electrolyte)

Temperature: 900 - 950 °C
Voltage used: 5-6 V

Overall reaction: 2 Al2O3 → 4 Al +3 O2↑

- Aluminium is deposited at the cathode and oxygen gas is liberated at the anode. Oxygen combines with graphite to form CO_2.

Physical Properties of Aluminium

- It is a silvery white metal
- It has low density (2.7) and it is light
- It is malleable and ductile
- It is a good conductor of heat and electricity.
- Its melting point is 660 °C.
- It can be polished to produce a shiny attractive appearance.

Chemical Properties of Aluminium

- **Reaction with air:** It is not affected by dry air. On heating at 800 °C, aluminium burns very brightly forming it's oxide and nitride.

  $4 Al + 3 O_2 \rightarrow 2 Al_2O_3$(Aluminium oxide)
  $2 Al + N_2 \rightarrow 2 AlN$ (Aluminium nitride)

- **Reaction with water:** Water does not react with aluminium due to the layer of oxide on it. When steam is passed over red hot aluminium, hydrogen is produced.

  $2 Al + 3 H_2O \rightarrow Al_2O_3 + 3 H_2\uparrow$

- **Reaction with alkalis:** It reacts with strong caustic alkalis forming aluminates.

  $2 Al + 2 NaOH +2 H_2O \rightarrow 2 NaAlO_2 + 3 H_2\uparrow$

- **Reaction with acids:** With dilute and con. HCl it liberates H2 gas.

  $2 Al + 6 HCl \rightarrow 2 AlCl_3 + 3 H_2\uparrow$

- Aluminium liberates hydrogen on reaction with dilute sulphuric acid. Sulphur dioxide is liberated with hot concentrated sulphuric acid

  $2 Al+ 3 H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3 H_2$
  $2 Al + 6 H_2SO_4 \rightarrow Al_2(SO_4)_3 + 6 H_2O + 3 SO_2 \uparrow$

As reducing agent: Aluminium is a powerful reducing agent. When a mixture of aluminium powder and iron oxide is ignited, the latter is reduced to metal. This process is known as **aluminothermic process**.

$Fe_2O_3 + 2\ Al \rightarrow 2\ Fe + Al_2O_3 + \text{Heat}.$

Uses

Aluminium is used in

- household utensils
- electrical cable industry
- making aeroplanes and other industrial mechine parts

EXTRACTIVE METALLURGY OF COPPER

Occurrence:

- It was named as cuprum by the Romans because they got it from the Island of Cyprus. Copper is found in the native state as well as combined state.

Ores of copper	**Formula**
Copper pyrites	CuFeS2
Cuprite or ruby copper	Cu2O
Copper glance	Cu2S

- The chief ore of copper is copper pyrite. It yields nearly 76% of the world production of copper. Extraction of copper from copper pyrites involves the following steps

- **Concentration of ore:** The ore is crushed and the concentrated by froth floatation process.

- **Roasting:** The concentrated ore is roasted in excess of air. During the process of roasting, the moisture and volatile impurities are removed. Sulphur, phosphorus, arsenic and antimony are removed as oxides. Copper pyrite is partly converted into sulphides of copper and iron.

  $2\ CuFeS_2 + O_2 \rightarrow Cu2S + 2\ FeS + SO_2 \uparrow$

- **Smelting:** The roasted ore is mixed with powdered coke and sand and is heated in a blast furnace to obtain matte (Cu2S + FeS) and slag. The slag is removed as waste.

- **Bessemerisation:** The molten matte is transferred to Bessemer converter in order to obtain blister copper. Ferrous sulphide from matte is oxidized to ferrous oxide, which is removed as slag using silica.

$2\ FeS + 3\ O_2\ 2\ FeO + 2\ SO_2 \uparrow$
$FeO + SiO_2\ FeSiO_3$ (slag)
$2\ Cu_2S + 3O_2\ 2\ Cu_2O + 2\ SO_2 \uparrow$
$2\ Cu_2O + Cu_2S\ 6\ Cu + SO_2\uparrow$

- **Refining:** Blister copper contains 98% of pure copper and 2% of impurities and is purified by **electrolytic refining**. This method is used to get metal of a high degree of purity. For electrolytic refining of copper, we use:

- **Cathode**: A thin plate of pure copper metal.
- **Anode**: A block of impure copper metal.
- **Electrolyte**: Copper sulphate solution acidified with sulphuric acid

- When electric current is passed through the electrolytic solution, pure copper gets deposited at the cathode and the impurities settle at the bottom of the anode in the form of sludge called anode mud.

Physical Properties of Copper

- Copper is a reddish brown metal, with high lustre, high density and high melting point (1356°C).

Chemical Properties of Copper

- **Action of Air and Moisture:** Copper gets covered with a green layer of basic copper carbonate in the presence of CO2 and moisture.

  $2\ Cu + O_2 + CO_2 + H_2O \rightarrow CuCO_3.Cu(OH)_2$

- **Action of Heat**: On heating at different temperatures in the presence of oxygen, copper forms two types of oxides CuO, Cu2O.

  $2\ Cu + O_2 \xrightarrow{\text{below 1370K}} 2\ CuO$
  $4\ Cu + O2 \xrightarrow{\text{above 1370K}} 2\ Cu_2O$

Action of Acids:
With dilute HCl and dilute H2SO4:

- Dilute acids such as HCl and H2SO4 have no action on these metals in the absence of air. Copper dissolves in these acids in the presence of air.

  $2\ Cu + 4\ HCl + O_2$ (air) $2\ CuCl_2 + 2\ H_2O$

With dil. HNO3:

- Copper reacts with dil. HNO3 with the liberation of Nitric Oxide gas.

  $3\,Cu + 8\,HNO_3 \rightarrow Cu(NO_3)_2 + 2\,NO \uparrow + 2H_2O$
  $Cu + 2\,H_2SO_4 \rightarrow CuSO_4 + SO_2 \uparrow + 2\,H_2O$

Action of Chlorine:

- Chlorine reacts with copper, resulting in the formation of copper(II) chloride.

  Cu + Cl2 CuCl2

Action of Alkalis:

- Copper is not attacked by alkalis.

Uses of Copper:

- It is extensively used in manufacturing electric cables and other electric appliances.
- It is used for making utensils, containers, calorimeters and coins,
- t is used in electroplating.
- It is alloyed with gold and silver for making coins and jewels

EXTRACTIVE METALLURGY OF IRON
Occurrence:

- Iron is the second most abundant metal available next to aluminium. It occurs in nature as oxides, sulphides and carbonates. The ores of iron are as follows:

Ores of iron	Formula
Haematite	Fe_2O_3
Magnetite	Fe_3O_4
Iron pyrite	FeS_2

- Iron is chiefly extracted from haematite ore (Fe_2O_3)

Concentration by Gravity Separation:

- The powdered ore is washed with a steam of water. As a result, the lighter sand particles and other impurities are washed away and the heavier ore particles settle down.

Roasting and Calcination:

- The concentrated ore is strongly heated in a limited supply of air in a reverberatory furnace. As a result, moisture is driven out and sulphur, arsenic and phosphorus impurities are oxidized off.

Smelting (in a Blast Furnace):

- The charge consisting of roasted ore, coke and limestone in the ratio 8:4:1 is smelted in a blast furnace by introducing it through the cup and cone arrangement at the top. There are three important regions in the furnace.

- The Lower Region (Combustion Zone)- The temperature is at 1500°C. In this region, coke burns with oxygen to form CO_2 when the charge comes in contact with a hot blast of air.

$$C + O_2 \xrightarrow[\Delta]{1500°C} CO_2 + Heat$$

- It is an exothermic reaction since heat is liberated.

- The Middle Region (Fusion Zone) – The temperature prevails at 1000°C. In this region, CO2 is reduced to CO.

$$CO2 + C \xrightarrow[\Delta]{1000°C} CO_2 + Heat$$

- Limestone decomposes to calcium oxide and CO_2

$$CaCO3 \xrightarrow[\Delta]{} CO2 - Heat$$

- These two reactions are endothermic due to absorption of heat. Calcium oxide combines with silica to form calcium silicate slag.

$$CaO + SiO_2 \rightarrow CaSiO_3$$

- The Upper Region (Reduction Zone)- The temperature prevails at 400°C . In this region carbon monoxide reduces ferric oxide to form a fairly pure spongy iron.

$$Fe_2O_3 + 3CO \xrightarrow{400°C} 2Fe + 3CO_2$$

- The molten iron is collected at the bottom of the furnace after removing the slag.

- The iron thus formed is called pig iron. It is remelted and cast into different moulds. This iron is called cast iron.

Physical properties:

- It is a lustrous metal, greyish white in colour.
- It has high tensility, malleability and ductility.
- It can be magnetized

Chemical properties:

- **Reaction with air or oxygen:** Only on heating in air, iron forms magnetic oxide.

  $3\ Fe + 2\ O_2 \rightarrow Fe_3O_4$ (black)

- **Reaction with moist air:** When iron is exposed to moist air, it forms a layer of brown hydrated ferric oxide on its surface. This compound is known as rust and the phenomenon of formation of rust is known as **rusting**.

  $4\ Fe + 3\ O_2 + x\ H_2O \rightarrow 2\ Fe_2O_3 .xH_2O$(rust)

- **Reaction with steam:** When steam is passed over red hot iron, magnetic oxide is formed.

  $3\ Fe + 4\ H_2O$ (steam) $\rightarrow Fe_3O_4 + 4\ H_2 \uparrow$

- **Reaction with chlorine:** Iron combines with chlorine to form ferric chloride.

  $2Fe + 3Cl_2 \rightarrow 2FeCl_3$ (ferric chloride)

- **Reaction with acids:** With dilute HCl and dilute H_2SO_4 it liberates H_2 gas.

  $Fe + 2HCl \rightarrow FeCl_2 + H_2 \uparrow$
  $Fe + H_2SO_4 \rightarrow FeSO_4 + H_2 \uparrow$

- With dilute HNO3 in cold condition it gives ferrous nitrate.

  $4\ Fe + 10\ HNO_3 \rightarrow 4\ Fe(NO_3)_2 + NH_4NO_3 + 3\ H_2O$

- With con. H2SO4 it forms ferric sulphate.

  $2\ Fe + 6\ H_2SO_4 \rightarrow Fe_2(SO_4)_3 + 3\ SO_2 + 6\ H_2O$

- When iron is dipped in con. HNO3 it becomes chemically passive or inert due to the formation of a layer of iron oxide (Fe3O4) on its surface.

Uses of iron

- **Pig iron** (Iron with 2-4.5% of carbon)**:** It is used in making pipes, stoves, radiators, railings, manhole covers and drain pipes.

- **Steel** (Iron with < 0.25% of carbon)**:** It is used in the construction of buildings, machinery, transmission cables and T.V towers and in making alloys.

- **Wrought iron** (Iron with 0.25-2% of wraught carbon)**:** It is used in making springs, anchors and electromagnets.

ALLOYS

- An alloy is a homogeneous mixture of two or more metals or of one or more metals with certain non-metallic elements.

- The properties of alloys are often different from those of its components. Pure gold is too soft to be used. The addition of small percentage of copper enhances its strength and utility.

Amalgam

- An amalgam is an alloy of mercury with another metal. These alloys are formed through metallic bonding with the electrostatic force of attraction between the electrons and the positively charged metal ions. Silver tin amalgam is used for dental filling.

Reasons for alloying:

- To modify appearance and colour
- To modify chemical activity.
- To lower the melting point.
- To increase hardness and tensile strength.
- To increase resistance to electricity.

Method of making alloys

- By fusing the metals together. E.g. Brass is made by melting zinc and copper.

- By compressing finely divided metals. E.g. Wood metal: an alloy of lead, tin, bismuth and cadmium powder is a fusible alloy.

Alloys as solid solutions:

- Alloys can be considered solid solutions in which the metal with high concentration is solvent and other metals are solute.

- For example, brass is a solid solution of zinc (solute) in copper (solvent).

Types of Alloys

- Based on the presence or absence of Iron, alloys can be classified into:

- Ferrous alloys: Contain Iron as a major component. A few examples of ferrous alloys are Stainless Steel, Nickel Steel etc.
- Non-ferrous alloys: These alloys do not contain Iron as a major component. For example, Aluminium alloy, Copper alloy etc.

Copper Alloys (Non- ferrous)

Alloys	Uses
Brass (Cu, Zn)	Electrical fittings, medal, decorative items, hardware
Bronze (Cu, Sn)	Statues, coins, bells, gongs

Aluminium Alloys (Non- ferrous)

Alloys	Uses
Duralumin (Al, Mg, Mn, Cu)	Aircrafts, tools, pressure cookers
Magnalium (Al, Mg)	Aircraft, scientific instruments

Iron Alloys(Ferrous)

Alloys	Uses
Stainless steel (Fe,C, Ni,Cr)	Utensils, cutlery, automobile parts
Nickel steel (Fe,C,Ni)	Cables , aircraftparts, propeller

CORROSION

- It is the gradual destruction of metals by chemical or electrochemical reaction with the environment. It is a natural process which converts a metal into its oxide, hydroxide or sulphide so that it loses its metallic characteristics.

- Rust is chemically known as hydrated ferric oxide (it is formulated as $Fe_2O_3.xH_2O$). Rusting results in the formation of scaling reddish brown hydrated ferric oxide on the surface of iron and iron containing materials.

Types of Corrosion
Dry Corrosion or Chemical Corrosion:

- The corrosive action in the absence of moisture is called dry corrosion. It is the process of a chemical attack on a metal by a corrosive liquids or gases such as O_2, N_2, SO_2, H_2S etc. It occurs at high temperature. Of all the gases mentioned above O_2 is the most reactive gas to impart the chemical attack.

Wet Corrosion or Electrochemical Corrosion:

- The corrosive action in the presence of moisture is called wet corrosion. It occurs as a result of electrochemical reaction of metal with water or aqueous solution of salt or acids or bases.

Methods of preventing corrosion

1. **Alloying:** The metals can be alloyed to prevent the process of corrosion. E.g: Stainless Steel
2. **Surface Coating:** It involves application of a protective coating over the metal. It is of the following types:

Galvanization:

- It is the process of coating zinc on iron sheets by using electric current.

Electroplating:

- It is a method of coating one metal over another metal by passing electric current.

Anodizing:

- It is an electrochemical process that converts the metal surface into a decorative, durable and corrosion resistant. Aluminium is widely used for anodizing process.

Cathodic Protection:

- It is the method of controlling corrosion of a metal surface protected is coated with the metal which is easily corrodible. The easily corrodible metal is called Sacrificial metal to act as anode ensuring cathodic protection

PAMBAN BRIDGE

It is a railway bridge which connects the town of Rameshwaram on Pamban Island to mainland India. Opened on 1914, it was India's first sea bridge in India until the opening of the BandraWorli Sea Link in 2010. We can control the corrosion and renovation of historical pamban bridge by a periodical protective coating which will be the strong example for applied chemistry to uphold our history.

11thvol 1
Unit – 3
Periodic Classification of Elements

"An awareness of the periodic table is essential to anyone who wishes to disentangle the word and see how it is built up from the fundamental building blocks of the chemistry, the chemical elements "

- Glenn T. Seaborg

Glenn Theodore Seaborg received Nobel Prize in 1951 in chemistry for the discoveries of trans-uranium elements. He was the co-discoverer of plutonium and other trans uranium elements. He along with his colleagues has discovered over a hundred isotopes of other elements. He demonstrated that actinide elements are analogues to rare earth series of lanthanide elements.

Introduction

There are millions of chemicalcompounds existing in nature withdifferent compositions and properties,formed from less than 100 naturallyoccurring elements.

The discovery of elements is linkedwith human civilization. In stone age manhas used some metals to suit his needswithout knowing that they are elements.Soon he learnt to extract elements fromores and fashion them into his daily life.Over the years, more and more elementswere discovered. In 1789, Lavoisier fromFrance, published the first list of chemicalelements containing 23 elements afterseveral experimental investigations.

Antoine Lavoisier classified thesubstances into four groups of elementsnamely acid-making elements, gas-likeelements, metallic elements and earthyelements

Lavoisier table

acid-making elements	gas-like elements
Sulphur	Light
Phosphorus	Caloric (heat)
Charcoal (carbon)	Oxygen
	Azote (nitrogen)
	Hydrogen

metallic elements	earthy elements
cobalt, mercury, tin	lime (calcium oxide)
copper, nickel, iron	magnesia (magnesium oxide)

gold, lead, silver, zinc	barytes (barium sulphate)
manganese, tungten	argilla (aluminium oxide)
platina (platinum)	silex (silicon dioxide)

Classification of Elements

During the 19th century, scientists have isolated several elements and the list of known elements increased. Currently, we have 118 known elements. Out of 118 elements, 92 elements with atomic numbers 1 to 92 are found in nature. Scientists have found out there are some similarities in properties among certain elements. This observation has led to the idea of classification of elements based on their properties. In fact, classification will be beneficial for the effective utilization of these elements. Several attempts were made to classify the elements. However, classification based on the atomic weights led to the construction of a proper form of periodic table.

In 1817, J. W. Döbereiner classified some elements such as chlorine, bromine and iodine with similar chemical properties into the group of three elements called as triads. In triads, the atomic weight of the middle element nearly equal to the arithmetic mean of the atomic weights of the remaining two elements. However, only a limited number of elements can be grouped as triads.

Dobereiner Triads

S. No.	Elements in the Triad	Atomic weight of middle element	Average atomic weight of the remaining Elements
1.	Li, Na, K	23	$\frac{7+39}{2} = 23$
2.	Cl, Br, I	80	$\frac{35.5+127}{2} = 81.25$
3.	Ca, Sr, Ba	88	$\frac{40+137}{2} = 88.5$

This concept can not be extended to some triads which have nearly same atomic masses such as [Fe, Co, Ni], [Ru, Rh, Pd] and [Os, Ir, Pt].

In 1862, A. E. B. de Chancourtois reported a correlation between the properties of the elements and their atomic weights. He said ‘the properties of bodies are the properties of numbers’. He intended the term numbers to mean the value of atomic weights. He designed a helix by tracing at an angle 45˚ to the vertical axis of a cylinder with circumference of 16 units. He arranged the elements in the increasing atomic weights along the helix on the surface of this cylinder. One complete turn of a helix corresponds to an atomic weight increase of 16. Elements which lie on the 16 equidistant vertical lines drawn on the surface of cylinder shows similar properties. This was the first reasonable attempt towards the creation of periodic table. However, it did not attract much attention.

In 1864, J. Newland made an attempt to classify the elements and proposed the law of octaves. On arranging the elements in the increasing order of atomic weights, he observed that the properties of every eighth element are similar to the properties of the first element. This law holds good for lighter elements up to calcium.

Newlands' Octaves

^{7}Li	^{9}Be	^{11}B	^{12}C	^{14}N	^{16}O	^{19}F
^{23}Na	^{24}Mg	^{27}Al	^{29}Si	^{31}P	^{32}S	$^{35.5}Cl$
^{39}K	^{40}Ca					

Mendeleev's Classification

In 1868, Lothar Meyer had developed a table of the elements that closely resembles the modern periodic table. He plotted the physical properties such as atomic volume, melting point and boiling point against atomic weight and observed a periodical pattern.

During same period Dmitri Mendeleev independently proposed that "the properties of the elements are the periodic functions of their atomic weights" and this is called periodic law. Mendeleev listed the 70 known elements at that time in several vertical columns in order of increasing atomic weights. Thus, Mendeleev constructed the first periodic table based on the periodic law.

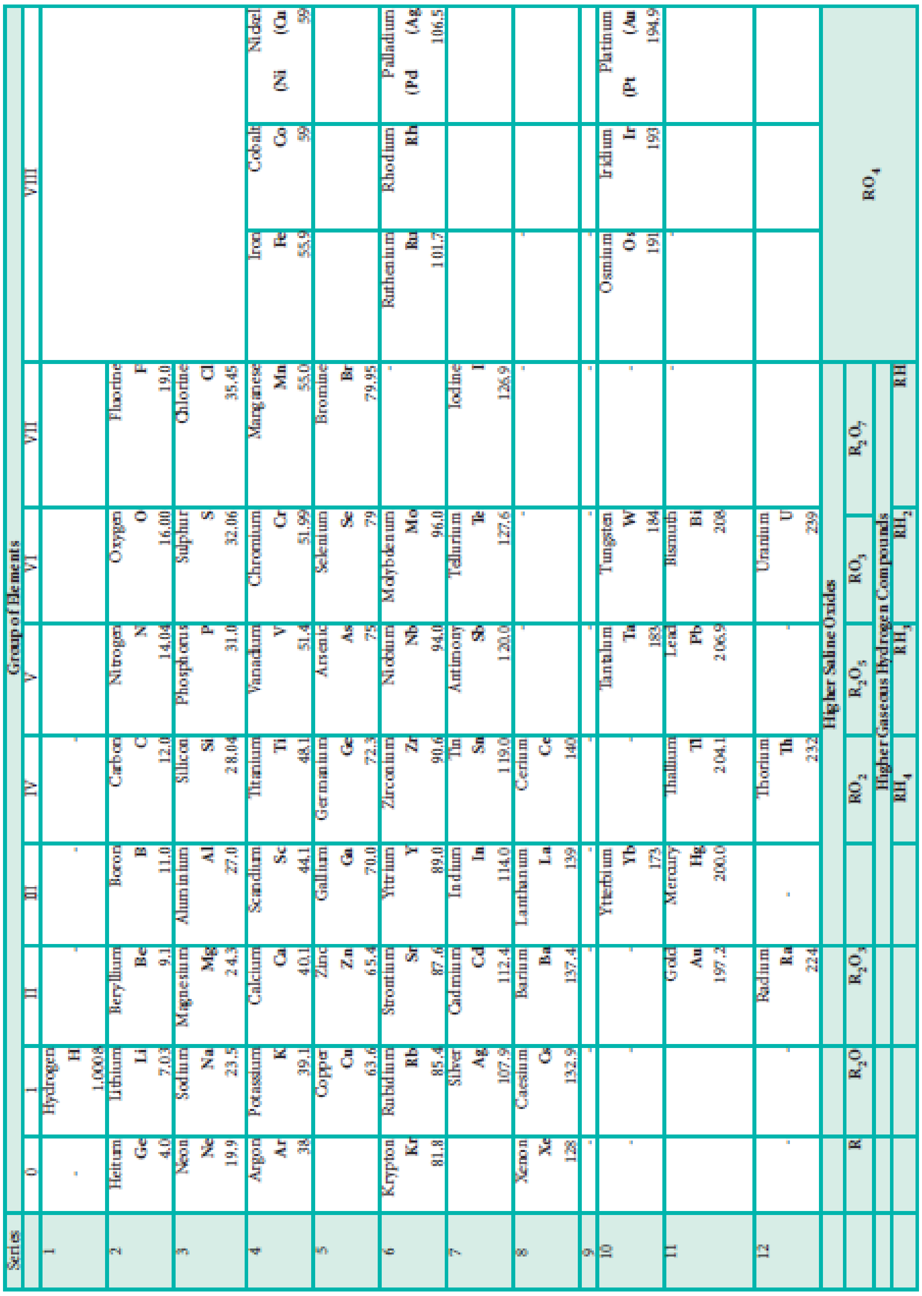
As shown in the periodic table, he left some blank spaces since there were no known elements with the appropriate properties at that time. He and others predicted the physical and chemical properties of the missing elements. Eventually these missing elements were discovered and found to have the predicted properties. For example, Gallium (Ga) of group III and germanium (Ge) of group IV were unknown at that time. But Mendeleev predicted their existence and properties. He referred the predicted elements as eka-aluminium and eka-silicon. After discovery of the actual elements, their properties were found to match closely to those predicted by Mendeleev.

Properties predicted for Eka-aluminium and Eka-silicon

S. NO.	Property	Ekaaluminium (Predicted)	Gallium (Observed)	Ekasilicon (Predicted)	Germanium (Observed)
1.	Atomic weight	68	70	72	72.59
2.	Density	5.9	5.94	5.5	5.35

	(g/cm3)				
3.	Melting point	low	29.78^O^C	High	947^O^C
4.	Formula of oxide	E_2O_3	Ga_2O_3	EO_2	GeO_2
5.	Formula of chloride	ECl_3	$GaCl_3$	ECl_4	$GeCl_4$

Mendeleev's periodic table



Series	Group of Elements										
	0	I	II	III	IV	V	VI	VII	VIII		
1	-	Hydrogen H 1.0008	-	-	-						
2	Heitum Ge 4.0	Lithium Li 7.03	Beryllium Be 9.1	Boron B 11.0	Carbon C 12.0	Nitrogen N 14.04	Oxygen O 16.00	Fluorine F 19.0			
3	Neon Ne 19.9	Sodium Na 23.5	Magnesium Mg 24.3	Aluminium Al 27.0	Silicon Si 28.04	Phosphorus P 31.0	Sulphur S 32.06	Chlorine Cl 35.45			
4	Argon Ar 38	Potassium K 39.1	Calcium Ca 40.1	Scandium Sc 44.1	Titanium Ti 48.1	Vanadium V 51.4	Chromium Cr 51.99	Manganese Mn 55.0	Iron Fe 55.9	Cobalt Co 59	Nickel (Ni (Cu 59
5		Copper Cu 63.6	Zinc Zn 65.4	Gallium Ga 70.0	Germanium Ge 72.3	Arsenic As 75	Selenium Se 79	Bromine Br 79.95			
6	Krypton Kr 81.8	Rubidium Rb 85.4	Strontium Sr 87.6	Yttrium Y 89.0	Zirconium Zr 90.6	Niobium Nb 94.0	Molybdenum Mo 96.0	-	Ruthenium Ru 101.7	Rhodium Rh	Palladium (Pd (Ag 106.5
7		Silver Ag 107.9	Cadmium Cd 112.4	Indium In 114.0	Tin Sn 119.0	Antimony Sb 120.0	Tellurium Te 127.6	Iodine I 126.9			
8	Xenon Xe 128	Caesium Cs 132.9	Barium Ba 137.4	Lanthanum La 139	Cerium Ce 140	-	-	-	-		
9	-	-	-	-	-	-	-	-	-		
10	-	-	-	Ytterbium Yb 173	-	Tantalum Ta 183	Tungsten W 184	-	Osmium Os 191	Iridium Ir 193	Platinum (Pt (Au 194.9
11			Gold Au 197.2	Mercury Hg 200.0	Thallium Tl 204.1	Lead Pb 206.9	Bismuth Bi 208	-	-		
12	-	-	Radium Ra 224	-	Thorium Th 232	-	Uranium U 239				
				Higher Saline Oxides					RO_4		
	R	R_2O	R_2O_3		RO_2	R_2O_5	RO_3	R_2O_7			
				Higher Gaseous Hydrogen Compounds							
					RH_4	RH_3	RH_2	RH			

Anomalies of Mendeleev's Periodic Table

Some elements with similar properties were placed in different groups and those with dissimilar properties were placed in same group.

Example: Tellurium (127.6) was placed in VI group but Iodine (127.0) was placed in VII group.

Similarly elements with higher atomic weights were placed before lower atomic weights based on their properties in contradiction to his periodic law. Example 59Co27 was placed before $^{58.7}Ni_{28}$

Moseley's Work and Modern Periodic Law

In 1913, Henry Moseley studied the characteristic X-rays spectra of several elements by bombarding them with high energy electrons and observed a linear correlation between atomic number and the frequency of X-rays emitted which is given by the following expression.

$$\upsilon = a(Z - b)$$

Where, υ is the frequency of the X-rays emitted by the element with atomic number 'Z'; a and b are constants and have same values for all the elements.

The plot of $\sqrt{\upsilon}$ against Z gives a straight line. Using this relationship, we can determine the atomic number of an unknown (new) element from the frequency of X-ray emitted.

Based on his work, the modern periodic law was developed which states that, "the physical and chemical properties of the elements are periodic functions of their atomic numbers." Based on this law, the elements were arranged in order of their increasing atomic numbers. This mode of arrangement reveals an important truth that the elements with similar properties recur after regular intervals. The repetition of physical and chemical properties at regular intervals is called periodicity.

Modern Periodic Table

The physical and chemical properties of the elements are correlated to the arrangement of electrons in their outermost shell (valence shell). Different elements having similar outer shell electronic configuration possess similar properties. For example, elements having one electron in their valence shell s-orbital possess similar physical and chemical properties. These elements are grouped together in the modern periodic table as first group elements.

Electronic configuration of
alkali metals (ns1)

Elements	**Atomic**	Number of	Valence

in Group 1	**number**	electrons in various shells in the order K L M N P	shell configuration
Li	**3**	2,1	$2s^1$
Na	**11**	2,8,1	$3s^1$
K	**19**	2,8,8,1	$4s^1$
Rb	**37**	2,8,18,8,1	$5s^1$
Cs	**55**	2,8,18,18,8,1	$6s^1$
Fr	**87**	2,8,18,32,18,1	$7s^1$

Similarly, all the elements are arranged in the modern periodic table which contains 18 vertical columns and 7 horizontal rows. The vertical columns are called groups and the horizontal rows are called periods. Groups are numbered 1 to 18 in accordance with the IUPAC recommendation which replaces the old numbering scheme IA to VIIA, IB to VIIB and VIII.

Each period starts with the element having general outer electronic configuration ns1 and ends with np6. Here 'n' corresponds to the period number (principal quantum number). The aufbau principle and the electronic configuration of atoms provide a theoretical foundation for the modern periodic table.

Nomenclature of Elements with Atomic Number Greater than 100

Usually, when a new element is discovered, the discoverer suggests a name following IUPAC guidelines which will be approved after a public opinion. In the meantime, the new element will be called by a temporary name coined using the following IUPAC rules, until the IUPAC recognises the new name.

1. The name was derived directly from the atomic number of the new element using the following numerical roots.

Notation for IUPAC Nomenclature of elements

Digital	0	1	2	3	4	5	6	7	8	9
Root	Nil	Un	Bi	Tri	Quad	Pent	Hex	Sept	Oct	enn
Abbrevation	N	U	B	T	Q	P	H	S	O	e

2. The numerical roots corresponding to the atomic number are put together and 'ium' is added as suffix

3. The final 'n' of 'enn' is omitted when it is written before 'nil' (enn + nil = enil) similarly the final 'i' of 'bi' and 'tri' is omitted when it written before 'ium' (bi + ium = bium; tri + ium = trium)

4. The symbol of the new element is derived from the first letter of the numerical roots.

The following table illustrates these facts.

Name of elements with atomic number above 100

Atomic number	Temp. Name	Temp. Symbol	Name of the element	Symbol
101	Unnilunium	Unu	Mendelevium	Md
102	Unnilbium	Unb	Nobelium	No
103	Unniltrium	Unt	Lawrencium	Lr
104	Unnilquadium	Unq	Rutherfordium	Rf
105	Unnilpentium	Unp	Dubnium	Db
106	Unnilhexium	Unh	Seaborgium	Sg
107	Unnilseptium	Uns	Bohrium	Bh
108	Unniloctium	Uno	Hassium	Hs
109	Unnilennium	Une	Meitnerium	Mt
110	Ununnilium	Unn	Darmstadium	Ds
111	Ununnilium	Uuu	Roentgenium	Rg
112	Ununbium	Uub	Copernicium	Cn
113	Ununtrium	Uut	Nihonium	Nh
114	Ununquadium	Uup	Flerovium	Lv
115	Ununpentium	Uup	Moscovium	Ts
116	Ununhexium	Uuh	Livermorium	Lv
117	Ununseptium	Uus	Tennessine	Ts
118	Ununoctium	Uco	Oganesson	Og

Grouping of Elements based on Electronic Configurations

In the modern periodic table, the elements are organised in 7 periods and 18 groups based on the modern periodic law. The placement of element in the periodic table is closely related to its outer shell electronic configuration. Let us analyse the change in the electronic configuration of elements along the periods and down the groups.

Variation of Electronic Configuration along the periods

We have already learnt that each period starts with the element having general outer electronic configuration ns1 and ends with ns2, np6 where n is the period number. The first period starts with the filling of valence electrons in 1s orbital, which can accommodate only two electrons. Hence, the first period has two elements, namely hydrogen and helium. The second period starts with the filling of valence electrons in 2s orbital followed by three 2p orbitals with eight elements from lithium to neon. The third period starts with filling of valence electrons in the 3s orbital followed by 3p orbitals. The fourth period starts with filling of valence electrons from 4s orbital followed by 3d and 4p orbitals in accordance with Aufbau principle. Similarly, we can explain the electronic configuration of elements in the subsequent periods.

Table 3.10 Electronic configuration of elements in a period

Period number (n)	Filling of electrons in orbitals		Number of elements	Outer shell Electronic configuration	
	Starts from	Ends with		First element	Last element
1	1s →	1s	2	H – $1s^1$	He – $1s^2$
2	2s →	2p	8	Li – $2s^1$	Ar– $2s^22p^6$
3	3s →	3p	8	Na – $3s^1$	Ne – $3s^23p^6$
4	4s → 3d →	4p	18	K – $4s^1$	Kr– $4s^24p^6$
5	5s → 4d →	5p	18	Rb – $5s^1$	Xe – $5s^25p^6$
6	6s → 4f → 5d →	6p	32	Cs – $6s^1$	Rn – $6s^26p^6$
7	7s → 5f → 6d →	7p	32	Fr – $7s^1$	Og – $7s^27p^6$

In the fourth period the filling of 3d orbitals starts with scandium and ends with zinc. These 10 elements are called first transition series. Similarly 4d, 5d and 6d orbitals are filled in successive periods and the corresponding series of elements are called second, third and fourth transition series respectively.

In the sixth period the filling of valence electrons starts with 6s orbital followed by 4f, 5d and 6p orbitals. The filling up of 4f orbitals begins with Cerium (Z=58) and ends at Lutetium (Z=71). These 14 elements constitute the first inner-transition series called Lanthanides. Similarly, in the seventh period 5f orbitals are filled, and it's -14 elements
constitute the second inner-transition series called Actinides. These two series are placed separately at the bottom of the modern periodic table.

Variation of Electronic Configuration in the Groups:

Elements of a group have similar electronic configuration in the outer shell. The general outer electronic configurations for the 18 groups are listed in the Table 3.11. The groups can be combined as s, p, d and f block elements on the basis of the orbital in which the last valence electron enters.

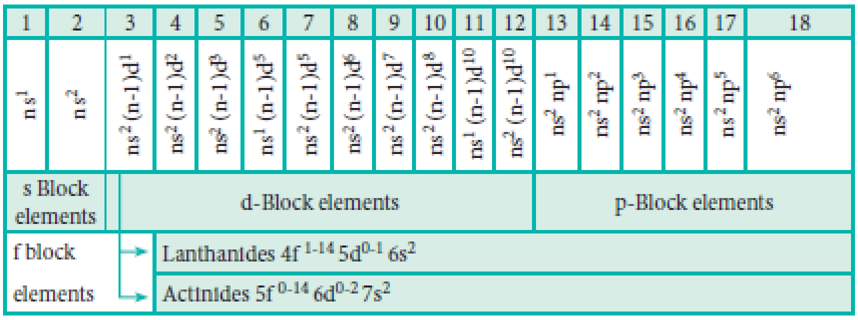
The elements of group 1 and group 2 are called s-block elements, since the last valence electron enters the ns orbital. The group 1 elements are called alkali metals while the group 2 elements are called alkaline earth metals. These are soft metals and possess low melting and boiling points with low ionisation enthalpies. They are highly reactive and form ionic compounds. They are highly electropositive in nature and most of the elements imparts colour to the flame. We will study the properties of these group elements in detail in subsequent chapters.

The elements of groups 13 to 18 are called p-block elements or representative elements and have a general electronic configuration ns2, np1-6. The elements of the group 16 and 17 are called chalcogens and halogens respectively. The elements of 18th group contain completely filled valence shell electronic configuration (ns2, np6) and are called inert gases or nobles gases. The elements of p-block have high negative electron gain enthalpies. The ionisation energies are higher than that of s-block elements. They form mostly covalent compounds and shows more than one oxidation states in their compounds.

The elements of the groups 3 to 12 are called d-block elements or transition elements with general valence shell electronic configuration ns1-2, (n-1)d1-10. These elements also show more than one oxidation state and form ionic, covalent and co-ordination compounds. They can form interstitial compounds and alloys which can also act as catalysts. These elements have high melting points and are good conductors of heat and electricity.

The lanthanides (4f1-14, 5d0-1, 6s2) and the actinides (5f0-14, 6d0-, 7s2) are called f-block elements. These elements are metallic in nature and have high melting points. Their compounds are mostly coloured. These elements also show variable oxidation states.

Table 3.11 General outer electronic configuration of elements in groups:



1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
ns^1	ns^2	$ns^2 (n-1)d^1$	$ns^2 (n-1)d^2$	$ns^2 (n-1)d^3$	$ns^1 (n-1)d^5$	$ns^2 (n-1)d^5$	$ns^2 (n-1)d^6$	$ns^2 (n-1)d^7$	$ns^2 (n-1)d^8$	$ns^1 (n-1)d^{10}$	$ns^2 (n-1)d^{10}$	$ns^2 np^1$	$ns^2 np^2$	$ns^2 np^3$	$ns^2 np^4$	$ns^2 np^5$	$ns^2 np^6$
s Block elements		d-Block elements										p-Block elements					
f block elements		Lanthanides $4f^{1-14} 5d^{0-1} 6s^2$															
		Actinides $5f^{0-14} 6d^{0-2} 7s^2$															

Periodic Trends in Properties

As discussed earlier, the electronic configuration of the elements shows a periodic variation with increase in atomic numbers. Similarly a periodic trend is observed in physical and chemical behaviour of elements. In this section, we will study the periodic trends in the following properties of elements.

1. Atomic radius
2. Ionic radius
3. Ionisation enthalpy (energy)
4. Electron gain enthalpy (electron affinity)
5. Electronegativity

Atomic radius

Atomic radius of an atom is defined as the distance between the centre of its nucleus and the outermost shell containing the valence electron.

It is not possible to measure the radius of an isolated atom directly. Except for noble gases, usually atomic radius is referred to as covalent radius or metallic radius depending upon the nature of bonding between the concerned atoms.

Covalent radius

It is one-half of the internuclear distance between two identical atoms linked together by a single covalent bond. Inter nuclear distance can be determined using x-ray diffraction studies.

Example:

The experimental internuclear distance in Cl_2 molecule is 1.98 Å. The covalent radius of chlorine is calculated as below.

$$d_{Cl-Cl} = r_{Cl} + r_{Cl}$$

$$\Rightarrow \quad d_{Cl-Cl} = 2r_{Cl}$$

$$\Rightarrow \quad r_{Cl} = \frac{d_{Cl-Cl}}{2}$$

$$= \frac{1.98}{2} = 0.99\,\overset{o}{A}$$

The covalent radius of chlorine $= \frac{198}{2}\,pm$

$= 99\,pm$

The formation of covalent bond involves the overlapping of atomic orbitals and it reduces the expected internuclear distance. Therefore covalent radius is always shorter than the actual atomic radius.

The covalent radius of individual atom can also be calculated using the internuclear distance (d_{A-B}) between two different atoms A and B. The simplest method proposed by Schomaker and Stevenson is as follows.

$$d_{A-B} = r_A + r_B - 0.09\left(\chi_A - \chi_B\right)$$

where χ_A and χ_B are the electronegativities of A and B respectively in Pauling units. Here $\chi_A > \chi_B$ and radius is in Å.

Let us calculate the covalent radius of hydrogen using the experimental dH-Cl value is 1.28 Å and the covalent radius of chlorine is 0.99 Å. In pauling scale the electronegativity of chlorine and hydrogen are 3 and 2.1 respectively.

$d_{H\text{-}Cl} = r_H + r_{Cl} - 0.09\ (\chi_{Cl} - \chi_H)$
$1.28 = r_H + 0.09 - 0.09\ (3 - 2.1)$
$1.28 = r_H + 0.09 - 0.09\ (0.9)$
$1.28 = r_H + 0.09 - 0.081$
$1.28 = r_H + 0.909$
$\therefore r_H = 1.28 - 0.909 = 0.317$ Å

Metallic radius

It is defined as one-half of the distance between two adjacent metal atoms in the closely packed metallic crystal lattice.

For example, the distance between the adjacent copper atoms in solid copper is 2.56 Å and therefore the metallic radius of copper is

$$\frac{2.56}{2} = 1.28\text{Å}$$

The metallic radius can be calculated using the unit cell length of the metallic crystal. You will study the detailed calculation procedure in XII standard solid state unit.

Periodic Trends in Atomic Radius

Variation in Periods

Atomic radius tends to decrease in a period. As we move from left to right along a period, the valence electrons are added to the same shell. The simultaneous addition of protons to the nucleus, increases the nuclear charge, as well as the electrostatic attractive force between the valence electrons and the nucleus. Therefore atomic radius decreases along a period.

Effective nuclear charge

In addition to the electrostatic forces of attraction between the nucleus and the electrons, there exists repulsive forces among the electrons. The repulsive force between the inner shell electrons and the valence electrons leads to a decrease in the electrostatic attractive forces acting on the valence electrons by the nucleus. Thus, the inner shell electrons act as a shield between the nucleus and the valence electrons. This effect is called shielding effect.

The net nuclear charge experienced by valence electrons in the outermost shell is called the effective nuclear charge. It is approximated by the below mentioned equation.

$$Z_{eff} = Z - S$$

Where Z is the atomic number and 'S' is the screening constant which can be calculated using Slater's rules as described below.

Step 1:

Write the electronic configuration of the atom and rearrange it by grouping ns and np orbitals together and others separately in the following form. (1s) (2s, 2p) (3s, 3p) (3d) (4s, 4p) (4d) (4f) (5s, 5p)...

Step 2:

Identify the group in which the electron of interest is present. The electron present right to this group does not contribute to the shielding effect.

Each of the electrons within the identified group (denoted by 'n') shields to an extent of 0.35 unit of nuclear charge. However, it is 0.30 unit for 1s electron.

Step 3 :

Shielding of inner shell electrons.

If the electron of interest belongs to either s or p orbital,

i) each electron within the (n-1) group shields to an extent of 0.85 unit of nuclear charge, and

ii) each electron within the (n-2) group (or) even lesser group (n-3, (n-4) etc... completely shields i.e. to an extent of 1.00 unit of nuclear charge. If the electron of interest belongs to d or f orbital, then each of electron left of the group of electron of interest shields to an extent of 1.00 unit of nuclear charge.

Step 4 :

Summation of the shielding effect of all the electrons gives the shielding constant 'S'

Example: Let us explain the calculation of effective nuclear charge on 4s electron and 3d electron in scandium. The electronic configuration of scandium is $1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$, $3d^1$. we can rearrange as below.

$$\underbrace{(1s)^2}_{(n-3)}\quad \underbrace{(2s,2p)^8}_{(n-2)}\quad \underbrace{(3s,3p)^8\,(3d)^1}_{(n-1)}\quad \underbrace{(4s)^2}_{n}$$

Group	number of electron in the group	contribution of each electron to 'S'value	contribution of aparticular groupto electrons to 'S'value
(n)	1	0.35	0.35
(n-1)	9	0.85	7.65
(n-2) & others	10	1	10.00
S value			18.00

$Z_{eff} = Z - S$ i.e.= 21- 18 $\therefore Z_{eff} = 3$

Calculation of effective nuclear charge on 3d electron

$$\underbrace{(1s)^2}_{(n-3)}\quad \underbrace{(2s,2p)^8}_{(n-2)}\quad \underbrace{(3s,3p)^8}_{(n-1)}\ \underset{n}{(3d)^1}\quad \underbrace{(4s)^2}$$

Group	number of electron in the group	contribution of each electron to 'S' value	contribution of a particular group to electrons to 'S' value
n	0	0.35	0
(n-1) & others	18	1	18
S Value			18

$\therefore Z_{eff} = Z - S$ i.e. =21 - 18 $\therefore \therefore Z_{eff} = 3$

Shielding effect from inner shell electrons (Slater's rules)

Electron Group	Electron of interest either S or P	Electron of interest either d or f
n	0.35 (0.30 for (S electron)	0.35
(n-1)	0.85	1.00
(n-2) and others	1.00	1.00

Atomic radius (covalent radius) of second period elements.

Elements	Effective nuclear charge	Covalent radius (pm)
Li^3	1.30	167
Be^4	1.95	112
C^6	2.60	87
N^7	3.25	67
O^8	3.25	56
F^9	4.55	48
Ne^{10}	5.85	38

Variation in Group

In the periodic table, the atomic radius of elements increases down the group. As we move down a group, new shells are opened to accommodate the newly added valence electrons. As a result, the distance between the centre of the nucleus and the outermost shell containing the valence electron increases. Hence, the atomic radius increases. The trend in the variation of the atomic radius of the alkali metals down the group os shown below.

Variation of covalent radius of group 1 elements

Element	Outermost shell containing valence electron	Covalent radius (Å)
Li	L (n=2)	1.35
Na	M (n=3)	1.54
K	N (n=4)	1.96
Rb	O (n=5)	2.11
Cs	P(n=6)	2.25

Ionic radius

It is defined as the distance from the centre of the nucleus of the ion up to which it exerts its influence on the electron cloud of the ion. Ionic radius of uni-univalent crystal can be calculated using Pauling's method from the inter ionic distance between the nuclei of the cation and anion. Pauling assumed that ions present in a crystal lattice are perfect spheres, and they are in contact with each other therefore,

Where d is the distance between the centre of the nucleus of cation C+ and anion A- and rC +, rA - are the radius of the cation and anion respectively.

Pauling also assumed that the radius of the ion having noble gas electronic configuration (Na+ and Cl^- having $1s^2$ $2s^2$, $2p^6$ configuration) is inversely proportional to the effective nuclear charge felt at the periphery of the ion.

$$i.e. r_C + \alpha \frac{1}{\left(Z_{eff}\right)C+} ....(1)$$

$$r_A - \alpha \frac{1}{\left(Z_{eff}\right)_{A^-}} ....(3)$$

Where Z_{eff} is the effective nuclear charge and $Z_{eff} = Z - S$

Dividing the equation 1 by 3

$$\frac{r_{C^+}}{r_{A^-}} = \frac{\left(Z_{eff}\right)_{A^-}}{\left(Z_{eff}\right)_{C^+}} ....(4)$$

On solving equation (1) and (4) the values of r_C and r_A can be obtained

Let us explain this method by calculating the ionic radii of Na+ and F- in NaF crystal whose interionic distance is equal to 231 pm .

$$d = r_{Na^-} + r_F ....(5)$$

$$i.e. r_{Na^+} + r_{F^-} = 231 pm$$

we know that

$$\frac{r_{Na^+}}{r_{F^-}} = \frac{\left(Z_{eff}\right)_{F^-}}{\left(Z_{eff}\right)_{Na^+}}$$

$$\left(Z_{eff}\right)_{F^-} = Z - S$$

$$= 9 - 4.15$$

$$= 4.85$$

$$\left(Z_{eff}\right)_{Na^+} = 11 - 4.15$$

$$= 6.85$$

$$\therefore \frac{r_{Na^+}}{r_{F^-}} = \frac{4.85}{6.85}$$

$$= 0.71$$

$$\Rightarrow r_{Na^+} = 0.71 \times r_{F^-}$$

Substitting (3) in (1)

$$(1) \Rightarrow 0.71 r_{F^-} + r_{F^-} = 231 pm$$

$$1.71 r_{F^-} = 231 pm$$

$$r_{F^-} = \frac{231}{1.71} = 135.1 pm$$

Substituting the value of r_{F^-} in equation (1)

$$r_{Na^+} + 135.1 = 231$$

$$r_{Na^+} = 95.9\, pm$$

Ionisation energy

It is defined as the minimum amount of energy required to remove the most loosely bound electron from the valence shell of the isolated neutral gaseous atom in its ground state. It is expressed in kJ mol-1 or in electron volts (eV).

$$M_{(g)} + IE_l \rightarrow M^+{}_{(g)} + 1e^-$$

Where IE_1 represents the first ionisation energy.

Successive Ionisation energies

The minimum amount of energy required to remove an electron from a unipositivecation is called second ionisation energy. It is represented by the following equation.

$$M^{+}{}_{(g)} + IE_2 \rightarrow M^{2+}{}_{(g)} + 1e^{-}$$

In this way we can define the successive ionisation energies such as third, fourth etc.

The total number of electrons are less in the cation than the neutral atom while the nuclear charge remains the same. Therefore the effective nuclear charge of the cation is higher than the corresponding neutral atom. Thus the successive ionisation energies, always increase in the following order

$$IE_1 < IE_2 < IE_3 < .....$$

Periodic Trends in Ionisation Energy

The ionisation energy usually increases along a period with few exceptions. As discussed earlier, when we move from left to right along a period, the valence electrons are added to the same shell, at the same time protons are added to the nucleus. This successive increase of nuclear charge increases the electrostatic attractive force on the valence electron and more energy is required to remove the valence electron resulting in high ionisation energy.

Let us consider the variation in ionisation energy of second periodelements. The plot of atomic number vs ionisation energy is given below.

In the following graph, there are two deviation in the trends of ionisiation energy. It is expected that boron has higher ionisation energy than beryllium since it has higher nuclear charge. However, the actual ionisation energies of beryllium and boron are 899 and 800 kJ mol^{-1}respectively contrary to the expectation. It is due to the fact that beryllium with completely filled 2s orbital, is more stable than partially filled valence shell electronic configuration of boron. ($2s^2$,$2p^1$)

The electronic configuration of beryllium (Z=4) in its ground state is $1s^2$, $2s^2$ and that of boran (Z = 5) $1s^2$ $2s^2$ $2p^1$

Similarly, nitrogen with $1s^2$, $2s^2$, $2p^3$ electronic configuration has higher ionisation energy (1402 kJ mol^{-1}) than oxygen (1314 kJ mol-1). Since the half filled electronic configuration is more stable, it requires higher energy to remove an electron from 2p orbital of nitrogen. Whereas the removal one 2p electron from oxygen leads to a stable half filled configuration. This makes comparatively easier to remove 2p electron from oxygen.

Periodic variation in group

The ionisation energy decreases down a group. As we move down a group, the valence electron occupies new shells, the distance between the nucleus and the valence electron increases. So, the nuclear forces of attraction on valence electron decreases and hence ionisation energy also decreases down a group.

Ionisation energy and shielding effect

As we move down a group, the number of inner shell electron increases which in turn increases the repulsive force exerted by them on the valence electrons, i.e. the increased shielding effect caused by the inner electrons decreases the attractive force acting on the valence electron by the nucleus. Therefore the ionisation energy decreases.

Let us understand this trend by considering the ionisation energy of alkali metals.

Electron affinity

It is defined as the amount of energy released (required in the case noble gases) when an electron is added to the valence shell of an isolated neutral gaseous atom in its ground state to form its anion. It is expressed in kJ mol^{-1}

$$A + 1e^{-} \rightarrow A^{-} + E_A$$

Variation of Electron Affinity in a period:

The variation of electron affinity is not as systematic as in the case of ionisation energy. As we move from alkali metals to halogens in a period, generally electron affinity increases, i.e. the amount of energy released will be more. This is due to an increase in the nuclear charge and decrease in size of the atoms. However, in case of elements such as beryllium ($1s^2$, $2s^2$), nitrogen ($1s^2$, $2s^2$, $2p^3$) the addition of extra electron will disturb their stable electronic configuration and they have almost zero electron affinity.

Variation of electron affinity (electron gain energy) along I period

Noble gases have stable ns^2, np^6 configuration, and the addition of further electron is unfavourable and requires energy. Halogens having the general electronic configuration of ns2, np5 readily accept an electron to get the stable noble gas electronic configuration (ns^2, np^6), and therefore in each period the halogen has high electron affinity. (high negative values)

Variation of Electron affinity in a group:

As we move down a group, generally the electron affinity decreases. It is due to increase in atomic size and the shielding effect of inner shell electrons. However, oxygen and fluorine have lower affinity than sulphur and chlorine respectively. The sizes of oxygen and fluorine atoms are comparatively small and they have high electron density. Moreover, the extra electron added to oxygen and fluorine has to be accommodated in the 2p orbital which is relatively compact compared to the 3p orbital of sulphur and chlorine so, oxygen and fluorine have lower electron affinity than their respective group elements sulphur and chlorine.

Electronegativity

It is defined as the relative tendency of an element present in a covalently bonded molecule, to attract the shared pair of electrons towards itself.

Electronegativity is not a measurable quantity. However, a number of scales are available to calculate its value. One such method was developed by Pauling, he assigned arbitrary value of electronegativities for hydrogen and fluorine as 2.2 and 4.0 respectively. Based on this the electronegativity values for other elements can be calculated using the following expression

$$\left(X_A - X_B\right) = 0.182\sqrt{E}_{AB} - \left(E_{AA} * E_{BB}\right)^{1/2}$$

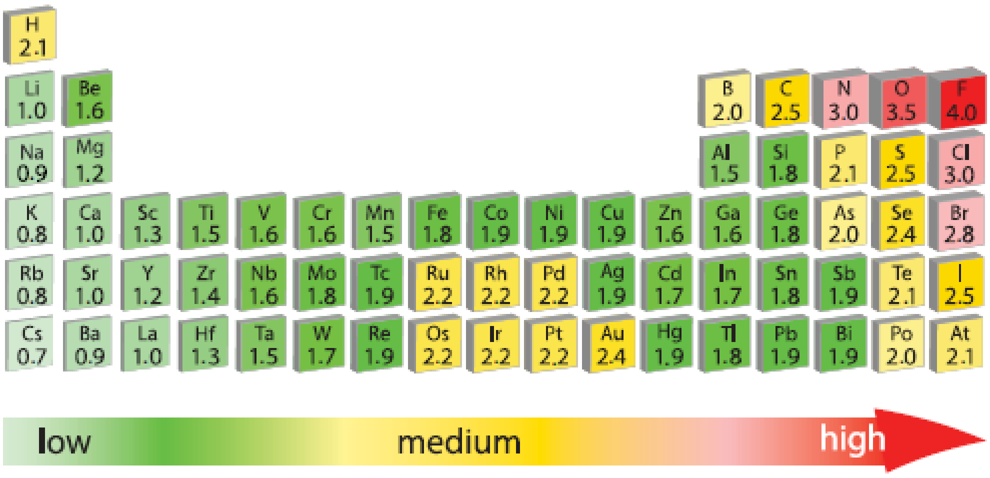
Where E_{AB}, E_{AA} and E_{BB} are the bond dissociation energies of AB, A_2 and B_2 molecules respectively.

The electronegativity of any given element is not a constant and its value depends on the element to which it is covalently bound. The electronegativity values play an important role in predicting the nature of the bond.

Variation of Electronegativity in a period:

The electronegativity generally increases across a period from left to right. As discussed earlier, the atomic radius decreases in a period, as the attraction between the valence electron and the nucleus increases. Hence the tendency to attract shared pair of electrons increases. Therefore, electronegativity also increases in a period.

Paulings scale of electronegativity valuse of elements



H 2.1																
Li 1.0	Be 1.6											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
Cs 0.7	Ba 0.9	La 1.0	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.1

Variation of Electronegativity in a group:

The electronegativity generally decreases down a group. As we move down a group the atomic radius increases and the nuclear attractive force on the valence electron decreases. Hence, the electronegativity decreases.

Noble gases are assigned zero electronegativity. The electronegativity values of the elements of s-block show the expected decreasing order in a group. Except 13th and 14th group all other p-block elements follow the expected decreasing trend in electronegativity.

Periodic Trends in Chemical Properties:

So far, we have studied the periodicity of the physical properties such as atomic radius, ionisation enthalpy, electron gain enthalpy and electronegativity. In addition, the chemical properties such as reactivity, valence, oxidation state etc… also show periodicity to certain extent.

In this section, we will discuss briefly about the periodicity in valence (oxidation state) and anomalous behaviour of second period elements (diagonal relationship).

Valence or Oxidation States

The valence of an atom is the combining capacity relative to hydrogen atom. It is usually equal to the total number of electrons in the valence shell or equal to eight minus the number of valence electrons. It is more convenient to use oxidation state in the place of valence.

Periodicity of Valence or Oxidation States

The valence of an atom primarily depends on the number of electrons in the valence shell. As the number of valence electrons remains same for the elements in same group, the maximum valence also remains the same. However, in a period the number of valence electrons increases, hence the valence also increases.

Variation of valence in groups

Alkali Metals (Group 1)			**Group 15**		
Element	**No. of electrons in valence shell**	**Valence**	**Element**	**No. of electrons in valence shell**	**Valence**
Li	1	1	N	5	3,5
Na	1	1	P	5	3,5
K	1	1	As	5	3,5
Rb	1	1	Sb	5	3,5
Cs	1	1	Bi	5	3,5
Fr	1	1			

Variation of valence in period (1st period)

Element	Li	Be	B	C	N	O	F	Ne
No. of electrons in valence shell	1	2	3	4	5	6	7	8
Valence (Combining capacity)	1	2	3	4	5,3	6,2	7,1	8,0

In addition to that some elements have variable valence. For example, most of the elements of group 15 which have 5 valence electrons show two valences 3 and 5. Similarly transition metals and inner transition metals also show variable oxidation states.

Anomalous properties of second period elements:

As we know, the elements of the same group show similar physical and chemical properties. However, the first element of each group differs from other members of the group in certain properties. For example, lithium and beryllium form more covalent compounds, unlike the alkali and alkali earth metals which predominantly form ionic compounds. The elements of the second period have only four orbitals (2s & 2p) in the valence shell and have a maximum co-valence of 4, whereas the other members of the subsequent periods have more orbitals in their valence shell and shows higher valences. For example, boron forms $BF_4^-$and aluminium forms AlF_6^{3-}.

Diagonal Relationship

On moving diagonally across the periodic table, the second and third period elements show certain similarities. Even though the similarity is not same as we see in a group, it is quite pronounced in the following pair of elements.

The similarity in properties existing between the diagonally placed elements is called 'diagonal relationship'.

Periodic Trends and Chemical Reactivity:

The physical and chemical properties of elements depend on the valence shell electronic configuration as discussed earlier. The elements on the left side of the periodic table have less ionisation energy and readily loose their valence electrons. On the other hand, the elements on right side of the periodic table have high electron affinity and readily accept electrons. As a consequence of this, elements of these extreme ends show high reactivity when compared to the elements present in the middle. The noble gases having completely filled electronic configuration neither accept nor lose their electron readily and hence they are chemically inert in nature.

The ionisation energy is directly related to the metallic character and the elements located in the lower left portion of the periodic table have less ionisation energy and therefore show metallic character. On the other hand the elements located in the top right portion have very high ionisation energy and are non-metallic in nature.

Let us analyse the nature of the compounds formed by elements from both sides of the periodic table. Consider the reaction of alkali metals and halogens with oxygen to give the corresponding oxides.

$4\ Na + O_2 \rightarrow 2\ Na_2O$
$2\ Cl_2 + 7\ O_2 \rightarrow 2\ Cl_2O_7$

Since sodium oxide reacts withwater to give strong base sodiumhydroxide, it is a basic oxide. ConverselyCl_2O_7 gives strong acid called perchloricacid upon reaction with water So, it is anacidic oxide.

$Na_2O + H_2O \rightarrow 2NaOH$
$Cl_2O_7 + H_2O \rightarrow 2\ HClO_4$

Thus, the elements from the two extreme ends of the periodic table behave differently as expected.

As we move down the group, the ionisation energy decreases and the electropositive character of elements increases. Hence, the hydroxides of these elements become more basic. For example,let us consider the nature of the second group hydroxides:

$Be(OH)_2$ amphoteric; $Mg(OH)_2$ weakly basic; $Ba(OH)_2$ strongly basic Beryllium hydroxide reacts with both acid and base as it is amphoteric in nature.

$Be(OH)_2 + HCl \rightarrow BeCl_2 + 2H_2O$
$Be(OH)_2 + 2\ NaOH \rightarrow Na_2BeO_2 + 2H_2O$