CHEMICAL REACTIONS AND EQUATIONS

Introduction: There are two type of changes in our surroundings. These are:

(i) Physical Changes. (ii) Chemical Changes.

Physical Changes: Those changes in which no new chemical substances are formed with new properties are called physical changes. **For example :** melting of ice, evaporation of water etc.

Chemical Changes: Those changes in which new chemical substances are formed with new properties are called chemical changes. **For example :** magnesium burns in the presence of oxygen of air, Burning of coke in air, rusting of iron etc.

Chemical Reactions: Chemical reactions are the processes in which new substences with new properties are formed or whenever a chemical change occurs. **For example:** magnesium burns in the presence of oxygen of air then new compound magnesium oxide is formed with new properties $Mg + O_2 \rightarrow 2MgO$.

Note: During a chemical reaction, atoms of one element do not change into those of another element only a rearrangement of atoms takes place in a chemical reaction.

Reactants & Products

(i) **Reactants:** The substances which take part in a chemical reaction are called reactants in the above examples magnesium and oxygen are reactants.

(ii) **Products:** The new substances produced as a result of chemical reaction are called products in the above example magnesium oxide is a product

Daily examples of chemical reactions:

• Ripening of fruits • digestion of food • souring of milk • rusting of iron • burning of a candle • combustion of glucose (or respiration) • photosynthesis in green plants • growth of plants • formation of curd • burning of fuels • cooking of food etc. are the example of chemical reaction.

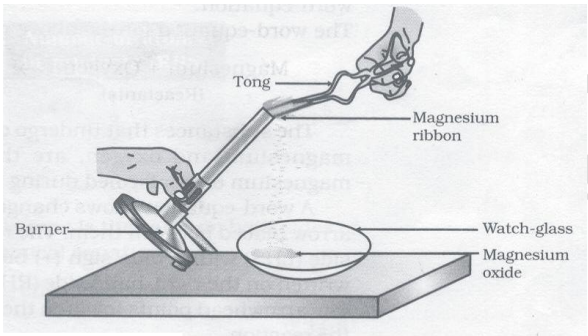
ACTIVITY 1.1

Aim: To demonstrate that burning of magnesium in air is a chemical reaction.

Material Required: Sand paper, Magnesium ribbon, Spirite, lamp and watch glass.

magnesium + oxygen → magnesium oxide
(As ribbon) (white powder)

$$Mg + O_2 \longrightarrow MgO$$
(asribbon) (white powder)



Burning of a magnesium ribbon in air and collection of magnesium oxide in a watch-glass

Procedure:

Step 1 Clean a magnesium ribbon with sand paper.

Step 2 Hold it with a pair of tongs. Burn it in air with a spirit lamp and collect the ash so formed in a watch glass as shown in figure.

Step 3 Collect the white ash in a watch glass, cool it and put a small portion of the white ash on a moist red litmus paper.

Step 4 Record your observations and answer the following questions:

Q.1 How does magnesium burn?

Ans Magnesium burns with dazzling white flame.

Q.2 What is the colour of the product formed when magnesium burns?

Ans A white ash of magnesium oxide is formed.

Q.3 What happens when white ash of magnesium oxide is brought in contact with a moist red litmus paper? What conclusion do you draw from your observation?

Ans The moist red litmus paper turns blue. This shows that magnesium oxide has basic character.

Q.4 Why should magnesium ribbon be cleaned before burning it in air?

Ans To remove dust and a layer of basic magnesium carbonate [$MgCO_3 \cdot Mg(OH)_2$] or Magnesium oxide (MgO)

Observation: We observe that magnesium ribbon burns with a dazzling white flame and changes into a white powder. The powder is magnesium oxide.

Characteristics of Chemical Reactions

There are fine characteristics of chemical reactions. These are:

(i) Evolution of gas.
(ii) Formation of precipitate.
(iii) Change in colour.
(iv) Change in temperature.
(v) Change in state.

1. Evolution of a gas

There are some chemical reacations are characterised by the evolutaion of a gas. **For example**, when zinc granules react with Dil sulphuric acid, then bubbles of hydrogen gas are produced.

$$Zn(s) + H_2SO_4(aa) \rightarrow ZnSO_4(ar) + H_2 .$$

ACTIVITY 1.2

Aim: To study the nature of the reaction between zinc metal and hydrochloric acid.

Material required: Conical flast, zinc granules, oil Hydrochloric.

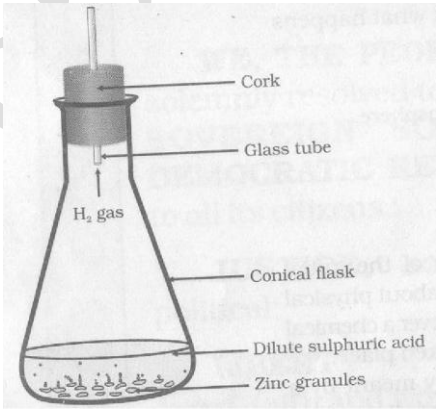
$$\underset{\text{Zinc}}{Zn} + \underset{\text{Hydrochloric acid}}{2HCl} \longrightarrow \underset{\text{Zinc chloride}}{ZnCl_2} + \underset{\text{Hydrogen}}{H_2}$$

Procedure:

Step 1 Take a few zinc granules in a conical flask or a test tube.

Step 2 Add dilute hydrochloric acid or sulphuric acid to this.

Observation: We observe that a colourless gas is produced which burns with the pop sound and test tube becomes slightly warm.



Note: Above chemical reaction is associated with two characteristics

(i) Evolution of gas
(ii) Change in temperature

Precipitate: A precipitate is 'solid product' which separates out from the solution during a chemical reaction.

*A precipitate is formed by mixing two aqueous solutions or passing a gas into an aqueous solution.

2. Formation of a Precipitate

There are some chemical reaction are characterised by the formation of a precipitate. **For Example 1:** when potassium iodide solution is added to a solution of lead nitrate, a yellow precipitate of lead iodide is formed.

$$\underset{\text{Colourless}}{Pb(NO_3)_2} + \underset{\text{Colourless}}{2KI} \rightarrow \underset{\text{Yellow ppt}}{PbI_2} + \underset{\text{Colourless}}{2KNO_3}$$

Precipitation Reactions: Those chemical reactions in which a product is obtained in the form of solid are called precipitation reactions. **For example:** When dilute sulphuric acid is added to bariun chloride solution taken in a test-tube, then a white precipitate of barium sulphate is formed.

$$BaCl_2 + H_2SO_2 \longrightarrow \underset{\text{White ppt}}{BaSO_4} + HCl$$

ACTIVITY 1.3

Aim: To study the nature of the reaction between lead nitrate and potassium iodide solution.

Material required: Beaker, Lead Nitrate Solution, Potassium Iodide solution and conical flock.

Procedure:

Step 1: Take lead nitrate solution in a conical flask.

Step 2: Add potassium iodide (KI) solution to this.

Observation: We observe that a yellow precipitate of lead iodide (PbI_2) is formed.

3. Change in Colour :

There are some chemical reaction are characterised by a change in colour. **For Examples -1:** When citric acid reacts with potassium permanganate solution $(KMnO_4)$ then, puple colour of potassium permanganate becomes colourless.

Example-2: When sulphur dioxide gas (SO_2) is passed through acidified potassium dicromate solution $(K_2Cr_2O_7)$, then, orange colour of **potassium** dichromate solution changes to **green**.

ACTIVITY 1.4

Aim: To study the nature of the reaction between potassium permagnate solution ($KMnO_4$) and citric acid.

Material Required: Potassium Permanganate Solution, Lemon Juice and test tube.

Procedure:

Step 1: Take some dilute potassium permanganate solution ($KMnO_4$) in a test-tube. It has purple colour.

Step 2: Add lemon juice to it dropwise with the help of a dropper and shake the test tube.

Observation: The purple colour of potassium permanganate solution ($KMnO_4$) becomes colourless.

4. Change in Temperature

There are some chemical reactions are characterised by a change in temperature. **For Examples:** When quicklime (CaO) reacts with water, then slaked lime is formed and a lot of heat energy is produced.

$$\underset{\text{Calcium oxide (Quick lime)}}{CaO(s)} + \underset{\text{Water}}{H_2O(l)} \longrightarrow \underset{\text{Calcium hydroxide (Slaked lime)}}{Ca(OH)_2(aq)} + \text{Heat}$$

ACTIVITY 1.5

Aim: To study the nature of the reaction between calcium oxide and water.

Material required: Calcium oxide, glass beaker and water.

Procedure:

Step 1: Take a little quicklime in a hard-glass beaker.

Step 2: Add water to it slowly.

Step 3: Touch a beaker very carefully.

Observation: The beaker feels to quite hot.

5. Change in State

There are some chemical reactions are characterised by a change in state. **For example**, when wax (in the form of solid) is burned then water and carbon dioxide gas are formed. In this Rx. solid substance changes into gaseous substances. When Petrol or Diseal or Lubricanting oil or Kerosene burns in the presence of Oxygen of air then CO_2 and water vapours are formed in this Reaction A liquid substance changes into gaseous form.

Note: There are some chemical reactions associated with more than one characteristic . **For example,** the chemical reaction between zinc granules and dilute sulphuric acid shows two characteristics: evolution of hydrogen gas and change in temperature. Similarly, the chemical reaction between potassium iodide solution and lead nitrate solution shows two characteristics formation of a precipitate and change in colour.

Topic : Chemical Equations

The method of representing a chemical reaction with the help of symbols and formulae of the substances involved in it is known as chemical equation. There are two types of chemical equations. These are:

(1) Word equations

(2) symbol equations.

1. **Word equations:** A word equation links together the names of the reactants with those of the products. **For example.** Magnesium burns in the process of oxygen of air to form a white powder of magnesium oxide, may be written as:

$$\underset{\text{(Reactan ts)}}{\text{Magnesium} + \text{oxygen}} \longrightarrow \underset{\text{(product)}}{\text{Magnesium oxide}}$$

In a word equation:

- The reactant are written on the left hand side (LHS) with plus (+) sign between them.
- The products are written on the Right hand side (RHS) with plus sign (+) sign between them.
- The arrow $(\rightarrow)$ Separates the reactant with the products.
- The direction of the arrow has point towards the product.

Note: Although word equations are quite useful. But they do not give the true Picture of chemical equations.

2. **Symbol Equations:** A chemical equation can be represented in terms of symbol and formulae of the reactants and products. This is known as symbol equations.

For Example: Zinc react with sulphuric acid form of Zinc sulphate and Hydrogen gas. The symbol of Zinc is Zn; The formula of Sulphuric acid is H_2SO_4. The formula of Zinc Sulphate is $ZnSO_4$ and Formula of Hydrogen gas is H_2.

Zinc+Sulphuric acid $\rightarrow$ Zinc Sulphate + Hydrogen gas.

$$Zn + H_2SO_4 \rightarrow SnSO_4 + H_2$$

Note: The symbol equation give the true picture of the chemical reaction than word equation.

Topic: Balanced and unbalanced equations

Balanced Chemical equation: Those chemical equations in which number of atoms of each elements are equal on both the sides are called balanced chemical equations.

$$2Mg + O_2 \longrightarrow 2MgO \text{ (balanced equation)}$$

Name of element	No. of atoms in Reactant	No of atoms in Product
Mg	2	2
O	2	2

Note: Need of balancing a chemical equation is necessary to verify the law of conservation of mass.

Unbalanced chemical equation: Those chemical equations in which number of atoms of each element are not equal on both the sides are called unbalanced chemical equations. **For example:**

$$Mg + O_2 \longrightarrow MgO \text{ (unbalanced equation)}$$

Name of element	No. of atoms in Reactant	No of atoms in Product
Mg	1	1
O	2	1

Balancing of Chemical equations: The process of making the number of atoms each elements are equal on both sides of the equation is called balancing of chemical equations.

Topic: To make equations more informative

The chemical equations can be made more informative by following ways:

(1) By indicating the "Physical States" of the reactant and product

(2) By indicating the "heat exchanges" during chemical reaction.

(3) By indicating the "conditions' under which the reaction takes place.

(1) **To indicate the physical states of reactants and products:** in order to make the equation more informative. The physical state mentioned with the help of certain specific symbols. These symbols are

(a) if the substance in the form of solid then it is respected by symbol 's'.

(b) if the substance in the form of liquid then it is represented by symbol '*l*'.

(c) if the substance in the form of gas then it is represented by symbol 'g' or upward arrow ($\uparrow$).

(d) if the solution of substance prepared in water, then it is represented by (aq).

(e) If the substance obtained in the form of precipitate then it is represented by symbol (ppt) or downward arrow ($\downarrow$).

Example: Zinc metal reacts with dilute sulphuric acid to form zinc sulphate solution and hydrogen gas:

$$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$$

(2) **To indicate the Heat Exchanges in an equations:** There are two types of reactions on the basis of heat exchanges during chemical reactions. These are:

(i) Exothermic reactions

(ii) Endothermic reactions.

(i) Exothermic reactions: Those chemical reactions in which heat energy is evolved during chemical reaction are known as exothermic reactions in an exothemic reactions heat evolved is indicated by writting "+ Heat" or "+ Energy " or " +Heat Energy" on the Products side of an equation or – "Heat" or – "Energy" or –"Heat Energy" on the reactant side

Examples of Exothermic Reaction:

Ex.1: When carbon burns in the present oxygen of air to form carbon dioxide and a lot of Heat Energy is produced.

$$C(s) + O_2(g) \longrightarrow CO_2(q) + \text{Heat}$$

Note: All combustion reactions are the examples of exothermic reactions.

Ex.2: When hydrogen combines with oxygen, water is formed and heat is also produced.

$$\underset{\text{Hydrogen}}{2H_2(g)} + \underset{\text{Oxygen}}{O_2(g)} \longrightarrow \underset{\text{Water}}{2H_2O(I)} + \text{Heat}$$

Ex.3: When coke, charcoal, diamond and graphite burns in the presence of oxygen of air to form carbon dioxide and lot of heat energy produced.

$$\underset{\text{Carbon}}{C(s)} + \underset{\text{Oxygen}}{O_2(g)} \longrightarrow \underset{\text{Carbon dioxide}}{CO_2(g)} + \text{Heat}$$

Note: The formula of cok, carbon, charcoal, diamond and graphid is same i.e. 'c'.

Ex.4: When methane (CH_4) burns in the presence of oxygen of the air to form carbon dioxide water vapours and large amount of heat energy is produced.

$$\underset{\text{Methane}}{CH_4(g)} + \underset{\text{Oxygen}}{2O_2(g)} \longrightarrow \underset{\text{Carbon dioxide}}{CO_2(g)} + \underset{\text{Water}}{2H_2O(l)} + \text{Heat}$$

Note: All the compounds made up of atoms of carbon & hydrogen or carbon, hydrogen, and oxygen burns in the presence of oxygen of air. To form water vapours, carbon dioxide gas and heat energy.

Ex.5: In the respiration process glucose Burns in the presence of oxygen of air to form CO_2, H_2O and a large quantity of energy is produced.

$$\underset{\text{Glucose}}{C_6H_{12}O_6} + \underset{\text{Oxygen}}{6O_2} \longrightarrow \underset{\text{Carbon dioxide}}{6CO_2} + \underset{\text{Water}}{6H_2O} + \text{Heat}$$

Note: The decomposition of vegetable matter into composed is an example of exothermic process

(ii) Endothermic reactions: Those chemical reactions in which heat energy is absorbed during chemical reactions are known as endothermic reactions. In an endothermic reaction heat energy absorbed is indicated by writing "+Heat" or "+heat energy" or "+Energy" on the reactant side or " – heat" or "–Energy" on "–Heat Energy" on the product side"

Examples of Endothermic reactions:

Ex.1 When Nitrogen and oxygen are heated to a temperature of 3000^{o} they combine to form Nitrogen mono oxide.

$$N_2(q) + O_2(q) + \text{Heat} \longrightarrow 2NO$$

Ex. 2 During photosynthesis plants use sun's energy to prepare their food.

$$6CO_2 + 6H_2O \xrightarrow{\text{sun light}} C_6H_{12}O_6 + 6O_2$$

Ex. 3 When electric current passing through water then water decomposes to form hydrogen and oxygen gas

$$H_2O \xrightarrow{\text{electric current}} H_2(g) + O_2(g)$$

Ex. 4 When calcium carbonate is heated it decomposes to form calciumoxide and carbon dioxide

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2$$

Note: All decomposition reactions are the examples of endothermic reactions.

(3) To indicate the conditons under which the reactions take place:

(a) If heat is required for reaction, then the heat sign delta (Δ) is put over the arrow of the equation.

(b) If the reaction take place in the presence of catalyst, the symbol or formula of the catalyst write above or below the arrow sign in the equation.

Examples of Conditions Reactions:

Ex.1: When potassium chlorate $(KClO_3)$ is heated in the presence of maganese dioxide catalyst, it decomposes to form potassium chloride and oxygen gas.

$$2KClO_3(s) \xrightarrow[MnO_2]{\Delta} 2KCl(s) + 3O_2(g)$$

Ex.3: During photosynthesis, plants combine carbon di oxide and water in the presence of 'sunlight' and 'chlorophyll' to make glucose and oxygen gas is given out.

$$6CO_2 + 6H_2O \xrightarrow[\text{Chlorophyll}]{\text{Sunlight}} C_6H_{12}O_6 + 6O_2$$

Topic : Types of Chemical Reactions

There are five types of chemical reactions. These are:

1. Combination reactions
2. Decomposition reactions
3. Displacement reactions
4. Double Displacement reactions
5. Oxidation and reduction reactions or redox reaction.

1. Combination Reactions

Those chemical reactions in which two or more than two substances combine to form a single substance, are called combination reactions. **For Example:** Calcium oxide (CaO) (lime or quick lime) reacts vigorously with water to form calcium Hydroxide $Ca(OH)_2$ (slaked lime or lime water) and produce lot of heat energy.

$$\underset{\text{quick lime}}{CaO(s)} + H_2O(l) \xrightarrow{\text{combination}} \underset{\text{slaked lime}}{Ca(OH)_2} + \text{Heat}$$

Other Examples of Combination Reactions

Ex.1 Magnesium Combine with Oxygen on Heating to form magnesium oxide.

$$\underset{\text{Magnesium}}{2Mg(s)} + \underset{\text{Oxygen}}{O_2(g)} \rightarrow \underset{\text{Magnesium oxide}}{2MgO(s)}$$

Ex.2 Hydrogen burns in oxygen to form water:

$$\underset{\text{Hydrogen}}{2H_2(g)} + \underset{\text{Oxygen}}{O_2(g)} \rightarrow \underset{\text{Water}}{2H_2O(l)}$$

Ex.3 Carbon (coal) burns in air to form carbon dioxide:

$$\underset{\text{(Coal)}}{C(s)} + \underset{\text{Oxygen}}{O_2(g)} \rightarrow \underset{\text{Carbon dioxide}}{CO_2(g)}$$

Ex.4 Hydrogen combines with chlorine to form hydrogen chloride:

$$\underset{\text{Hydrogen}}{H_2(g)} + \underset{\text{Chlorine}}{Cl_2(g)} \longrightarrow \underset{\text{Hydrogen Chloride}}{2HCl(g)}$$

Ex.5 Sodium metal burns in chlorine to form sodium chloride:

$$\underset{\text{Sodium}}{2Na(s)} + \underset{\text{Chlorine}}{Cl_2(g)} \longrightarrow \underset{\text{Sodium Chloride}}{2NaCl(s)}$$

Ex.6 When iron is heated with sulphur, iron sulphide is formed:

$$\underset{\text{Iron}}{Fe(s)} + \underset{\text{Sulphur}}{S(s)} \longrightarrow \underset{\text{Iron sulphide}}{FeS(s)}$$

Ex.7 Ammonia reacts with hydrogen chloride to form ammonium chloride.

$$\underset{\text{Ammonia}}{NH_3(g)} + \underset{\text{Hydrogen Chloride}}{HCl(g)} \longrightarrow \underset{\text{Ammonium chloride}}{NH_4Cl(s)}$$

Ex.8 Carbon Mono Oxide reacts with oxygen to form carbon dioxide:

$$\underset{\text{Carbon Monoxide}}{2CO(g)} + \underset{\text{Oxygen}}{O_2(g)} \longrightarrow \underset{\text{Carbon dioxide}}{2CO_2(g)}$$

Ex.9 Sulphur dioxide reacts with oxygen to form sulphur trioxide.

$$\underset{\text{Sulphur dioxide}}{2SO_2(g)} + \underset{\text{Oxygen}}{O_2(g)} \longrightarrow \underset{\text{Sulphur trioxide}}{2SO_3(g)}$$

ACTIVITY 1.6

Aim: To study the nature of the reaction between calcium oxide and water.

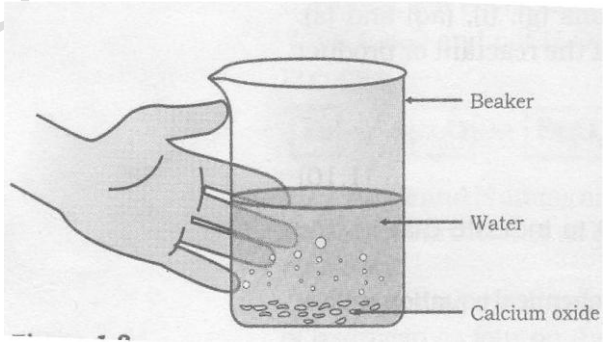
Material Required: Calcium oxide (CaO), beaker and water

Procedure:

Step-1: Take 2g of calcium oxide or quick (CaO) lime in a beaker.

Step-2: Add water to this slowly.

Step-3: Touch the beaker.



Chemical Reaction:

$$CaO + H_2O \longrightarrow Ca(OH)_2 + \text{Heat}$$

Observation: We observe water in beaker becomes hot.

Note: A solution of slaked lime is used for white washing. Calcium Hydroxide reacts slowly with the carbon dioxide in air to form a thin shining layer of calcium carbonate on the walls.

$$Ca(OH)_2(aq) + CO_2(g) \longrightarrow$$

$$CaCO_3(s) + H_2O(l)$$

2. Decomposition Reactions

Those reactions in which a compound splits up into two or more then two simpler substances under suitable conditions, are known as decomposition reactions. There are three types of decomposition reactions, these are:

(i) Thermal Decomposition Reaction.
(ii) Electrolytic Decomposition Reaction.
(iii) Photo Decomposition Reaction.

(i) Thermal Decomposition Reaction: Those chemical reactions in which a single substance break down into two or more than two simpler substances in the presence of heat are called thermal decomposition reaction. **Example-1:** When Ferrous sulphate is heated strongly, it decomposes to form Ferric oxide, sulphur dioxide and Sulphur Trioxide

$$\underset{\text{Green Colour}}{2FeSO_4} \rightarrow \underset{\text{Reddish Brown}}{Fe_2O_3} + SO_2 + SO_3$$

Note: In science laboratory ferrous sulphate crystals are actually hepta hydrate ferrous sulphate $(FeSO_4.7H_2O)$. When the green coloured ferrous sulpahte crystals are heated, they loose 7 molecules of water of crystallisation to form anhydrous ferrous sulphate $(FeSO_4)$ which is white in colour and then anhydrous ferrous sulphate decomposes to give ferric oxide, sulphur dioxide and sulphur tri oxide.

ACTIVITY 1.7

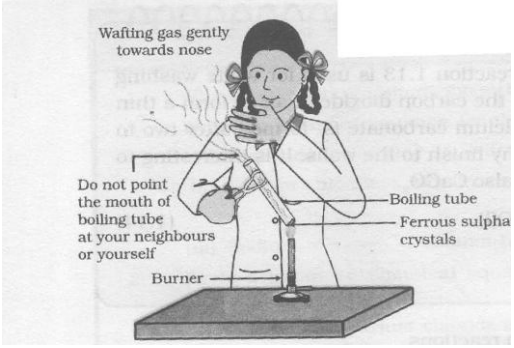
Aim: To study the nature of reaction when ferrous sulphate is heated.

Procedure:

Step 1 Take 2g ferrous sulphate crystals in a dry test tube.

Step 2 Heat the crystals of ferrous sulphate over the flame of a burner.

Reaction: $2FeSO_4 \xrightarrow{\Delta} Fe_2O_3 + SO_2 + SO_3$



Observations: Before heating, the sample of ferrous sulphate is light green. As ferrous sulphate is heated a colourless gas (SO_2) is formed. The green sample changes into a reddish brown solid due to formation of ferric oxide (Fe_2O_3).

Ex.2 When lead Nitrate is heated strongly it breaks down to form simpler substance like lead monoxide, Nitrogen dioxide and oxygen.

$$\underset{\text{Colourless}}{2Pb(NO_3)_2} \xrightarrow[\text{Decomposition}]{\text{heat}}$$

$$\underset{\text{Yellow Colour}}{2PbO(s)} + \underset{\text{Brown Fumes}}{4NO_2(g)} + O_2(g)$$

ACTIVITY 1.8

Aim: To study the type of of chemical reaction involved when lead nitrate is heated.

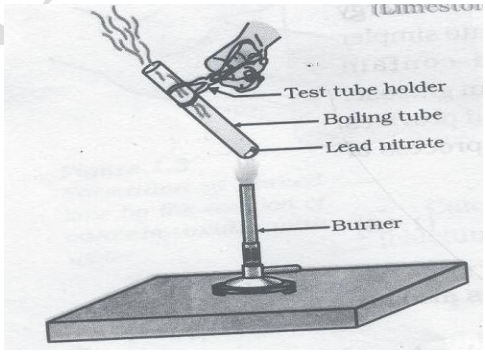
Procedure:

Step 1 Take 2g lead nitrate powder in a test tube.

Step 2 Hold the test tube with the help of a pair of tongs and heat it over the flame.

Reaction:

$$2Pb(NO_3)_2 \xrightarrow{\Delta} 2PhO^{(s)} + 9NO_2^{(q)} + O_2(q)$$



Observation: On heating, lead nitrate breaks up with cracking sound a brown and irritating gas (NO_2) comes out and a yellow solid (PbO) is left in the test tube. The gas produces a dark blue colour on the starch-iodide filter paper. Glowing splint burns brilliantly.

Other Examples of decomposition reaction:

Ex.2 When calcium carbonate is heated, it decomposes to give calcium oxide and carbon dioxide:

$$\underset{\substack{\text{Calcium carbonate}\\\text{(Limestone)}}}{CaCO_3(s)} \xrightarrow[\text{(Decomposition)}]{\text{Heat}} \underset{\substack{\text{Calcium oxide}\\\text{(Lime)}}}{CaO(s)} + \underset{\text{Carbon dioxide}}{CO_2(g)}$$

Ex.3 When potassium chlorate is heated in the presence of manganese dioxide catalyst, it decomposes to give potassium chloride and oxygen:

$$\underset{\text{Potassium chlorate}}{2KClO_3\,(s)} \xrightarrow[\text{(Decomposition)}]{\text{Heat}} \underset{\text{Potassium Chloride}}{2KCl(s)} + \underset{\text{Oxygen}}{3O_2\,(g)}$$

(ii) Electrolytic Decomposition: Those chemical reactions in which a single substance break down into two or more than two simpler substances in the presence of electric current are called Electrolytic decomposition reactions.

Ex.1 When electric current is passes through acidified water, it decomposes to give hydrogen gas and oxygen gas.

$$2H_2O\,(l) \xrightarrow{\text{Electricity}} H_2\,(g) + O_2\,(g)$$

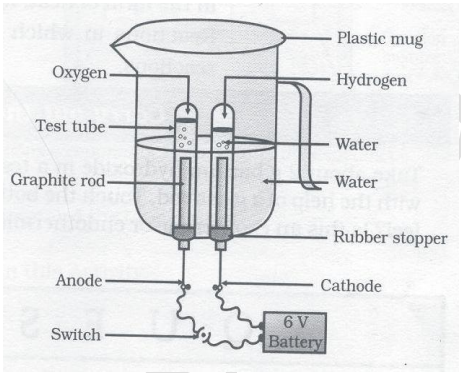
ACTIVITY 1.9

Aim: To study the nature of reaction when electricity is passed through acidified water.

Procedure:

Step-1: Take a plastic mug. Drill two holes at the base and insert carbon electrodes as shown in figure.

(iii)



Step-2: Connect these electrodes to a 6 volt battery. This is voltmeter.

Step-3: Fill the mug with water to its half and add a few drops of dilute sulphuric acid.

Step-4: Fill a graduated test tube with water and invert it in the apparatus so that one of two electrodes in enclosed within the test tube as shown in figure. Similarly, fill the other graduated test tube with water and invert it to enclose the other electrode.

Step-5: Allow the current to pass through the voltmeter and leave the apparatus undisturbed for some time.

Step-6: The bulb formation starts at both the electrodes.

Observation: From all ions.

Other Examples of decomposition reaction:

Ex.2 When electric current is passed through molten sodium chloride, it decomposes to give sodium metal and chlorine gas:

$$\underset{\substack{\text{Sodium Chlorine}\\\text{(Molten)}}}{2NaCl(l)} \xrightarrow[\text{(Decomposition)}]{\text{Electricity}} \underset{\text{Sodium metal}}{2Na\,(s)} + \underset{\text{Chlorine gas}}{Cl_2\,(g)}$$

Ex.3 When electric current is passed through molten aluminium oxide, it composes to give aluminium metal and oxygen gas:

$$\underset{\substack{\text{Aluminium oxide}\\\text{(Molten)}}}{2Al_2O_3\,(l)} \xrightarrow[\text{(Decomposition)}]{\text{Electricity}} \underset{\text{Aluminium metal}}{4Al(l)} + \underset{\text{Oxygen gas}}{3O_2\,(g)}$$

Photo Decomposition or photolysis: Those chemical reactions in which a single substance break down into two or more than two simpler substances in the presence of light are called photo decomposition reactions.

Ex.1: When silver chloride is exposed to light, it decomposes to form silver metal and chlorine gas.

$$\underset{\text{(white)}}{2AgCl(s)} \xrightarrow[\text{decomposition}]{} \underset{\text{(grey)}}{2Ag(s)} + Cl_2\,(g)$$

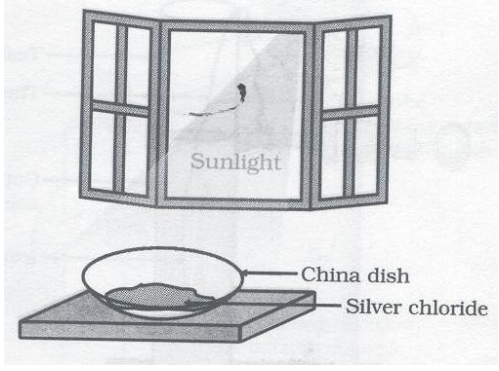
ACTIVITY 1.10

Aim: To study the nature of reaction when silver chloride is exposed to sunlight.

Procedure:

Step 1 Take 2g silver chloride which is white in colour on a watch glass.

Step 2 Place this watch glass under sunlight for some time.



Observation: After some time the colour of white silver chloride turns grey in sunlight.

Ex.2 When silver Bromide is exposed to light, it decomposes to form silver metal and Bromine gas.

$$\underset{\text{Silver chloride (White)}}{2AgBr} \xrightarrow[\text{(Decomposition)}]{} \underset{\text{Silver (Grey)}}{2Ag(s)} + \underset{\text{Bromine}}{Br_2(g)}$$

Reactivity series of metals

The arrangement of metals in a vertical column according to decreasing. Their reactivity is called reactivity or activity series of metals:

Decomposition Reactions in our body: The digestion of food in our body takes place through a number of decomposition reactions. The food that we eat mainly contains starch (from wheat, rice, potatoes etc) and proteins (from pulses, egg, meat etc). In the presence of "enzymes", starch decomposes into simple sugars called **glucose**. Similarly, **proteins decompose** into simpler molecules, called **amino acids.**

Uses of decomposition reactions

(i) The decomposition reactions are used to extract metals from their ores. **For examples:**

Ex. (1) When electric current is passed through molten sodium chloride, it decomposes to give sodium metal and chlorine gas:

$$\underset{\text{Sodium Chlorine (Molten)}}{2NaCl(l)} \xrightarrow[\text{(Decomposition)}]{\text{Electricity}} \underset{\text{Sodium metal}}{2Na(s)} + \underset{\text{Chlorine gas}}{Cl_2(g)}$$

Ex. (2) When electric current passes through molten Aluminium Oxide, it decomposes to given Aluminium metal and chlorine gas.

$$2Al_2O_3 \xrightarrow[\text{Decomposer}]{\text{Electricity}} \underset{\text{Aluminium Metal}}{4Al} + 3O_2$$

(ii) The decomposition reaction is used to obtain calcium oxide from calcium carbonate which is used to make cement.

$$CaCO_3 \xrightarrow{\Delta} CaO + CO_2$$

3. Displacement Reactions

Those chemical reactions in which highly reactive metal displaces less reacts metal from their solution.

Ex.1 When a piece of iron metal is placed in copper sulphate solution, then blue colour of copper sulphate solution becomes turns to dirty green due to formation of ferrous sulphate. A red brown layer of copper metal is formed on the surface of Iron metal

$$CuSO_4(aq) + Fe(s) \longrightarrow FeSO_4(aq) + Cu(s)$$

ACTIVITY 1.11

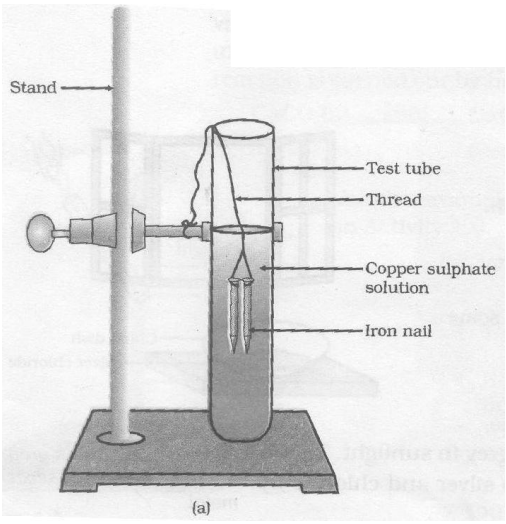
Aim: To study the nature of reaction between iron and copper sulphate solution

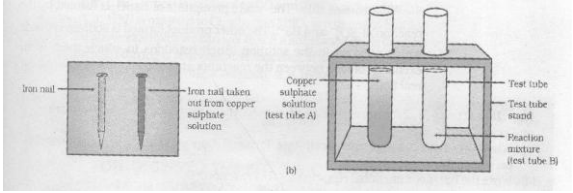
Procedure:

Step 1 Take three iron nails and clean them by rubbing with sand paper.

Step 2 Take 10ml each of copper sulphate solution in test tubes (A) and (B).

Step 3 Tie two iron nails with a thread and immerse them carefully for comparison. After 30 minutes, take out the iron nail from copper sulphate solution.





Observation: We observe that blue colour of copper sulphate solution gradually changes into greenish colour.

Nature of metal	Name of Metal	Symbols	
Metals more reactive than Hydrogen	Potassium	K	Most Reactive
	Sodium	Na	
	Calcium	Ca	
	Magnesium	Mg	
	Aluminium	Al	REACTIVE DECREASES
	Zinc	Zn	
	Iron	Fe	
	Nickel	Ni	
	Tin	Sn	
	Lead	Pb	
	Hydrogen H		
Metals less reactive than hydrogen	Copper	Cu	
	Mercury	Hg	
	Silver	Ag	
	Gold	Au	
	Platinum	Pt	Last Reactive

Note: Element hydrogen (Non metal) is placed in the reactivity series of metal because hydrogen is also form positive ions like metals.

Ex.2 When a strip of zinc metal is placed in copper sulphate solution, then blue colour of Copper Sulphate solution becomes colourless due to formation of Zinc Sulphate. A red brown layer of copper metal is formed on the surface of Zinc Strip.

$$\underset{\text{Copper sulphate (Blue solution)}}{CuSO_4(aq)} + \underset{\text{Zinc (Silvery White)}}{Zn(s)} \longrightarrow$$

$$\underset{\text{Zinc sulphate (Colourless Solution)}}{ZnSO_4(aq)} + \underset{\text{Copper (reddish brown)}}{Cu(s)}$$

Ex.3 When a piece of magnesium metal is placed in copper sulphate solution, then blue coper of Copper Sulphate becomes colourless due to formation of Magnesium Sulphate and a red brown layer of Copper Metal is formed on the Magnesium piece.

$$\underset{\text{Copper sulphate (Blue solution)}}{CuSO_4(aq)} + \underset{\text{Magnesium (Silvery White)}}{Mg(s)} \longrightarrow$$

$$\underset{\text{Magnesium sulphate (Colourless solution)}}{MgSO_4(aq)} + \underset{\text{Copper (red brown)}}{Cu(s)}$$

Ex.4 When a strip of lead metal is placed in a solution of copper chloride, then blue colour of Copper Sulphate solution becomes colourless due to formation of Lead Chloride and a red brown layer of copper metal is deposited on the Lead Strip.

$$\underset{\text{Copper chloride (Blue solution)}}{CuCl_2(aq)} + \underset{\text{Lead (Bluish Grey)}}{Pb(s)} \longrightarrow$$

$$\underset{\text{Lead chloride (Colourless solution)}}{PbCl_2(aq)} + \underset{\text{Copper (red brown)}}{Cu(s)}$$

Ex.5 When a copper strip is placed in a solution of silver nitrate, then solution of Silver Nitrate turned into Blue due to formation of Copper Nitrate and a shining greyish white deposite of Silver is formed on the Copper Strip.

$$\underset{\text{Silver Nitrate (Colourless solution)}}{2AgNO_3(aq)} + \underset{\text{Copper (red brown)}}{Cu(s)} \longrightarrow$$

$$\underset{\text{Copper nitrate (Blue solution)}}{Cu(NO_3)_2(aq)} + \underset{\text{Silver (Greyish white)}}{2Ag(s)}$$

Note: The reaction with metal oxide with other metal which is more reactive then metal present in oxide are the examples of displacement of reaction.

4. Double Displacement Reaction

Those chemical reactions in which two compounds react by an exchange of ions to form to two new compounds are called double displacement reactions.

Examples of Double Displacement

Ex.1 When barium chloride solution is added to sodium sulphate solution, then a white ppt of Barium Sulphate is formed along with Sodium Chloride Solution.

$$BaCl_2 + Na_2SO_4 \longrightarrow \underset{\text{(White ppt)}}{BaSO_4} + 2NaCl$$

Ex.2 If barium chloride solution is added to copper sulphate solution, then a white precipitate of barium sulphate is produced Along with copper chloride solution:

$$\underset{\text{Barium chloride}}{BaCl_2(aq)} + \underset{\text{Copper sulphate}}{CuSO_4(aq)} \longrightarrow$$

$$\underset{\text{Barium sulphate (White ppt.)}}{BaSO_4(s)} + \underset{\text{Copper chloride}}{CuCl_2(aq)}$$

Ex.3 When ammonium hydroxide solution is added to aluminium chloride solution, then a white precipitate of aluminium hydroxide is formed alongwith amomonium chloride solution:

$$\underset{\text{Aluminium Cloride}}{AlCl_3(aq)} + 3\underset{\text{Ammonium hydroxide}}{NH_4OH(aq)} \longrightarrow$$

$$\underset{\text{Aluminium hydroxide (White ppt.)}}{Al(OH)_3(s)} + \underset{\text{Ammonium Chloride}}{3NH_4Cl(aq)}$$

Ex.4 When potassium iodide solution is added to lead nitrate solution, then a yellow precipitate of lead iodide is produced alongwith potassium nitrate solution:

$$\underset{\text{Lead nitrate}}{Pb(NO_3)_2(aq)} + \underset{\text{Potassium iodide}}{2KI(aq)} \longrightarrow$$

$$\underset{\text{Leadiodide (Yellow ppt.)}}{PbI_2(s)} + \underset{\text{Potassium nitrate}}{2KNO_3(aq)}$$

Note: The reaction between base and acid to form salt and water is called neutralisation reaction.

Ex.5 When silver nitrate react with sodium chloride solution then white precipitate of silver chloride is formed along with sodium nitrate solution:

$$AgNO_3(aq) + NaCl(aq) \longrightarrow$$

$$\underset{\text{white ppt}}{AgCl} + NaNO_3(aq)$$

Ex.6 When copper sulphate react with ammonium hydroxide solution then bluish white ppt of copper hydroxide is formed along with ammonium sulphate solution:

$$CuSO_4(aq) + 2NH_4OH(aq) \longrightarrow$$

$$\underset{\text{Bluish white ppt}}{Cu(OH)_2} + (NH_4)_2SO_4(aq)$$

ACTIVITY 1.12

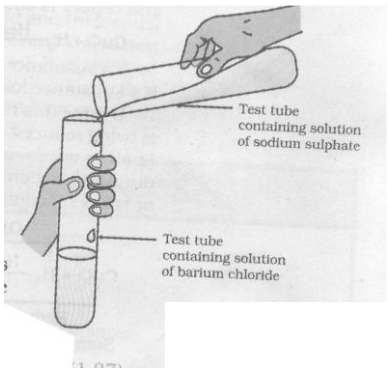
Aim: To study the nature of reaction between barium chloride and Sodium sulphate in aqueous solution.

Procedure:

Step 1 Take 3ml. of sodium sulphate solution in a test tube.

Step 2 In another test tube, take 3ml. of barium chloride.

Step 3 Mix the two solutions with gentle shaking and leave it undisturbed for some time and observe.



Observations:

After mixing the solutions in the conical flask: A white precipitate is formed and settles at the bottom. There is clear and colourless solution above the white solid.

5. Topic : Oxidation and reduction Rx.

Concept: 1 – In terms of gain or loss of oxygen or hydrogen.

Oxidation: The addition of oxygen to a substance or removal of hydrogen from a substance is called oxidation.

Reduction: The addition of Hydrogen to a substance or removal of oxygen from a substance is called reduction.

Oxidising agent:

The substance which is responsible for the addition of oxygen to substance or removal of Hydrogen from a substance is called oxidising agent.

Reducing agent:

The substance which is responsible for the addition of Hydrogen to substance or removal of Oxygen from a substance is called reducing agent.

Note: The oxidation and reduction reactions are called redox reactions

Ex: When copper is heated in air, it reacts with the oxygen to form black compound copper oxide:

$$\underset{\text{(Red brown)}}{2Cu} + O_2 \longrightarrow \underset{\text{(Black)}}{2CuO}$$

If hydrogen gas is passed over heated copperoxide, then the black copperoxide is reduced and red-brown copper metal is formed:

$$\underset{\text{(Black)}}{CuO} + H_2 \longrightarrow \underset{\text{(red brown)}}{Cu} + H_2O$$

In the above Rx oxygen is removed from copperoxide. So CuO is reduced and Hydrogen is responsible for removal of oxygen from copper oxidised. So Hydrogen is reducing agent on the other hand oxygen is added with Hydrogen. So, Hydrogen is oxidised and CuO is responsible for addition of oxygen with Hydrogen. So copper oxide is oxidising agent.

ACTIVITY 1.13

Aim: To demonstrate the addition of oxygen is oxidation and removal of oxygen is reduction.

Procedure:

Step 1 Fix a wide long glass tube horizontally on a stand.

Step 2 Introduce copper powder from one end of the long tube.

Step 3 Insert one cork with a narrow bore glass tube at one end of the wide bore glass tube.

Step 4 Insert another cork with a narrow bore glass tube at the other end of the wide bore glass tube.

Step 5 Heat the copper powder kept in the long tube with a burner.

Step 6 Pass air over heated copper powder.

Step 7 Record your observations.

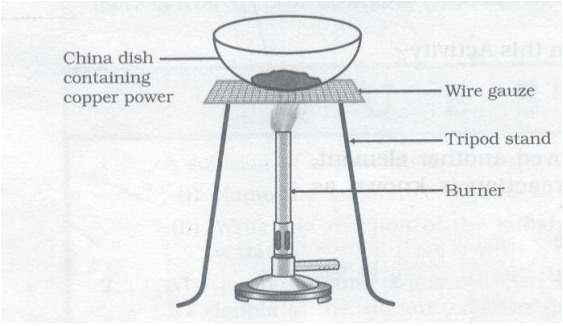
Observations: Copper powder is brown before heating. A black mass is formed when air is passed over heated copper powder.

(i) in above activity oxygen removed from CuO. So copperoxide is reduced to copper.

(ii) in above example oxygen is added to hydrogen. So hydrogen is being oxidised to water.

(iii) in above activity, copper oxide is giving the oxygen for the oxidation of hydrogen. So copper oxide is the oxidising agent.

(iv) in above activity, Hydrogen responsible for removing oxygen from copper oxide. So hydrogen is the reducing agent.



Important conclusions:

(i) substance oxidised: H_2

(ii) substance reduced: CuO

(iii) oxidising agent: CuO

(iv) Reducing agent: H_2

Ex.1 When Hydrogen sulphide reacts with chlorine, then sulphur and hydrogen chloride are formed:

$$H_2S + Cl_2 \longrightarrow S + 2HCl$$

(i) substance oxidised: H_2S

(ii) substance Reduced: Cl_2

(iii) oxidising agent: Cl_2

(iv) Reducing agent: H_2S

Concept: 2 – Oxidation and Reduction RX. On the basis of metals and non-metals.

Oxidation: The addition of non - metal with metal is called oxidation.

Reduction: The addition of metal with non - metal is called reduction.

Concept: 3 – Oxidation and Reduction Rx. On the basis of electrons.

Oxidation: The removal of electrons from an atom or ion is called oxidation.

Reduction: The addition of electrons to an atom or an ion is called reduction.

Remember the following points:

(a) The substance which get oxidised is the reducing agent.

(b) The substance which gets reduced is the oxidising agent

Examples of redox reactions in every day life

(1) Burning of fuel.

(2) Corrosion of metals.

(3) Rusting of iron.

(4) Respiration.

(5) Rancidity of food items.

Topic : Corrosion

Corrosion is the process in which metals are eaten up gradually by the action of air, moisture or a chemical on their surface.

Familiar Instances of Corrosion of Metals

(1) **Corrosion of Iron:** When an iron object exposed to moist air, then a Red brown coating is deposited on the surface of Iron object is called Rust.

$$4Fe + 3O_2 + x\,H_2O \longrightarrow \underset{\text{Rust}}{2Fe_2O_3.xH_2O}$$

(2) **Corrosion of Copper:** When a copper metal exposed to atmosphere, it react with the oxygen, water and carbon dioxide and form a green coating of basic copper carbon on the surface of copper metal.

$$\underset{\text{Copper}}{2Cu} + O_2 + CO_2 + H_2O \rightarrow \underset{(\text{Basic Copper Carbonate})}{CuCO_3\ Cu(OH)_2}$$

(3) **Corrosion of Silver:** When silver metal exposed to atmosphere it react with H_2S gas and gradually it becomes black due to formation of Ag_2S layer on the surface of silver metal.

$$Ag + H_2S \rightarrow \underset{(\text{Black Colour})}{Ag_2S} + H_2$$

(4) **Corrosion Aluminium:** When an aluminium metal exposed to atmosphere it react with the oxygen of air and formed a very thin and protective layer of aluminium oxide (Al_2O_3).

$$4Al + 3O_2 \longrightarrow 2Al_2O_3$$

Note: In the above example case corrosion is beneficial. This is because thin layer of aluminium oxide protect the aluminium metal from further corrosion.

Rusting of iron: When an iron object is left in damp air for a considerable time, if gets covered with a red - brown flaky substance called Rust.

$$4Fe + 3O_2 + xH_2O \longrightarrow \underset{\text{Rust (Red brown)}}{2Fe_2O_3.\ xH_2O}$$

Note: Corrosion weakens the iron and steel object and structure such as railings, car bodies, bridges and ships etc, and cuts short their life.

Prevention of Rusting:

(1) Rusting can be preventive by painting, oiling or greasing.

(2) Rusting can be prevented by galvanization. The depositing a lager of zinc metal over iron object is called electricity.

(3) Rusting can be prevented by making alloys

(4) Rusting can be prevented by electricity of tin or alluminium metal.

Rancidity: Destruction of food item containing fat or oil. By oxygen air and give unpleasant smell and tastes is called rancidity.

(i) Light, water, bacteria and molds.

Prevention of Rancidity:

(i) Rancidity can be prevented by adding anti oxidants to food containing facts and oils. There are two anti oxidents used which are

(i) **BHA**(Butylated Hydroxy Anisole) $C_{11}H_{16}O_2$

(ii) **BHT**(Butylated Hydroxy Toulene) $C_{15}H_{24}O$

(ii) Rancidity can be prevented by Packaging fat and oil containing foods in nitrogen gas. Because of nitrogen gas is unreactive gas.

(iii) Rancidity can be retarted by keeping food in refrigerator.

(iv) Rancidity can be retarded by storing food in air-tight container.

(v) Rancidity can be retarded by storing foods away from light.

Note:

ELECTRONINC CONFIGURATION OF FIRST 20 ELEMENTS

								Distribution of electrons				Valence electrons	Valance
S.No.	**Element**	**Symbol**	**At. No.**	**Mass No.**	**P**	**N**	**E**	**K**	**L**	**M**	**N**		
1.	**Hydrogen**	H	1	1	1	0	1	1				1	1
2.	**Helium**	He	2	4	2	2	2	2				2	2
3.	**Lithium**	Li	3	7	3	4	3	2	1			1	1
4.	**Beryllium**	Be	4	9	4	5	4	2	2			2	2
5.	**Boron**	B	5	11	5	6	5	2	3			3	3
6.	**Carbon**	C	6	12	6	6	6	2	4			4	4
7.	**Nitrogen**	N	7	14	7	7	7	2	5			5	3
8.	**Oxygen**	O	8	16	8	8	8	2	6			6	2
9.	**Fluorine**	F	9	19	9	10	9	2	7			7	1
10.	**Neon**	Ne	10	20	10	10	10	2	8			8	Ine
11.	**Sodium**	Na	11	23	11	12	11	2	8	1		1	1
12.	Magnesium	Mg	12	24	12	12	12	2	8	2		2	2
13.	**Aluminium**	Al	13	27	13	14	13	2	8	3		3	3
14.	**Silicon**	Si	14	28	14	14	14	2	8	4		4	4
15.	Phosphorus	P	15	31	15	16	15	2	8	5		5	3
16.	**Sulphur**	S	16	32	16	16	16	2	8	6		6	2
17.	**Chlorine**	Cl	17	35	17	18	17	2	8	7		7	1
18.	**Argon**	Ar	18	40	18	22	18	2	8	8		8	Ine
19.	**Potassium**	K	19	39	19	20	19	2	8	8	1	1	1
20.	**Calcium**	Ca	20	40	20	20	20	2	8	8	2	2	2