1. PHYSICAL AND CHEMICAL CHANGES

1.1 CHANGE IN MATTER

Change in matter can be studied in two major ways. The two types of changes are as follows:

TYPES OF CHANGES

Physical Changes
(No new substance is formed)

e.g. Melting of ice, boiling of water, breaking of a glass tumbler.

Chemical Changes
(A new substance is formed)

e.g. Souring of milk in summers, Rusting of iron articles, Burning of any substance, Digestion of food in our body.

1.2 DIFFERENCE BETWEEN PHYSICAL CHANGE AND CHEMICAL CHANGE

Physical Change	**Chemical Change**
(i) Those changes in which no new substances are formed are called physical changes.	(i) Those changes in which the original substances lose their chemical nature and identity and form new chemical substances with different properties are called chemical changes.
(ii) It is a temporary change	(ii) It is a permanent change.
(iii) It is easily reversible	(iii) It is usually irreversible.
(iv) In a physical change the mass of substance does not alter.	(iv) In a chemical change the mass of the substance does alter.

2. CHEMICAL REACTION

The process involving a chemical change is called a **chemical reaction**. **OR**

The process in which a substance or substances undergo change, to produce new substances with new properties, is known as **chemical reaction**.

Reactants: The substances which take part in a chemical reaction are called reactants.

Products: The new substances formed as a result of chemical reaction are called products.

For example

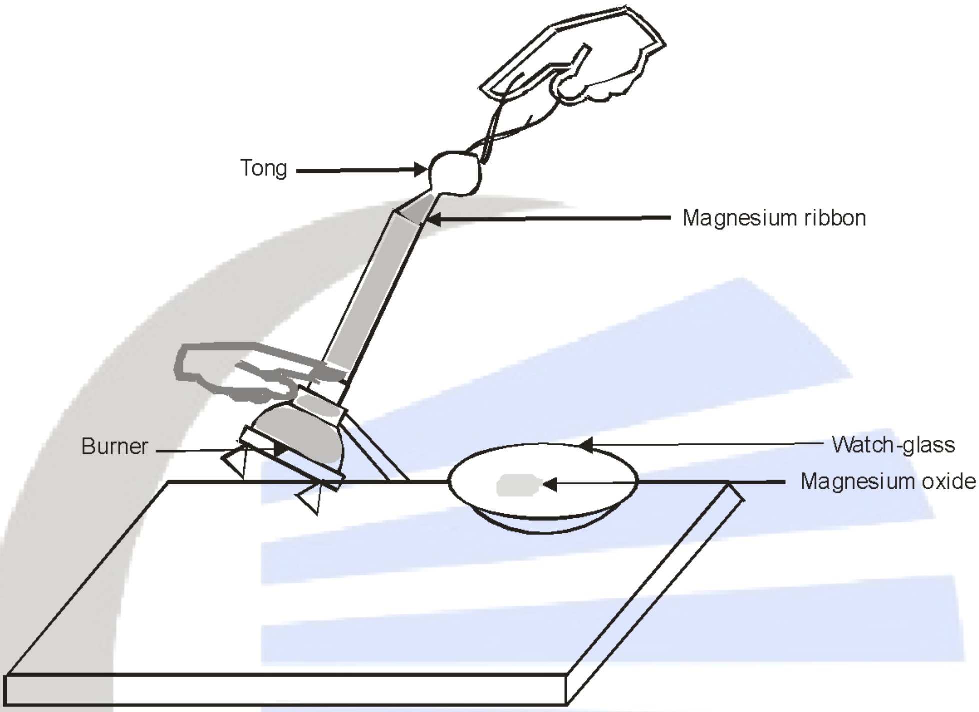
$$\underset{\text{(Reactant)}}{2H_2(g)} + \underset{\text{(Reactant)}}{O_2(g)} \longrightarrow \underset{\text{(Product)}}{2H_2O(l)}$$

In the above chemical reaction hydrogen and oxygen which are written on the left hand side are reactants and water which is written on the right hand side is a product.

2.1 EXAMPLES OF SOME CHEMICAL REACTIONS

(i) The burning of magnesium in air to form magnesium oxide

Take a magnesium ribbon and clear it by rubbing with a sand paper. Hold it with a pair of tongs. Burn it using a spirit lamp or burner and collect the white ash so formed in a watch-glass as shown in figure.



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

Figure 1

Magnesium ribbon burns with dazzling light and a white substance is formed which is magnesium oxide. This happens due to the following chemical reaction

$$\underset{\text{Magnesium}}{2\,Mg\,(s)} + \underset{\text{Oxygen (from air)}}{O_2\,(g)} \longrightarrow \underset{\text{Magnesium oxide}}{2\,MgO\,(s)}$$

Thus a chemical reaction has taken place in which magnesium has combined with oxygen of the air to form a new chemical substance, magnesium oxide (MgO). Here Mg and O_2 are reactants and MgO formed is product.

(ii) Reaction between lead nitrate and potassium iodide.

Take lead nitrate solution in a test tube and add some potassium iodide solution to this.

A yellow solid namely lead iodide in the form of precipitate appears. Another substance namely potassium nitrate is also formed which we cannot see as it remains in the solution. This happens due to the following chemical reaction.

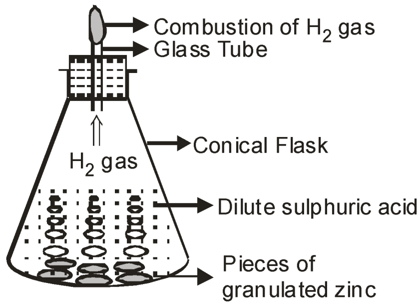
$$\underset{\text{Lead nitrate}}{Pb(NO_3)_2(aq)} + \underset{\text{Potassium Iodide}}{2KI\,(aq)} \rightarrow \underset{\substack{\text{Yellow ppt}\\\text{Lead Iodide}}}{PbI_2\,(s)} + \underset{\text{Pot. nitrate}}{2KNO_3(aq)}$$

In this reaction lead nitrate and potassium iodide are the reactants while lead iodide and potassium nitrate are the products.

(iii) Reaction between zinc and dilute sulphuric acid (or hydrochloric acid)

Take a few zinc granules in a conical flask and add some dilute hydrochloric acid (HCl) or sulphuric acid (H_2SO_4) to this. A gas is evolved very briskly. If we touch the flask, it is found to be hot.

On bringing a lighted candle near the upper end of the tube fitted in the flask (figure 2) the gas burns with a popping sound. This confirms that the gas evolved is hydrogen.



Reaction between zinc and dilute sulphuric acid (or hydrochloric acid) giving out hydrogen gas Figure 2

During this reaction, zinc sulphate or zinc chloride is also formed which we cannot see as it remains in the solution. The following chemical reaction takes place.

With HCl

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2 \uparrow$$

With H_2SO_4 :

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2 \uparrow$$

In this chemical reaction Zinc and hydrochloric acid (or sulphuric acid) are the reactants while hydrogen gas and zinc chloride (or zinc sulphate) are the products

2.2 CHARACTERISTICS OF CHEMICAL REACTIONS

The easily observable changes which take place as a result of chemical reactions are known as characteristics of chemical reactions.

CHARACTERISTICS OF CHEMICAL REACTIONS

- Change in state
- Change in colour
- Evolution of a gas
- Change in temperature

When a chemical reaction takes place, any one or more of the above changes or characteristics are observed. Let us discuss these characteristics in detail.

(i) **Change in state (formation of precipitate):** In the above discussed chemical reactions, in the burning of magnesium ribbon, **white powder of magnesium oxide** is formed. In the chemical reaction between lead nitrate and potassium iodide **yellow precipitate of lead iodide** is formed. Similarly when barium chloride solution reacts with sodium sulphate solution, a **white precipitate of barium sulphate is obtained.** Similarly in the chemical reaction of silver nitrate solution with sodium chloride solution, **white precipitate** of silver chloride is obtained.

$$\underset{\text{Lead nitrate}}{Pb(NO_3)_2(aq)} + \underset{\text{Potassium iodide}}{2KI(aq)} \rightarrow \underset{\text{Lead iodide (yellow ppt)}}{PbI_2(s)} + \underset{\text{Pot. nitrate}}{2KNO_3(aq)}$$

(ii) Change in colour : Some chemical reactions are characterized by change in colour

Example 1

When citric acid of citrus fruits (lemon and orange), reacts with potassium permagnate solution (purple), the colour changes and a colourless solution is produced.

$$\text{Citric acid} + \underset{\substack{\text{Pot. permagnate}\\ \text{(Purple)}}}{KMnO_4} \longrightarrow \text{Colourless solution}$$

As the purple colour of potassium permagnate solution changes to colourless by addition of citric acid (lemon juice) it is an example of CHANGE IN COLOUR.

Example 2

When sulphur dioxide gas reacts with acidified pot. dichromate solution, the orange colour of pot. dichromate changes to green.

$$\underset{\text{sulphur dioxide}}{SO_2} + \underset{\substack{\text{Pot. dichromate}\\ \text{(Orange)}}}{K_2Cr_2O_7} \longrightarrow \text{Green Colour}$$

Example 3

When ferrous sulphate crystals are heated, they lose their green colour and become white.

$$\underset{\text{Green}}{FeSO_4.7H_2O(s)} \xrightarrow{\text{Heat}} \underset{\text{White}}{FeSO_4(s)} + 7H_2O$$

(iii) Evolution of Gas: Some chemical reactions, when take place, a gas is evolved.

Example 1

In chemical reaction between zinc and dilute hydrochloric acid or dilute sulphuric acid, hydrogen gas is evolved.

i.e. $Zn + HCl \longrightarrow ZnCl_2 + H_2 \uparrow$

$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2 \uparrow$

Example 2

Metals like sodium and calcium react with water and hydrogen gas is evolved

$2Na + 2H_2O \longrightarrow 2\,NaOH + H_2 \uparrow$

$Ca + 2H_2O \longrightarrow Ca(OH)_2 + H_2 \uparrow$

Example 3

In the chemical reaction of calcium carbonate with dilute hydrochloric acid, carbon dioxide gas is evolved.

$CaCO_3 + 2HCl \longrightarrow CaCl_2 + CO_2 \uparrow + H_2O$

On heating calcium carbonate, carbon dioxide gas is evolved

Example 4

$$CaCO_3 \xrightarrow[\text{heat}]{} CaO + CO_2 \uparrow$$

(iv) Change in Temperature : Some chemical reactions are known by a change in temperature. This change in temperature may be rise in temperature or fall in temperature. For example

Example 1

When zinc pieces react with dilute hydrochloric acid or dilute sulphuric acid in a flask, it becomes hot (rise in temperature).

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2\uparrow + heat$$

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2\uparrow + heat$$

Example 2

In the reaction of quick lime with water, a lot of heat is produced. Which causes a rise in its temperature.

$$CaO + H_2O \longrightarrow Ca(OH)_2 + Heat.$$

Example 3

In the chemical reaction of barium hydroxide [$Ba(OH)_2$] with ammonium chloride (NH_4Cl) to form barium chloride ($BaCl_2$), ammonia (NH_3) and water, a lot of heat energy is absorbed due to which the temperature of the mixture falls and it becomes very cold.

$$Ba(OH)_2 + NH_4Cl + Heat \longrightarrow BaCl_2 + NH_3 + H_2O$$

3. CHEMICAL EQUATIONS

The short-hand method of representing a chemical reaction in terms of symbols and formulae of the different reactants and products is called a **chemical equation**.

A chemical reaction can be represented in two different ways :

REPRESENTATION OF CHEMICAL REACTIONS

Word Equation
e.g.
Zinc + Sulphuric acid → Zinc sulphate + Hydrogen (g)

Chemical Equation
$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2\,(g)$

3.1 STEPS FOR WRITING A CHEMICAL EQUATION

Writing of a chemical equation involves the following steps :

(i) The symbols and formulae of the reactants are written on the left hand side with plus (+) sign between them.

(ii) The symbols and formulae of the products are written on the right hand side with + sign between them.

(iii) An arrow sign (→) is put between the reactants and the products, pointing from reactants towards products.

3.2 BALANCED AND UN BALANCED CHEMICAL EQUATIONS

(i) Balanced chemical equation

The equation in which the number of atoms of each element in the reactants, and the products sides are equal, is called a balanced chemical equation.

The chemical equations are balanced to satisfy the **law of conservation of mass** in chemical reactions.

For example, in a chemical reaction between zinc and dilute sulphuric acid, giving zinc sulphate and hydrogen. The chemical equation can be written as

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2.$$

Here the No. of atoms of each element in the reactant and products sides are equal. i.e.

	In reactants	**In products**
No. of Zn atoms	1	1
No. of H atoms	2	2
No. of S atoms	1	1
No. of O atoms	4	4

Hence it is a balanced chemical equation.

(ii) Unbalanced chemical equation (skeletal equation)

The equation in which the number of atoms of different elements on the reactants and the product sides are not equal is called an unbalanced chemical equation. The unbalanced chemical equation is also known as **skeletal equation**.

For example

The burning of aluminium in oxygen to form aluminium oxide can be written as :

$$Al + O_2 \longrightarrow Al_2\,O_3$$

Here the No. of atoms of each element in the reactants and products side are not equal. i.e.

	In reactants	**In products**
No. of Al atoms	1	1
No. of O atoms	2	3

Hence it is an unbalanced chemical equation.

3.3 BALANCING A CHEMICAL EQUATION

Steps involved in the balancing of a chemical equation are as follows :

(i) Write the equation in the word form by keeping the reactants on the left hand side and the products on the right hand side of the arrow sign (→)

(ii) Write the symbols and formulae of all the reactants and the products in the word equation to get the skeletal chemical equation.

(iii) List the no. of atoms of different elements of reactants and products.

(iv) To balance the chemical equations, select the compound which has maximum number of atoms, irrespective of the fact whether it is a reactant or product. Multiply the symbols and formulae by the smallest possible number to balance the element having maximum atoms.

(v) Also balance the others elements one by one. Do not change the formulae to balance the chemical equation.

Example 1

Write the balanced chemical equation for the following reaction. Copper sulphate reacts with sodium hydroxide to form copper hydroxide and sodium sulphate.

(i) Writing the equation in the word-form.

$$\text{Copper Sulphate} + \text{Sodium Hydroxide} \longrightarrow \text{Copper Hydroxide} + \text{Sodium Sulphate}$$

(ii) Writing the skeletal chemical equation and enclosing the formulae in boxes.

$$\boxed{CuSO_4} + \boxed{NaOH} \longrightarrow \boxed{Cu(OH)_2} + \boxed{Na_2SO_4}$$

(iii)

Element	No. of atoms in LHS	No. of atoms in RHS
Cu	1	1
S	1	1
O	5	6
Na	1	2
H	1	2

(iv) Selecting the biggest formula (i.e. Na_2SO_4) and balancing the element with highest number of atoms i.e. oxygen.

There are 5 atoms of oxygen in LHS but 6 in RHS. To balance oxygen multiply NaOH by 2.

$$\boxed{CuSO_4} + 2\ \boxed{NaOH} \longrightarrow \boxed{Cu(OH)_2} + \boxed{Na_2SO_4}$$

By doing so sodium and hydrogen also get balanced.

(v) Checking the correctness of the balanced equation.

Element	In Reactants	In Products
Cu	1	1
S	1	1
O	6	6
Na	2	2
H	2	2

Hence the final balanced chemical equation may be written as

$$CuSO_4 + 2NaOH \longrightarrow Cu(OH)_2 + Na_2SO_4$$

3.4 LIMITATIONS OF A CHEMICAL EQUATION

It gives us no information about the following :

(i) The physical state of the reactants.

(ii) The concentration of the reactants.

(iii) The time taken for the reaction to complete.

(iv) The rate at which the reaction proceeds.

(v) The conditions necessary to start and carry on the reaction *e.g.,* Is any catalyst required? What is the temperature needed to start and continue the reaction?

(vi) Is the reaction exothermic or endothermic, *i.e,* is heat evolved or absorbed during the reaction?

3.5 ESSENTIALS OF A CHEMICAL EQUATION

A true chemical equation, therefore, must be in accordance with the following essentials :

(i) It should represent an actual chemical change.

(ii) It should be balanced, *i.e.,* number of atoms of different elements on the two sides of the equation must be equal.

(iii) It should be molecular *i.e.,* all the substances concerned should be expressed as molecules.

3.6 TO MAKE EQUATION MORE INFORMATIVE

It is quite helpful if an equation gives an idea about the physical state, heat changes and the conditions under which the reaction takes place.

This can be done in the following three ways.

(i) To indicate the physical states of reactants and products : By indicating the physical states of the reactants and products of a chemical reactions by writing letters (g) for gaseous state, (l) for liquid state, (s) for solid state and (aq) for aqueous solutions i.e. solutions in water, just after the formulae in an equation as shown in the examples below.

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

An arrow pointing upwards (↑) may also be used to indicate a gaseous product in the equation as follows :

$Zn\ (s) + H_2SO_4(aq) \rightarrow Zn\ SO_4(aq) + H_2 \uparrow$.

An arrow pointing downwards (↓) is used to indicate an insoluble product or precipitate (ppt) in an equation as follows :

$AgNO_3$ (aq)	+ NaCl (aq) →	AgCl(s) ↓ +	$NaNO_3$(aq)
Silver Nitrate	Sodium chloride	Silver chloride Precipitate	Sodium Nitrate

(ii) To indicate the heat changes (thermo chemical equations) : Chemical equations for exothermic or endothermic reactions representing the heat evolved or absorbed during the reaction are called thermo chemical equations.

(a) An exothermic reaction is indicated by writing + heat, or + heat energy or + energy, on the products side of an equation as follows :

$C(s) + O_2(g) \rightarrow CO_2(g) + heat$

$CaO\ (s) + H_2O(l) \rightarrow Ca(OH)_2 + heat$

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g) + heat$

(b) An endothermic reaction is indicated by writing "+ heat" or "+ heat energy" or "+ energy" on the reactants side of an equation as shown below.

$CaCO_3(s)$	+	heat	→	$CaO(s)$	+	$CO_2(g)$
Calcium carbonate				Calcium oxide		Carbondioxide

$N_2(g)$	+	$O_2(g)$	+	heat	→	$2NO(g)$
Nitrogen		Oxygen				Nitric oxide

$Ba(OH)_2 + 2\ NH_4Cl + heat \rightarrow BaCl_2 + 2\ NH_4OH$

Barium hydroxide Ammonium chloride Barium chloride Ammonium hydroxide

(iii) **To indicate the conditions under which the reaction takes place:** The conditions of temperature, pressure and the presence of catalyst if any are represented by writing these conditions above or below the arrow sign as follows:

$$\underset{\text{Nitrogen}}{N_2(g)} + \underset{\text{Hidrogen}}{3H_2(g)} \xrightarrow[\text{Fe}]{500°C,\ 200\,atm} \underset{\text{Ammonia}}{2NH_3(g)}$$

Here 500°C temperature, 200 atm pressure and Fe as catalyst are the condition for the reaction to take place.

$$\underset{\text{Carbon monooxide}}{CO(g)} + \underset{\text{Hydrogen}}{2H_2(g)} \xrightarrow[ZnO+CrO_3]{300°C,\ 300atm} \underset{\text{Methanol}}{CH_3OH(l)}$$

Here 300°C temperature, 300 atm pressure and zinc acid (ZnO) and chromium oxide CrO_3 as catalyst are the conditions needed for the reaction to take place.

If heat is required for the reaction to take place, then heat sign delta (Δ) is written over the arrow of the equation as follows :

$$\underset{\text{Potassium Chlorate}}{2KClO_3(s)} \xrightarrow[MnO_2]{\Delta} \underset{\text{Potassium Chloride}}{2KCl(s)} + \underset{\text{Oxygen}}{3O_2(g)}$$

Here Δ shows the heat needed and MnO_2 is the catalyst needed for the reaction to take place.

4. TYPES OF CHEMICAL REACTIONS

TYPES OF CHEMICAL REACTIONS

- Combination reactions
- Decomposition reactions
- Displacement reactions
- Double displacement reactions
- Oxidation – Reduction (Redox) Reactions

4.1 COMBINATION REACTIONS

Those reactions in which two or more substances (reactants) combine together to form a single substance (product) are called the combination reactions.

Example 1 : Formation of water from H_2(g) and O_2(g)

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

In this reaction, two substances hydrogen and oxygen (reactants) combine together to form a single substance i.e. water (product). So it is a **combination reaction**.

Example 2 :

$$\underset{\text{Carbon (Coal)}}{2CO\ (g)} + \underset{\text{Oxygen}}{O_2(g)} \rightarrow \underset{\text{Carbondioxide}}{2CO_2\ (g)}$$

In this reaction again two substances namely, carbon monoxide and oxygen combine together to form a single substance, carbon dioxide (product). Thus it is also a **combination reaction**.

Example 3 : Reaction of ammonia and hydrogen chloride

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

This reaction is a combination reaction, as ammonia and Hydrogen chloride combine together to form ammonium chloride as a single product.

Example 4 : Reaction of water on quick lime :

Experiment : If we take a small amount of quick lime (calcium oxide) in a beaker and add some water to it slowly, then they combine vigorously to form slaked lime i.e. (Calcium hydroxide).

The reaction may be represented by :

$$\underset{\text{Quick lime}}{CaO(s)} + \underset{\text{water}}{H_2O(l)} \rightarrow \underset{\substack{\text{Slaked lime} \\ \text{(Calcium hydroxide)}}}{Ca(OH)_2(aq)}$$

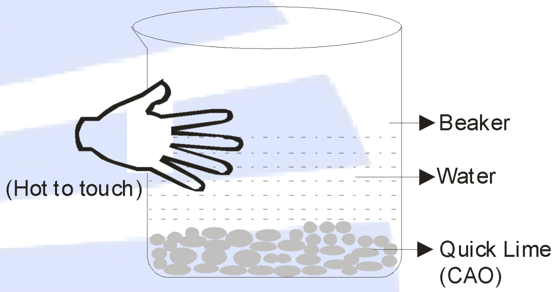


Figure 3

In this reaction, two substances (reactants), namely, calcium oxide and water have combined to form a single substance i.e., calcium hydroxide. Thus, it is a combination reaction.

Since a lot of heat is produced along with slaked lime so it is also called an exothermic reaction.

4.2 DECOMPOSITION REACTIONS

Those reactions in which a single substance (reactant) splits up into two or more simpler substances (products) are known as decomposition reactions.

These reactions are carried out by supplying energy in form of heat, electricity or light which breaks that substance into simpler substances. Thus decomposition reactions are classified as:

DECOMPOSITION REACTION

Thermolysis | Electrolysis | Photolysis

(i) Thermolysis or thermal decomposition reactions (decomposition by heat).

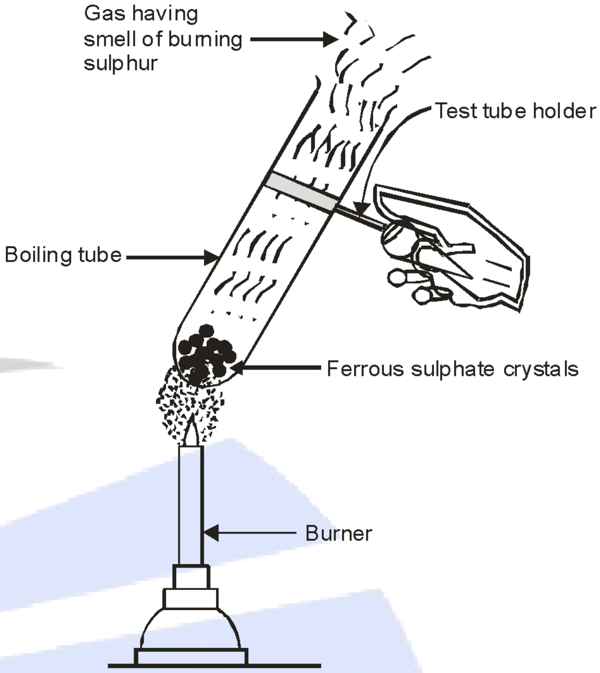
(ii) Electrolysis or electrolytic decomposition reactions (decomposition by electricity)

(iii) Photolysis or photodecomposition reactions (decomposition by light).

(i) Thermal decomposition reactions : Decomposition reactions which are carried out by heating are called thermal decomposition reactions.

Example 1 : Thermal decomposition of ferrous sulphate

Experiment : Take a small amount of ferrous sulphate crystals ($FeSO_4 . 7H_2O$) in a test tube. These crystals are green in colour. Heat the test tube over a burner. The green colour of crystals first changes to white due to formation of Anhydrous Ferrous sulphate ($FeSO_4$), as on heating, ferrous sulphate crystals lose the water molecules. On further heating anhydrous ferrous sulphate decomposes to brown solid (which is ferric oxide) along with characteristic smell of burning sulphur (fig. 4).



Decomposion reaction of ferrous sulphate

Figure 4

The reactions involved are :

$$\underset{\substack{\text{Ferrous Sulphate Crystals} \\ \text{(Green)}}}{FeSO_4.7H_2O(s)} \xrightarrow{\text{Heat}} \underset{\substack{\text{Ferrous Sulphate} \\ \text{(white)}}}{FeSO_4(s)} + \underset{\text{Water vapour}}{7H_2O(g)}$$

$$\underset{\substack{\text{Ferrous Sulphate} \\ \text{(white)}}}{2FeSO_4(s)} \xrightarrow{\text{Heat}} \underset{\substack{\text{Ferric oxide} \\ \text{(Brown colour)}}}{Fe_2O_3(s)} + \underset{\text{Sulphur dioxide}}{SO_2(g)\uparrow} + \underset{\text{Sulphur trioxide}}{SO_3(g)\uparrow}$$

In this Reaction, A single substance splits up into three simpler substances. Thus it is a decomposition reaction : Since this decomposition is brought about by heat, so, it is an example of **thermal decomposition reaction**.

Example 2 : Thermal decomposition of Potassium Chlorate : When potassium chlorate is heated in the presence of manganese dioxide as a catalyst, it decomposes to give potassium chloride and oxygen.

$$\underset{\text{Potassium Chlorate}}{2KClO_3(s)} \xrightarrow[(MnO_2)]{\text{Heat}} \underset{\text{Potassium Chloride}}{2KCl(s)} + \underset{\text{Oxygen}}{3O_2(g)}$$

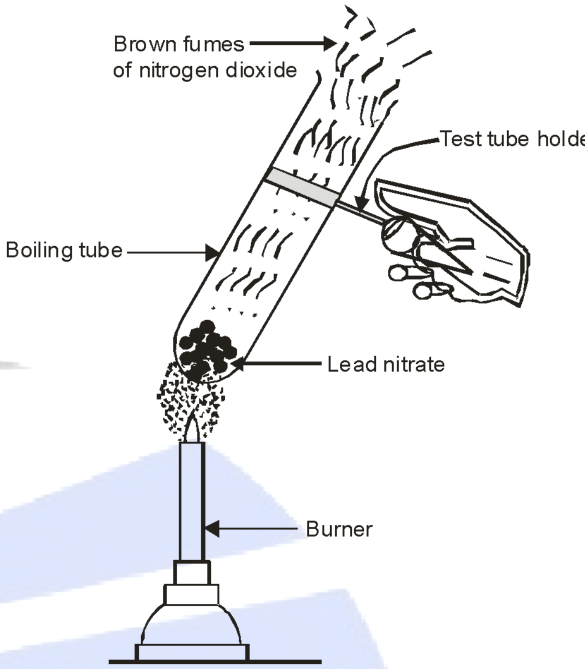
In this reaction, a single substance splits up into two simpler substances on heating. Thus it is a **thermal decomposition reaction.**

Example 3 : Thermal decomposition of Lead nitrate

Experiment: Take about 2 gm of powdered lead nitrate (colourless) in a dry test tube. Heat the tube over a burner. We observe that brown fumes of nitrogen dioxide are found to evolve and a yellow residue of lead oxide is left behind in the tube (fig. 5). If we put a burning candle over the mouth of test tube, it catches fire and starts burning again. This shows that the oxygen gas is also evolved during this reaction.

The reaction involved is :

$$\underset{\substack{\text{Lead Nitrate}\\\text{(Colourless)}}}{2\,Pb(NO_3)_2(s)} \xrightarrow{\text{Heat}} \underset{\text{Lead oxide (yellow)}}{2\,PbO(s)} + \underset{\substack{\text{Nitrogen dioxide}\\\text{(Brown fumes)}}}{4NO_2(g)} + \underset{\text{Oxygen}}{O_2(g)}$$



Decomposition reaction of lead nitrate
Figure 5

In this reaction a single substance i.e. lead nitrate decomposes to three simpler substances (i.e. lead oxide, nitrogen dioxide and oxygen).

Thus it is a decomposition reaction, Since this decomposition is carried about by heat, therefore, it is an example of **thermal decomposition reaction.**

Example 4: Thermal Decomposition of limestone: When calcium carbonate (limestone) is heated, it decomposes to give calcium oxide and carbon dioxide:

$$\underset{\substack{\text{Calcium Carbonate}\\\text{(Limestone)}}}{CaCO_3(s)} \xrightarrow{\text{Heat}} \underset{\text{Calciumoxide (Lime)}}{CaO(s)} + \underset{\text{Carbondioxide}}{CO_2(g)}$$

In this reaction, a single substance, limestone breaks up into two simpler substances, calcium oxide and carbon dioxide, so it is a decomposition reaction. Since the decomposition is carried out by heating, so it is an example of thermal decomposition reaction.

Example 5 : Thermal decomposition of zinc carbonate : When zinc carbonate is heated, it decomposes to give zinc oxide and carbon dioxide gas :

$$\underset{\text{Zinc Carbonate}}{ZnCO_3} \xrightarrow{\text{Heat}} \underset{\text{Zinc oxide}}{ZnO(s)} + \underset{\text{Carbon dioxide}}{CO_2(g)}$$

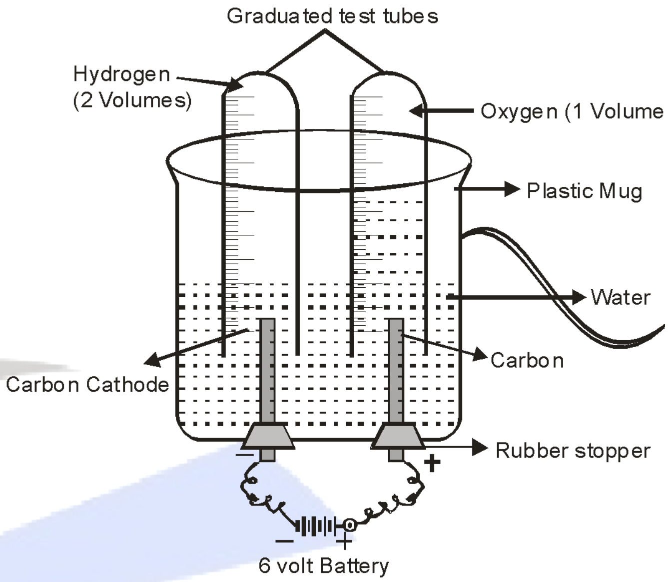
In this reaction zinc carbonate on heating decomposes to give two simpler substances, zinc oxide and carbon dioxide, thus it is an example of **thermal decomposition reaction.**

(ii) Electrolytic decomposition Reactions (or electrolysis)

The decomposition reactions which are carried out by using electricity, are called electrolytic decomposition reactions (or electrolysis)

Example 1: Electrolytic decomposition of water or electrolysis of water

Experiment: Take a plastic mug and make two holes at the base of mug. Fit two rubber stoppers in these holes, with carbon electrodes. Connect the electrodes to a 6-V battery (fig. 6). Fill the mug with water so that electrodes get immersed in it. Add few drops of dil H_2SO_4 to make it good conductor of electricity. Fill the two test tubes with water and invert them over the electrodes. Now switch on the battery and keep the apparatus undisturbed for some time.



Experiment set-up for electrolysis of water
Figure 6

OBSERVATION

1. There is formation of bubbles of two different gases over the electrodes which displaces the water in test tubes.
2. The gas collected in the test tube covering the cathode is double than that of gas collected in tube covering the anode.
   Keep on passing current till the test tubes are completely filled with respective gases. Now switch off the battery.

RESULT

On testing the gases by bringing a burning candle near mouth of tubes, it is found that the gas with double volume, burns with poping sound, so it is hydrogen gas. The other gas with lesser volume, makes the burning candle to glow more, so, it is oxygen gas.

Thus above experiment shows that on supplying electricity, water decomposes into hydrogen and oxygen according to the reaction :

$$\underset{\text{Water}}{2H_2O(l)} \xrightarrow[\text{current}]{\text{electric}} \underset{\text{hydrogen (2 Volume)}}{2H_2(g)} + \underset{\text{oxygen (1volume)}}{O_2(g)}$$

Since in this reaction decomposition is carried out by using electricity, so it is an example of **electrolytic decomposition reaction** or **electrolysis**.

Example 2: Electrolytic decomposition of molten sodium chloride : On passing electric current through molten sodium chloride, it decomposes to give sodium metal and chlorine gas :

$$\underset{\text{Sodium Chloride}}{2\,NaCl(l)} \xrightarrow[\text{current}]{\text{electric}} \underset{\text{Sodium metal}}{2Na(s)} + \underset{\text{Chlorine gas}}{Cl_2(g)}$$

In this reaction on passing electric current a single substance decomposes to give two simpler substances. Thus it an example of **electrolytic decomposition reaction.**

Example 3: Electrolytic decomposition of molten alumina (aluminium oxide) : On passing electric current through molten alumina, it decomposes to give aluminium metal and oxygen gas :

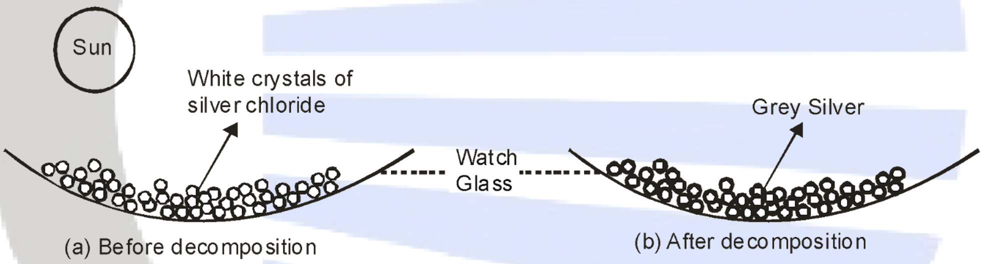
$$\underset{\text{(Alumina molten)}}{2\,Al_2O_3(l)} \xrightarrow[\text{current}]{\text{electric}} \underset{\text{Aluminium metal}}{4\,Al(s)} + \underset{\text{Oxygen gas}}{3O_2(g)}$$

In this reaction, on passing electric current, a single substance i.e. alumina decomposes to give two simpler substances, aluminium metal and oxygen gas, thus it is an example of **electrolytic decomposition reaction**.

(iii) **Photo-decomposition reactions (or photolysis) :** The decomposition reactions which take place by absorption of light are called the photo-decomposition reactions or photolysis.

Example 1: Photo-decomposition of silver chloride or Photolysis of Silver chloride

Experiment: Take a pinch of silver chloride on a watch glass and keep it in sunlight for some time.



Decomposition of silver chloride in the presence of light

Figure 7

It is observed that white silver chloride turns gray due to formation of silver metal.

$$\underset{\substack{\text{(Silver chloride)} \\ \text{(White)}}}{2\,AgCl(s)} \xrightarrow[\text{current}]{\text{Sunlight}} \underset{\substack{\text{Silver (Grey)} \\ \text{Metal}}}{2\,Ag(s)} + \underset{\text{Chlorine}}{Cl_2(g)}$$

In this reaction a single substance i.e. silver chloride decomposes (in pressure of sunlight) into two simpler substances, silver metal and chlorine gas. Thus this is an example of **photo-decomposition** reaction or **photolysis**.

Example 2: Photolysis of hydrogen iodide: Hydrogen iodide decomposes in the presence of ultraviolet light into hydrogen and iodine :

$$\underset{\text{Hydrogen iodide}}{2HI(s)} \xrightarrow{\text{UV light}} \underset{\text{hydrogen}}{H_2(g)} + \underset{\text{iodine}}{I_2(g)}$$

In this reaction single substance *i.e.*, hydrogen iodide on absorption of ultraviolet light decomposes into two simpler substances, hydrogen and iodine. Thus it is an example of **photo-decomposition reaction or photolysis.**

Example 3: Photolysis of hydrogen peroxide : In presence of light, hydrogen peroxide decomposes into water and oxygen.

$$\underset{\text{Hydrogen peroxide}}{2H_2O_2(l)} \xrightarrow{\text{Light}} \underset{\text{Water}}{2H_2O(l)} + \underset{\text{Oxygen}}{O_2(g)}$$

In this reaction, a single substance decomposes in the presence of light, into two simpler substances. Thus it is an example of **photo-decomposition reaction.**

4.2.1 Decomposition reactions are called the opposite of combination reactions

In a decomposition reaction, one substance breaks up into two or more chemical substances, while in a combination reaction two or more substances combine to form one single substance. So these two reactions are called opposite of each other.

4.2.2 Uses of Decomposition Reactions

The decomposition reactions are used in the extraction of metals in the following ways :

(i) Metals are extracted from their molten salts by electrolytic decomposition e.g. sodium from molten sodium chloride and aluminium from alumina (molten aluminium oxide).

(ii) Thermal decomposition reactions form one of the steps in extraction of metals.
For example, Zinc carbonate (the naturally occurring ore of zinc) is first decomposed to give zinc oxide and then reduced to obtain zinc metal i.e.,

$$ZnCO_3(s) \xrightarrow[\text{Absence of air}]{\text{heat}} ZnO(s) + CO_2(g)$$

$$ZnO(s) + \underset{\text{Coke}}{C(s)} \xrightarrow[\text{Reduction}]{\text{heat}} Zn(s) + CO(g)$$

4.2.3 Decomposition Reactions in our body

The digestion of food in the body is an example of decomposition reaction. When we eat foods like wheat, rice or potatoes, then the starch present in them decomposes to give simple sugars like glucose in the body and proteins decompose to form amino acids :

$$\underset{\text{starch}}{(C_6H_{10}O_5)_n} + H_2O \xrightarrow{\text{diastase}} \underset{\text{maltose}}{C_{12}O_{22}O_{11}}$$

$$\underset{\text{maltose}}{C_{12}H_{22}O_{11}} + H_2O \xrightarrow{\text{maltase}} \underset{\text{Glucose}}{2C_6H_{12}O_6}$$

$$\text{Proteins} \xrightarrow{\text{Pepsin}} \text{Amino acids}$$

4.2.4 Decomposition Reactions are endothermic reactions

All decomposition reactions require energy either in form of heat, light or electricity. Hence all decomposition reactions are endothermic (heat absorbing) reactions.

4.3 DISPLACEMENT REACTIONS

Those reactions in which more active element displaces a less active element from its compound are called displacement reactions.

4.3.1 Relative activities or reactivities of metals

Metals have been arranged in decreasing order of their activities (or reactivities) in the activity series as follows:

Metals More Reactive than Hydrogen

Element	Symbol	
Potassium	K	⟵ Most Reactive
Sodium	Na	
Barium	Ba	
Calcium	Ca	
Magnesium	Mg	
Aluminium	Al	
Zinc	Zn	
Iron	Fe	
Nickel	Ni	
Tin	Sn	
Lead	Pb	
Hydrogen	H	
Copper	Cu	
Mercury	Hg	
Silver	Ag	
Gold	Au	
Platinum	Pt	⟵ Least Reactive

Metals less Reactive than Hydrogen

Reactivity decreases downward

Activity series of metals

It is clear from the series that the metals lying above the hydrogen are more reactive than the metals lying below the hydrogen. Thus any metal can displace the metals lying below it from its solution.

4.3.2 Relative activities (or reactivities) of Non-metals

Relative activities of non-metals like halogens is in the order:

$$F > Cl > Br > I$$

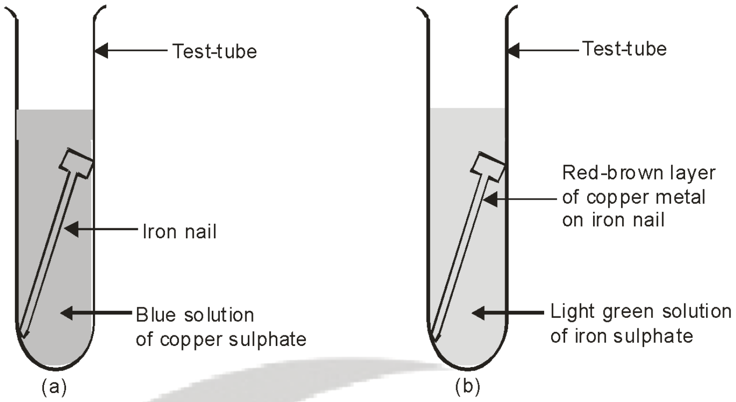
Thus, fluorine is most reactive and iodine is least reactive. So fluorine (F_2) can displace chlorine (Cl_2) from NaCl, Bromine (Br_2) from NaBr and so on. Similarly, chlorine can displace bromine (Br_2) from KBr and iodine (I_2) from KI and so on.

Let us discuss some displacement reactions on the basis of reactivity of metals and non-metals.

Example 1: Reaction between iron and copper sulphate solution :

Experiment : Take about 10 ml of copper sulphate solution in a test tube. It is deep blue in colour. Take two iron nails and clean their surface by rubbing with a sand paper.

Now put the one nail in test tube containing $CuSO_4$ solution. Keep another iron nail aside for comparison.



Displacement reaction between iron (nail) and copper sulphate solution.

Figure 8

After 30 minutes we observe the following changes :

(i) Blue colour of $CuSO_4$ has faded and it changes into light green due to formation of iron sulphate ($FeSO_4$)

(ii) The iron nail is covered with a red brown layer of copper metal.

These changes show that the following reaction has taken place :

$$\underset{\text{iron}}{Fe(s)} + \underset{\substack{\text{Copper Sulphate}\\ \text{(Blue solution)}}}{CuSO_4(aq)} \rightarrow \underset{\substack{\text{iron sulphate}\\ \text{(light green solution)}}}{FeSO_4(aq)} + \underset{\substack{\text{Copper}\\ \text{(Brownish deposit on iron)}}}{Cu(S)}$$

Thus, more reactive metal iron, (Fe), displaces the less reactive metal copper from copper sulphate solution, so this is an example of **displacement reaction**.

Example 1 : Displacement reactions in which a more reactive metal displaces a less reactive metal from its compound :

(i)

$$\underset{\text{Zinc}}{Zn(s)} + \underset{\substack{\text{Copper Sulphate}\\ \text{(Blue)}}}{CuSO_4(aq)} \rightarrow \underset{\substack{\text{Zinc Sulphate}\\ \text{(Colourless)}}}{ZnSO_4(aq)} + \underset{\text{Copper (Raddish brown)}}{Cu(s)}$$

Thus, when zinc pieces are added to copper sulphate solution, then, zinc being more reactive metal than copper, displaces copper from its solution ($CuSO_4$) so that Cu is set free. The blue colour of $CuSO_4$ solution fades due to formation of $ZnSO_4$ (colourless). A reddish brown deposit of copper metal is formed on the surface of zinc. Therefore it is an example of **displacement reaction**.

(ii)

$$\underset{\text{Copper}}{Cu(s)} + \underset{\substack{\text{Silver nitrate}\\ \text{(Colourless)}}}{2\,AgNO_3(aq)} \rightarrow \underset{\substack{\text{Copper nitrate}\\ \text{(Blue)}}}{Cu(NO_3)_2(aq)} + \underset{\text{Silver}}{2\,Ag(s)}$$

Thus, when a copper wire is dipped in silver nitrate solution, copper, being more reactive metal than silver, displaces silver from its solution ($AgNO_3$) so that silver is liberated. This silver is deposited on the copper wire giving it a white shining surface. The solution forms a blue colour due to formation of copper nitrate. Thus it is an example of **displacement reaction**.

(iii)

$$\underset{\text{Sodium}}{2Na(s)} + \underset{\text{water}}{2\,H_2O(l)} \rightarrow \underset{\text{Sodium hydroxide}}{2\,NaOH(aq)} + \underset{\text{Hydrogen}}{H_2(g)\uparrow}$$

In this reaction sodium, being more active than hydrogen, displaces hydrogen from water so that hydrogen gas is liberated along with the formation of sodium hydroxide. Thus, it is an example of **displacement reaction**.

Example 2 : The displacement reactions in which more reactive non-metal displaces less reactive non-metal from its compound :

$$\underset{\text{Chlorine}}{Cl_2(g)} + \underset{\text{Pot. bromide}}{2KBr(aq)} \rightarrow \underset{\text{Pot. Chloride}}{2KCl(aq)} + \underset{\text{Bromine (Raddish Brown)}}{Br_2(g)}$$

Thus, when Cl_2 gas is passed through an aqueous solution of potassium bromide, Chlorine being more reactive than bromine, displaces bromine from KBr so that bromine gas (Br_2) is liberated. The solution forms light brown colour due to dissolution of Br_2 gas in it. So it is an example of **displacement reaction**.

$$\underset{\text{Chlorine}}{Cl_2(g)} + \underset{\text{Pot. iodide}}{2KI(aq)} \rightarrow \underset{\text{Pot. Chloride}}{2KCl(aq)} + \underset{\text{Iodine (voilet)}}{I_2(s)}$$

Thus, when chlorine gas is passed through potassium iodide solution, chlorine, being more reactive than iodine, displaces iodine from KI and liberates I_2 gas.

The solution acquires violet colour due to dissolution of I_2 gas in it. So it is also an example of **displacement reaction**.

4.3.3 Uses of Displacement Reactions

Displacement reactions are used in the extraction of silver and gold. Silver or gold ore is dissolved in sodium cyanide solution. When zinc granules are added to the solution of the compound formed, zinc, being more active than silver and gold, displaces silver and gold from the solution of their compounds and thus silver and gold are extracted.

4.3.4 Displacement reactions are exothermic

All displacement reactions are exothermic (heat producing) reactions, For example :

(i) In the displacement reaction between zinc and dilute hydrochloric acid or dilute sulphuric acid, there is production of heat along with evolution of gas.

(ii) In displacement reaction between iron and copper sulphate solution, there is an increase in temperature due to production of heat.

Thus displacement reactions are heat producing or exothermic reactions.

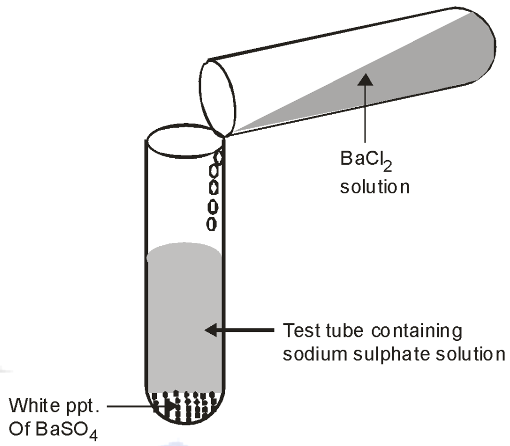
4.4 DOUBLE DISPLACEMENT REACTIONS

Those reactions in which two different atoms or groups of atoms are exchanged are called the **double displacement reactions.**

Example 1 : Reaction between Barium Chloride solution and sodium sulphate solution.

Experiment: Take sodium sulphate solution in a test tube and add few drops of solution of barium chloride to it and mix the two solution. There is formation of white precipitate of Barium sulphate: The reaction is as follows :

$$\underset{\text{Sodium Sulphate Sol.}}{Na_2SO_4(aq)} + \underset{\text{Barium Chloride Sol.}}{BaCl_2(aq)} \longrightarrow \underset{\text{Barium Sulphate (white ppt)}}{BaSO_4(s)} + \underset{\text{Sodium Chloride}}{2NaCl(aq)}$$



Formation of white ppt. of barium sulphate on adding barium chloride solution to sodium sulphate

Figure 10

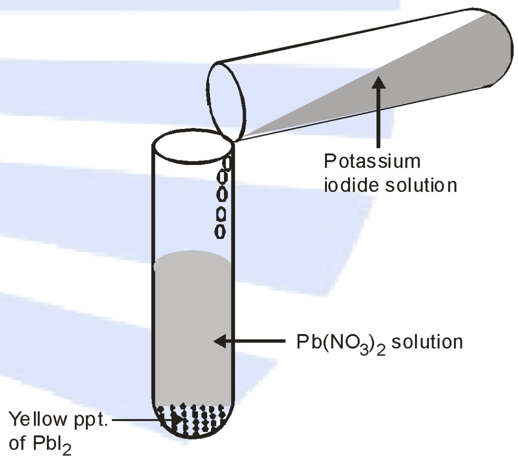
In this reaction, there is double displacement or exchange of ions, i.e. chloride ions of $BaCl_2$ have been replaced by sulphate ions of Na_2SO_4 whereas sodium ions of Na_2SO_4 have been replaced by chloride ions of $BaCl_2$. Hence it is a double displacement reaction.

Example 2: Reaction between lead nitrate solution and potassium Iodide solution.

Experiment : Take lead nitrate solution in a test tube and add potassium Iodide solution to it and mix the two solutions. There is formation of yellow precipitate of lead iodide (fig. 11).

The reaction involved is :

$$\underset{\text{Lead nitrate}}{Pb(NO_3)_2(aq)} + \underset{\text{Potassium iodide}}{2KI(aq)} \longrightarrow \underset{\text{Lead iodide (yellow ppt)}}{PbI_2(s)} + \underset{\text{Pot. nitrate}}{2KNO_3(aq)}$$



Formation of yellow ppt. of lead iodide on adding potassium iodide solution to lead nitrate

Figure 11

In this reaction, there is double displacement or exchange of ions, i.e. lead ions of $Pb(NO_3)_2$ have been replaced by potassium ions of KI and iodide ions of KI have been replaced by nitrate ions of $Pb(NO_3)_2$. Thus it is a **double displacement reaction**.

Some more examples of **double displacement reaction** are :

(i) $$\underset{\text{Copper Sulphate}}{CuSO_4(aq)} + \underset{\text{Amm. hydroxide}}{2NH_4OH(aq)} \rightarrow \underset{\substack{\text{Copper hydroxide}\\\text{(Bluish white ppt)}}}{Cu(OH)_2(s)\downarrow} + \underset{\text{Amm. sulphate}}{(NH_4)_2SO_4(aq)}$$

(ii) $$\underset{\text{Aluminium Chloride}}{AlCl_3(aq)} + \underset{\text{Amm. hydroxide}}{3NH_4OH(aq)} \rightarrow \underset{\substack{\text{Aluminium hydroxide}\\\text{(White ppt)}}}{Al(OH)_3(s)\downarrow} + \underset{\text{Amm. Chloride}}{3NH_4Cl(aq)}$$

(iii) $$\underset{\text{Copper Sulphate}}{CuSO_4(aq)} + \underset{\text{Hydrogen sulphide}}{H_2S(aq)} \rightarrow \underset{\substack{\text{Copper Sulphide}\\\text{(White ppt)}}}{CuS(s)\downarrow} + \underset{\text{Sulphuric Acid}}{H_2SO_4(aq)}$$

All the above reactions are double displacement reactions, as there is double exchange or displacement of ions.

4.5 OXIDATION – REDUCTION REACTIONS OR REDOX REACTIONS

4.5.1 Oxidation

Oxidation can be defined in three ways :

(i) Oxidation in terms of oxygen or electronegative element :

Oxidation is a process in which oxygen or electronegative element is added to a substance for example:

(a) $$2Mg + O_2 \rightarrow \underset{\text{magnesium oxide}}{2MgO} \quad \text{(Addition of oxygen)}$$

(b) $$4Na + O_2 \rightarrow \underset{\text{Sodium oxide}}{2Na_2O} \quad \text{(Addition of oxygen)}$$

(c) $$2Cu + O_2 \rightarrow \underset{\text{Copper (ii) oxide}}{2CuO} \quad \text{(Addition of oxygen)}$$

(d) $$S + O_2 \rightarrow \underset{\text{Sulphur dioxide}}{SO_2} \quad \text{(Addition of oxygen)}$$

(e) $$2H_2 + O_2 \rightarrow \underset{\text{water}}{2H_2O} \quad \text{(Addition of oxygen)}$$

(f) $$\underset{\text{iron}}{2Fe} + \underset{\substack{\text{Chlorine}\\\text{(electronegative)}\\\text{element}}}{3Cl_2} \rightarrow \underset{\text{Ferric chloride}}{2FeCl_3} \quad \text{(Addition of electronegative element)}$$

(g) $$\underset{\text{iron}}{Fe} + \underset{\substack{\text{Sulphur}\\\text{(electronegative)}\\\text{elemement}}}{S} \rightarrow \underset{\text{Ferrous sulphide}}{FeS} \quad \text{(Addition of electronegative element)}$$

(ii) Oxidation in terms of hydrogen or electropositive element

Oxidation is defined as a process in which hydrogen or an electropositive element is removed from the substance i.e.

(a) Removal of hydrogen

$$\underset{\text{Ethanol}}{CH_3CH_2OH} \xrightarrow[300^\circ C]{Cu} \underset{\text{Ethanal}}{CH_3CHO} + H_2$$

(b) Removal of electropositive element

$$\underset{\text{Mercurous Chloride}}{Hg_2Cl_2} \rightarrow \underset{\substack{\text{(electropositive)}\\\text{element}}}{Hg} + \underset{\text{mercuric chloride}}{HgCl_2}$$

(iii) Oxidation in terms of electronic concept

Oxidation is a process in which loss of electrons take place. or

Oxidation is a process in which electrons are lost by an atom, ion or a group of atoms taking part in the chemical reaction.

As a result, there is increase in positive charge or decrease in negative charge of the atom or group of atoms in an oxidation reaction e.g.,

(a)

$$\underset{\text{Sodium atom}}{Na} \rightarrow \underset{\text{Sodium ion}}{Na^+} + \underset{\text{(Electron lost)}}{e^{-1}} \text{ (Increase in positive charge) or (loss of electron)}$$

(b) $$\underset{\text{Copper atom}}{Cu} \rightarrow \underset{\text{Copper ions}}{Cu^{2+}} \underset{\text{(electrons lost)}}{+2e^{-}} \text{ (Increase in positive charge) or (loss of electron)}$$

(c) $$\underset{\text{Ferrous}}{Fe^{2+}} \rightarrow \underset{\text{Ferric (electron lost)}}{Fe^{3+} + e^-} \text{ (Increase in positive charge) or (loss of electron)}$$

(d) $$\underset{\text{Chloride ion}}{Cl^-} \rightarrow \underset{\text{Chlorine atom}}{Cl} + \underset{\text{electron lost}}{e^-}$$ (Decrease in negative charge) or (loss of electron)

(e) $$\underset{\text{Sulphide ion}}{S^{2-}} \rightarrow \underset{\text{Sulphur atom}}{S} + \underset{\text{e's lost}}{2e^-} \text{ (Decrease in negative charge) or (loss of electron)}$$

4.5.2 Reduction

Reduction can also be defined in three ways :

(i) Reduction in terms of oxygen or electronegative element: Reduction is a process in which oxygen or electronegative element is removed from a substance i.e.

(a) $$\underset{\text{Auric chloride}}{AuCl_3} \rightarrow \underset{\text{Aurous chloride}}{AuCl} + \underset{\text{Chorine}}{Cl_2} \text{ (Removal of electro negative element)}$$

(b) $$\underset{\text{Copper oxide}}{CuO} + H_2 \xrightarrow{\text{heat}} \underset{\text{Copper}}{Cu} + H_2O \text{ (Removal of oxygen from copper oxide)}$$

(c) $$\underset{\substack{\text{Zinc}\\\text{Oxide}}}{ZnO} + C \rightarrow \underset{\text{Zinc}}{Zn} + CO \text{ (Removal of oxygen from zinc oxide)}$$

(ii) Reduction in terms of hydrogen or electropositive element: Reduction is a process in which hydrogen or electropositive element is added to a substance i.e.

(a) $$\underset{\substack{\text{mercuric}\\\text{chloride}}}{HgCl_2} + \underset{\substack{\text{mercury}\\\text{(electropositive)}}}{Hg} \rightarrow \underset{\substack{\text{mercurous}\\\text{chloride}}}{Hg_2Cl_2} \text{ (Addition of electropositive element)}$$

(b) $$\underset{\substack{\text{Cupric}\\\text{chloride}}}{CuCl_2} + \underset{\substack{\text{Copper}\\\text{(electropositive)}}}{Cu} \rightarrow \underset{\text{Cuprous Chloride}}{Cu_2Cl_2} \text{ (Addition of electropositive element)}$$

(c) $\underset{\text{Sodium}}{2Na} + \underset{\text{Hydrogen}}{H_2} \rightarrow \underset{\text{Sodium hydride}}{2NaH}$ (Addition of hydrogen to Sodium)

(d) $\underset{\text{Copper (II) oxide}}{CuO} + \underset{\text{Hydrogen}}{H_2} \rightarrow \underset{\text{Copper}}{Cu} + \underset{\text{Water}}{H_2O}$ (Addition of hydrogen to an oxide)

(iii) **Reduction in terms of electronic concept :** Reduction is a process in which electrons are gained by an atom, ion or a group of atoms. As a result, there is an increase in negative charge or decrease in positive charge e.g.,

(a) $\underset{\substack{\text{Chlorine}\\ \text{atom}}}{Cl} + \underset{\substack{\text{electron}\\ \text{(gained)}}}{e^-} \rightarrow \underset{\substack{\text{Chloride}\\ \text{ion}}}{Cl^-}$ (gain of electron or increase in negative charge)

(b) $\underset{\substack{\text{magnesium}\\ \text{ion}}}{Mg^{2+}} + \underset{\substack{\text{electrons}\\ \text{(gained)}}}{2e^-} \rightarrow \underset{\substack{\text{magnesium}\\ \text{atom}}}{Mg}$ (gain of electrons or decrease in positive charge)

(c) $\underset{\substack{\text{Ferric}\\ \text{ion}}}{Fe^{3+}} + \underset{\substack{\text{electrons}\\ \text{(gained)}}}{3e^-} \rightarrow \underset{\text{iron}}{Fe}$ (gain of electrons or decrease in positive charge)

4.6 REDOX REACTIONS

In the term 'redox', 'red' stands for reduction and 'ox' stands for oxidation.

Thus the reactions in which oxidation and reduction take place simultaneously are called Redox reactions, i.e. in redox reactions one substance is oxidized and other is reduced.

4.6.1 Oxidizing agent

It is a substance which itself gets reduced but oxidizes the other substance by

(i) Adding oxygen or electronegative element to other substance.

(ii) Removing hydrogen or electropositive element from other substance.

(iii) Gaining electrons from other substances.

4.6.2 Reducing agent

It is a substance which itself gets oxidized but reduces the other substance by :

(i) Adding hydrogen or electropositive element to substance.

(ii) Removing oxygen or electronegative element from a substance.

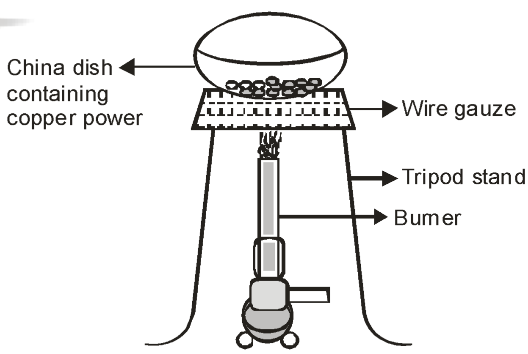
(iii) Losing or donating electrons to other substances.

Let us discuss oxidizing agents and reducing agents by taking some examples of oxidation-reduction or redox reactions.

Example 1: Oxidation of copper to copper oxide and reduction of copper oxide to copper.

Experiment: Take about 1 g of copper powder in a china dish and heat it as shown in figure.

It is observed that surface of copper powder becomes coated with black



copper (II) oxide and the reaction taking place is

Figure 12

$$\underset{\text{Copper Powder (brown)}}{2Cu(s)} + \underset{\text{Oxygen (from air)}}{O_2(g)} \xrightarrow{\text{heat}} \underset{\text{Copper (II) oxide (black)}}{2CuO}$$

It is clear from the reaction that there is addition of oxygen to copper, as a result copper has been oxidized to copper (II) oxide. Thus, it is an oxidation reaction.

Now, if hydrogen gas is passed over Cu(II) oxide, the black Copper (II) oxide changes to brown copper:-

$$\underset{\text{Copper (II) oxide (black)}}{CuO(s)} + \underset{\text{hydrogen}}{H_2(g)} \rightarrow \underset{\text{Copper (brown)}}{Cu(s)} + H_2O\,(g)$$

It is clear from the reaction that there is removal of oxygen from copper oxide and addition of oxygen to hydrogen, as a result, copper oxide has been reduced to copper and hydrogen has been oxidized to water. Since oxidation and reduction occur simultaneously, in this reaction, so it is an example of **redox reaction**. Which can be shown more clearly as follows :

Removal of oxygen: Reduction

$$\underset{\text{Oxidizing agent}}{CuO} + \underset{\text{Reducing agent}}{H_2} \xrightarrow{\text{heat}} Cu + H_2O$$

Addition of oxygen–oxidation

In the above reaction copper oxide (CuO) is giving oxygen required for the oxidation of hydrogen, therefore, CuO is oxidizing agent. Hydrogen is responsible for removing oxygen from CuO, therefore, hydrogen is the reducing agent.

(i) Substance oxidized : H_2
(ii) Substance reduced : CuO
(iii) Oxidizing agent : CuO
(iv) Reducing agent : H_2

Example 2*:* When hydrogen sulphide reacts with chlorine, then sulphur and hydrogen chloride are formed as:

$$\underset{\text{hydrogen sulphide}}{H_2S(g)} + \underset{\text{Chlorine}}{Cl_2(g)} \rightarrow \underset{\text{Sulphur}}{S(s)} + \underset{\text{Hydrogen chloride}}{2HCl(g)}$$

In this reaction there is removal of hydrogen from H_2S, which is being oxidized to S and Cl_2 is being reduced to HCl. Since both oxidation and reduction occur together, so it is an example of redox reaction. Which can be shown very clearly as follows:-

Removal of hydrogen : oxidation

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

Reducing agent (H₂S) oxidising agent (Cl₂)

Addition of hydrogen : Reduction

In the above reaction, H_2S is giving hydrogen required for reduction of Cl_2, therefore, H_2S is reducing agent. Cl_2 is responsible for removing hydrogen from H_2S, therefore Cl_2 is the oxidizing agent.

Thus,

(i) Substance oxidized : H_2S

(ii) Substance reduced : Cl_2

(iii) Oxidizing agent : Cl_2

(iv) Reducing agent : H_2S

Some more examples of oxidation reduction or Redox reactions:-

(i)

Addition of oxygen : oxidation

$$ZnO(s) + C(s) \xrightarrow{heat} Zn(s) + CO(g)$$

Oxidising agent (ZnO) Reducing agent (C)

Removal of oxygen : Reduction

(ii)

Addition of oxygen : oxidation

$$Fe_2O_3 + 2Al \xrightarrow{heat} 2Fe + Al_2O_3$$

Oxidising agent (Fe_2O_3) Reducing agent (Al)

Removal of oxygen : Reduction

(iii)

Gain of electrons : Reduction

$$2Mg + O_2 \xrightarrow{heat} 2MgO$$

Reducing agent (Mg) Oxidising agent (O_2)

Loss of electrons : oxidation

(iv)

$$Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$$

Loss of electrons : oxidation

$$Zn + Cu^{2+} \longrightarrow Zn^{2+} + Cu$$

Reducing agent (Zn) Oxiding agent (Cu^{2+})

Gain of electrons : Reduction

(v)

$$Cu + 2AgNO_3 \longrightarrow Cu(NO_3)_2 + 2Ag$$

Gain of electrons : Reduction

$$Cu + 2Ag^{+} \longrightarrow Cu^{2+} + 2Ag$$

Reducing agent (Cu) oxidising agent (Ag^+)

Loss of electrons : Oxidation

$$MnO_2 + 4HCl \longrightarrow MnCl_2 + Cl_2 + 2H_2O$$

Loss of electrons : oxidation

$$\underset{\text{oxidising agent}}{Mn^{4+}} + \underset{\text{Reducing agent}}{2Cl^-} \longrightarrow Mn^{2+} + Cl_2$$

(vi) Gain of electrons : Reduction

4.6.3 The effects of oxidation reactions in everyday life

The damaging effect of oxidation on metals is studied as **corrosion** and that on food is studied as **rancidity**. Thus, there are two common effects of oxidation reactions which we observe in everyday life:

(i) Corrosion of metals

(ii) Rancidity of food.

(i) **Corrosion :** The slow process of eating up of metals due to attack of atmospheric gases such as oxygen, carbon dioxide, hydrogen sulphide, water vapour etc. on the surface of the metals so as to convert the metal into oxide, carbonate, sulphide etc. is known as **corrosion**.

Example : Corrosion of iron (Rusting) :

The most common example of corrosion is rusting i.e., corrosion of iron. When an iron object is left in the moist air for a long time, its surface is covered with a brown, flaky (non-sticky) substance called rust. Rust is mainly hydrated ferric oxide ($Fe_2O_3.xH_2O$). It is formed due to attack of oxygen and water vapour present in moist air on the surface of iron :-

$$\underset{\text{iron}}{4Fe(s)} + \underbrace{3O_2(g) + 2xH_2O(l)}_{\text{from moist air}} \xrightarrow{\text{Oxidation}} \underset{\text{hydrated ferric oxide (Rust)}}{2Fe_2O_3.xH_2O(s)}$$

This reaction is called **corrosion of iron or rusting**.

Effect of Rusting in everyday life : Rusting is a serious problem because it weakens the structure of bridges, iron railings, automobile parts etc. Every year, a lot of money is spent to replace rusted iron and steel structures. The reason is that raddish brown crust of rust does not stick to the surface. It falls down exposing fresh surface for rusting. Thus, corrosion of iron or rusting is a continuous process which ultimately eats up the whole iron object.

Methods to prevent rusting or Prevention of Rusting : Rusting can be prevented if iron objects are not allowed to come in contact with the moist air. It can be done by :

(a) Painting the iron objects such as iron gates, steel furniture, bodies of cars etc.

(b) Greasing and oiling the iron objects such as machine parts etc.

(c) Galvanisation is coating the surface of iron objects with a thin layer of zinc, which is more resistant to corrosion.

SOME MORE EXAMPLES OF CORROSION OF METALS

1. **Corrosion of copper :** When a copper object (shiny brown) is left in moist air for a long time, then its surface is covered with green coating of basic carbonate, $CuCO_3 . Cu(OH)_2$. This is due to attack of O_2, CO_2 and water vapour present in moist air on the surface of copper :

$$\overbrace{2Cu(s)}^{\text{Oxidation}}_{\substack{\text{Copper}\\\text{(Brown)}}} + \underbrace{CO_2(g) + O_2(g) + H_2O(l)}_{\text{from most air}} \rightarrow \underset{\substack{\text{Basic Copper}\\\text{carbonate (Green)}}}{CuCO_3.Cu(OH)_2}$$

This reaction is called **corrosion of copper**.

2. **Corrosion of Silver :** When a silver object (shiny white) is kept in air for a long time, its surface is covered with coating of black silver sulphide (Ag_2S). This is due to the attack of H_2S gas present in air on the surface of silver.

$$\overset{\text{Oxidation}}{\underset{\substack{\text{Silver}\\\text{(White)}}}{2Ag(s)}} + \underset{\substack{\text{hydrogen}\\\text{sulphide}\\\text{(from air)}}}{H_2S(g)} \rightarrow \underset{\substack{\text{Silver}\\\text{sulphide}\\\text{(black)}}}{Ag_2S(s)} \rightarrow \underset{\text{Hydrogen}}{H_2(g)}$$

This reaction is called the corrosion of silver.

(ii) Rancidity : When fats and oils in food are oxidized, their smell and taste changes and they become rancid. This phenomenon is called **rancidity**.

Prevention from Rancidity :

(a) adding antioxidants (Reducing agents) like BHA (Butylated Hydroxy Anisole) and BHT (Butylated Hydroxy-Toluene) to foods containing fats and oils.

(b) Packaging fat and oil containing foods in nitrogen gas (inert gas)

(c) Keeping food in a refrigerator

(d) Storing food in air tight containers.

(e) Storing foods away from light.

5. TYPES OF CHEMICAL REACTIONS BASED ON HEAT CHANGES

Depending upon the kind of heat change during a reaction, the chemical reactions are classified into two types :

(i) Exothermic Reactions

(ii) Endothermic Reactions

(i) Exothermic Reactions : The term exothermic is taken from Greek word exotherm (exo-out, therm-heat) which means heat goes out. Thus a reaction in which heat is released or produced is called an exothermic reaction.

Examples :

(a) When nitrogen and hydrogen combine then, ammonia (NH_3) is formed and heat is liberated. Thus, formation of ammonia is an exothermic reaction.

$$\underset{\text{nitrogen}}{N_2(g)} + \underset{\text{hydrogen}}{3H_2(g)} \rightarrow \underset{\text{Ammonia}}{2NH_3(g)} + \underset{\text{(heat released)}}{92 kJ\,mol^{-1}}$$

(b) When methane gas is burnt, heat is liberated, along with CO_2 and H_2O. Thus, combustion of methane gas is an exothermic reaction.

$$\underset{\text{Methane gas}}{CH_4(g)} + \underset{\text{Oxygen (from air)}}{2O_2(g)} \rightarrow \underset{\text{Carbondioxide}}{CO_2(g)} + \underset{\text{water}}{2H_2O(l)} + \underset{\text{heat released}}{890kJ}$$

(c) Burning of Coke in air is an exothermic reaction.

$$\underset{\text{Coke}}{C(s)} + \underset{\text{Oxygen (from air)}}{O_2(g)} \rightarrow \underset{\text{Carbondioxide}}{CO_2(g)} + \underset{\text{(heat released)}}{393.5\,kJ\,mol^{-1}}$$

(d) Formation of water from $H_2(g)$ and $O_2(g)$ is an exothermic reaction.

$$\underset{\text{hydrogen}}{H_2(g)} + \underset{\text{Oxygen}}{1/2\,O_2(g)} \rightarrow \underset{\text{(steam)}}{H_2O(g)} + \underset{\text{(heat released)}}{241\,kJ\,mol^{-1}}$$

The steam $H_2O(g)$ so produced condenses into liquid water and 45 kJ heat is released.

$$\underset{\text{steam}}{H_2O(g)} \rightarrow \underset{\text{Water}}{H_2O(l)} + \underset{\text{heat released}}{45\,kJ\,mol^{-1}}$$

(e) During respiration process glucose undergoes combustion to form CO_2 and H_2O. The reaction is exothermic.

$$\underset{\text{Glucose}}{C_6H_{12}O_6} + \underset{\text{Oxygen}}{6O_2} \rightarrow \underset{\text{Carbondioxide}}{6CO_2} + \underset{\text{Water}}{6H_2O} + \underset{\text{Energy (heat)}}{2820\,kJ\,mol^{-1}}$$

WHY AND WHEN A CHEMICAL REACTION IS EXOTHERMIC

We know that in a chemical reaction, heat is supplied to break the bonds in the reactants and heat is released due to bond formation in the products. When the heat released due to bond formation in products is greater than the heat supplied to break the bonds in reactants, then the reaction is exothermic.

For example formation of H_2O from (H_2) and (O_2) is exothermic because the energy released in the formation of 2 covalent bonds in H_2O is more than the energy, required to break

H – H bonds in H_2 and O = O bonds in O_2.

5.1 ENDOTHERMIC REACTIONS

The term endothermic is taken from the Greek word endotherm (endo-in, therm-heat) which means heat is taken in. Thus a reaction in which heat is absorbed is called the endothermic reaction.

Examples:

(a) Reaction between coke and steam is endothermic i.e., heat is absorbed by reactants to give the products :

$$\underset{\text{Coke}}{C(s)} + \underset{\text{Steam}}{H_2O(g)} + \underset{\text{heat absorbed}}{128\,kJ} \rightarrow \underset{\text{Carbon monoxide}}{CO(g)} + \underset{\text{Hydrogen}}{H_2(g)}$$

(b) Thionyl chloride ($SOCl_2$) is a useful drying agent. Its reaction with water is endothermic.

$$\underset{\text{Thionyl Chloride}}{SOCl_2(l)} + \underset{\text{Water}}{H_2O(l)} + \underset{\text{heat absorbed}}{50\text{ kJ}} \rightarrow \underset{\text{Sulphurdioxide}}{SO_2(g)} + \underset{\text{Hydrogen chloride}}{2HCl(g)}$$

(c) When nitrogen and oxygen combine to form nitric oxide (NO), heat is absorbed. This reaction is endothermic :

$$\underset{\text{nitrogen}}{N_2(g)} + \underset{\text{Oxygen}}{O_2(g)} + \underset{\text{heat absorbed}}{180\text{ kJ}} \rightarrow \underset{\text{Nitric oxide}}{2NO}$$

(d) Melting of ice into liquid water is an endothermic reaction (heat absorbing reaction)

$$\underset{\text{ice}}{H_2O(s)} + \underset{\text{Heat absorbed}}{6KJ} \rightarrow \underset{\text{water}}{H_2O(l)}$$

WHY AND WHEN A CHEMICAL REACTION IS ENDOTHERMIC

In a chemical reaction when heat required to break the bonds in all reactants is greater than the heat released in the bond formation of the products, then the reaction is said to be endothermic.

For example formation of nitric oxide from N_2 and O_2 is endothermic because the heat required to break N – N bonds in N_2 and (O=O) bonds in O_2 is more than the heat released in the formation of 2 covalent bonds in NO.