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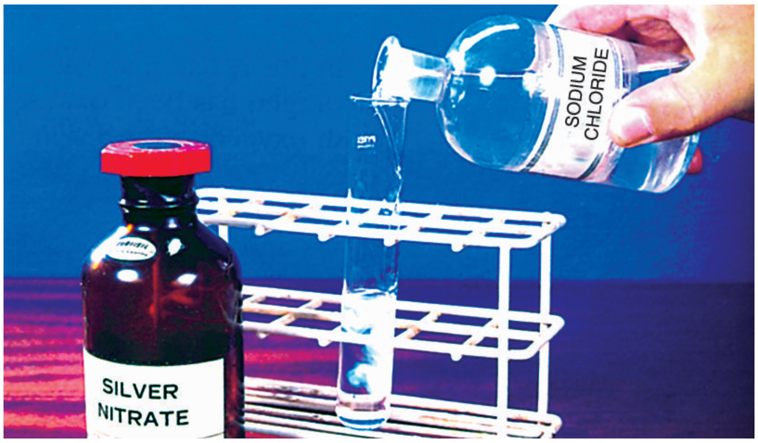


CHEMICAL REACTIONS AND EQUATIONS

Chemical reactions are the processes in which new substances with new properties are formed. Chemical reactions involve chemical changes. During chemical reactions, a rearrangement of atoms takes place between the reacting substances to form new substances having entirely different properties. Chemical reactions involve breaking of old chemical bonds which exist between the atoms of reacting substances, and then making of new chemical bonds between the rearranged atoms of new substances. **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.** We will now discuss reactants and products of a chemical reaction.

(*i*) **The substances which take part in a chemical reaction are called reactants.**

(*ii*) **The new substances produced as a result of chemical reaction are called products.**

In a chemical reaction, reactants are transformed into products. The products thus formed have properties which are entirely different from those of the reactants. We will now give an example of a chemical reaction. Before we do that please note that magnesium is a silvery-white metal. Magnesium metal is available in a science laboratory in the form of a magnesium ribbon (or magnesium wire). Let us study the chemical reaction of 'magnesium' with the 'oxygen' of air now.

When a magnesium ribbon is heated, it burns in air with a dazzling white flame to form a white powder called magnesium oxide. Actually, on heating, magnesium combines with oxygen present in air to form magnesium oxide :

$$\underset{\text{(As ribbon)}}{\text{Magnesium}} + \underset{\text{(From air)}}{\text{Oxygen}} \xrightarrow{\text{Heat}} \underset{\text{(White powder)}}{\text{Magnesium oxide}}$$

The burning of magnesium in air to form magnesium oxide is an example of a chemical reaction. In this chemical reaction there are two reactants 'magnesium and oxygen' but only one product 'magnesium oxide'. The properties of the product magnesium oxide are entirely different from those of the reactants magnesium and oxygen.

The magnesium ribbon which we use usually has a coating of 'magnesium oxide' on its surface which is formed by the slow action of oxygen of air on it. So, **before burning in air, the magnesium ribbon is cleaned by rubbing with a sand paper. This is done to remove the protective layer of magnesium oxide from the surface of magnesium ribbon** so that it may readily combine with the oxygen of air (on heating). Another point to be noted is that the dazzling (very bright) white light given out during the burning of magnesium ribbon is harmful to the eyes. So, the magnesium ribbon should be burned by keeping it as far as possible from the eyes. We can perform the chemical reaction involved in the burning of magnesium ribbon as follows :

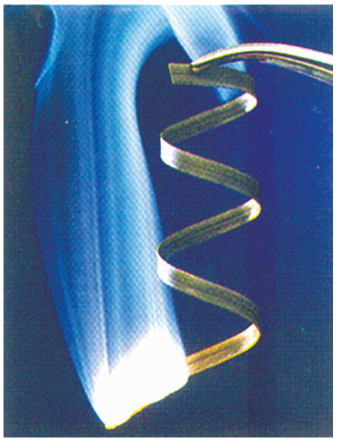
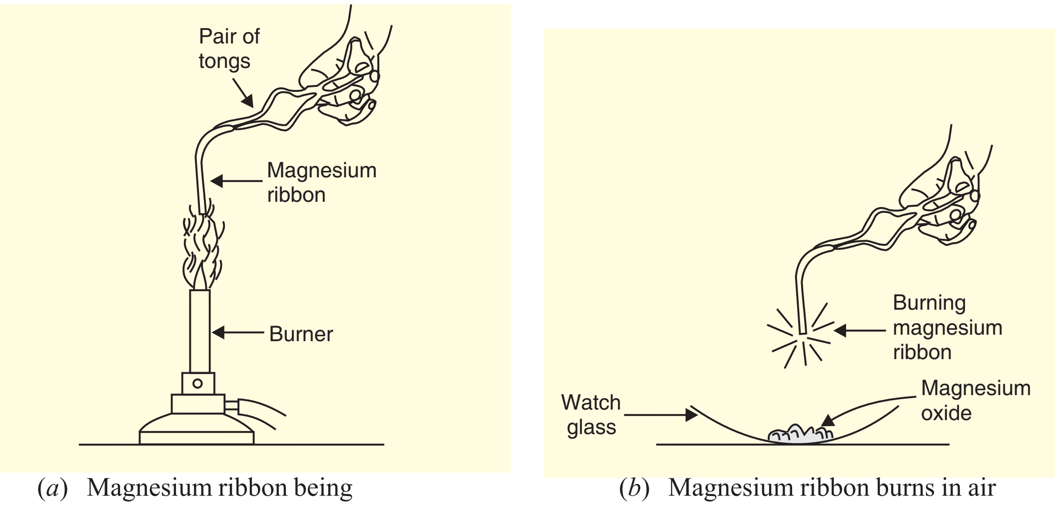


Figure 1. When magnesium burns in air, it combines with the oxygen of air to form magnesium oxide.

(*i*) Take about 2 cm long magnesium ribbon and clean it by rubbing its surface with sand paper.

(*ii*) Hold the magnesium ribbon with a pair of tongs at one end, and heat its other end over a burner [see Figure 2(*a*)]



(*a*) Magnesium ribbon being heated over a burner

(*b*) Magnesium ribbon burns in air to form magnesium oxide

Figure 2.

(*iii*) The magnesium ribbon starts burning with a dazzling white flame.

(*iv*) Hold the burning magnesium ribbon over a watch glass so that the magnesium oxide powder being formed collects in the watch glass [see Figure 2(*b*)].

It is not that chemical reactions can be carried out only in a science laboratory. A large number of chemical reactions keep on occurring in our daily life. **Souring of milk** (when left at room temperature during summer), **Formation of curd from milk, Cooking of food, Digestion of food in our body, Process of respiration, Fermentation of grapes, Rusting of iron** (when left exposed to humid atmosphere), **Burning of fuels** (like wood, coal, kerosene, petrol and LPG), **Burning of candle wax,** and **Ripening of fruits,** are all chemical changes which involve chemical reactions. In all these cases, the nature and identity of the initial substance changes because of the chemical reaction which takes place in it. An important question now arises : How do we come to know that a chemical reaction has taken place ? This will become clear from the following discussion on the characteristics of chemical reactions.

Figure 3. Ripening of fruits is a chemical change which involves chemical reactions.

Characteristics of Chemical Reactions

In a chemical reaction, the substances known as reactants are converted into new substances called products. *The conversion of reactants into products in a chemical reaction is often accompanied by some features which can be observed easily.* The easily observable features (or changes) which take place as a result of chemical reactions are known as characteristics of chemical reactions. **The important characteristics of chemical reactions are :**

(*i*) **Evolution of a gas,**

(*ii*) **Formation of a precipitate,**

(*iii*) **Change in colour,**

(*iv*) **Change in temperature,** and

(*v*) **Change in state**

Any one of these general characteristics can tell us whether a chemical reaction has taken place or not. For example, if on mixing two substances a gas is evolved, then we can say that a chemical reaction has taken place. We will now give examples to show all the characteristics of chemical reactions, one by one.

1. Evolution of a Gas

Some chemical reactions are characterised by the evolution of a gas. For example, when zinc granules react with dilute sulphuric acid, then bubbles of hydrogen gas are produced. So, **the chemical reaction between zinc and dilute sulphuric acid is characterised by the evolution of hydrogen gas.** (Please note that we can also use dilute hydrochloric acid in place of dilute sulphuric acid in this reaction.) We can perform this chemical reaction in the laboratory as follows :

(*i*) Take some zinc granules in a conical flask (or a test-tube).

(*ii*) Add dilute sulphuric acid over zinc granules.

(*iii*) We will see the bubbles of hydrogen gas being formed around zinc granules (see Figure 4).

(*iv*) If we touch the conical flask with our hand, we will find that it is somewhat hot. So, a change in temperature (rise in temperature) also occurs in this chemical reaction.

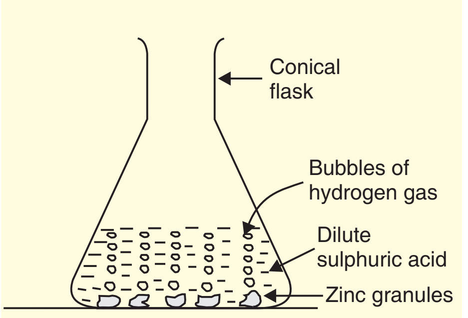


Figure 4. Dilute sulphuric acid reacts with zinc to evolve hydrogen gas.

When magnesium reacts with a dilute acid (like dilute hydrochloric acid or dilute sulphuric acid), even then hydrogen gas is evolved (see Figure 5). Let us take another example of a chemical reaction in which a gas is evolved. When dilute hydrochloric acid is poured over sodium carbonate in a test-tube, then carbon dioxide gas is evolved. So, **the chemical reaction between sodium carbonate and dilute hydrochloric acid is characterised by the evolution of carbon dioxide gas.**

Figure 5. When magnesium reacts with a dilute acid, then hydrogen gas is evolved.

Before we go further, we should know the meaning of the term 'precipitate'. **A precipitate is a 'solid product' which separates out from the solution during a chemical reaction.** A precipitate can be formed by mixing aqueous solutions (water solutions) of reactants when one of the products is insoluble in water. A precipitate can also be formed by passing a gas into an aqueous solution of a substance (like passing carbon dioxide gas into lime water).

2. Formation of a Precipitate

Some chemical reactions are characterised by the formation of a precipitate. For example, when potassium iodide solution is added to a solution of lead nitrate, then a yellow precipitate of lead iodide is formed (see Figure 6). Thus, **the chemical reaction between potassium iodide and lead nitrate is characterised by the formation of a yellow precipitate of lead iodide.** We can carry out this chemical reaction as follows :

(*i*) Take some lead nitrate solution in a test-tube (or a beaker).

(*ii*) Add potassium iodide solution to it.

(*iii*) A yellow precipitate of lead iodide is formed at once.

(*iv*) A change in colour (from colourless to yellow) also takes place in this chemical reaction.

Figure 6. Precipitation is the formation of an insoluble product. When colourless solutions of lead nitrate and potassium iodide are mixed, then a yellow precipitate of lead iodide is formed.

Let us take another example of a chemical reaction in which a precipitate is formed. When dilute sulphuric acid is added to barium chloride solution taken in a test-tube, then a white precipitate of barium sulphate is formed. Thus, **the chemical reaction between sulphuric acid and barium chloride solution is characterised by the formation of a white precipitate of barium sulphate.**

3. Change in Colour

Some chemical reactions are characterised by a change in colour. For example, when citric acid reacts with potassium permanganate solution, then the purple colour of potassium permanganate solution disappears (it becomes colourless). So, **the chemical reaction between citric acid and purple coloured potassium permanganate solution is characterised by a change in colour from purple to colourless.** We can perform the reaction between citric acid and potassium permanganate solution as follows. (Lemon juice contains citric acid, so we will use lemon juice as a source of citric acid for carrying out this reaction).

(*i*) Take some dilute potassium permanganate solution in a test-tube. It has purple colour.

(*ii*) Add lemon juice to it dropwise with the help of a dropper and shake the test-tube.

(*iii*) The purple colour of potassium permanganate solution goes on fading and ultimately it becomes colourless.

Let us take another example of a chemical reaction in which a change in colour takes place. When sulphur dioxide gas is passed through acidified potassium dichromate solution, then the orange colour of potassium dichromate solution changes to green (see Figure 7). Thus, **the chemical reaction between sulphur dioxide gas and acidified potassium dichromate solution is characterised by a change in colour from orange to green.**

(*a*) Potassium dichromate solution is orange in colour.

(*b*) When sulphur dioxide gas is passed through potassium dichromate solution, the orange colour of potassium dichromate solution changes to green.

Figure 7.

Before we go further, we should know why temperature changes take place in chemical reactions. Chemical reactions often produce heat energy. When a chemical reaction produces heat energy, then the temperature of reaction mixture rises (or increases) and it becomes hot. In some cases, however, chemical reactions absorb heat energy. When a chemical reaction absorbs heat energy, then the temperature of reaction mixture falls (or decreases) and it becomes cold. So, when we talk of 'change in temperature' in a chemical reaction, it can be 'rise in temperature' or 'fall in temperature'. Another point to be noted is that the compound 'calcium oxide' is known by two common names 'lime' as well as 'quicklime'. And the compound 'calcium hydroxide' is known as 'slaked lime'. Keeping these points in mind, we will now describe the change in temperature in chemical reactions.

4. Change in Temperature

Some chemical reactions are characterised by a change in temperature. For example, when quicklime reacts with water, then slaked lime is formed and a lot of heat energy is produced. This heat raises the temperature due to which the reaction mixture becomes hot. So, we can say that **the chemical reaction between quicklime and water to form slaked lime is characterised by a change in temperature (which is rise in temperature).** The reaction between quicklime and water to form slaked lime is an ***exothermic reaction*** (which means ***heat producing reaction***). We can perform this chemical reaction as follows :

(*i*) Take a little of quicklime in a hard-glass beaker [Figure 8(*a*)].

(*ii*) Add water to it slowly [Figure 8(*b*)].

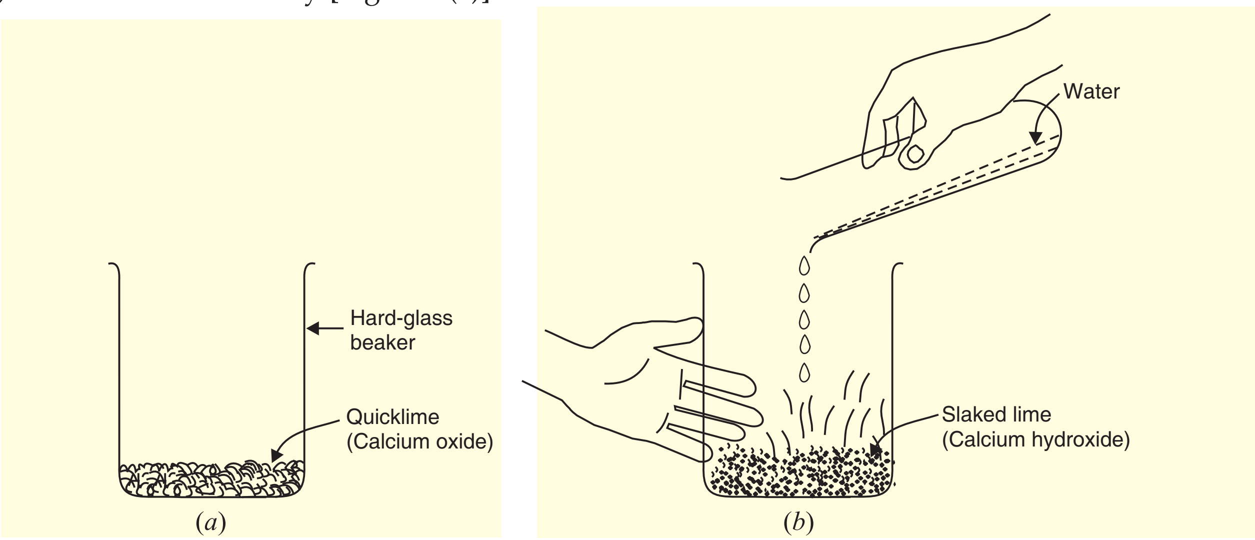


Figure 8. Quicklime reacts with water to form slaked lime releasing a lot of heat. The beaker becomes hot. Its temperature rises.

(*iii*) Touch the beaker carefully.

(*iv*) The beaker feels to be quite hot (Its temperature is high).

We have already studied the chemical reaction between zinc granules and dilute sulphuric acid to produce hydrogen gas. If we touch the conical flask containing zinc granules and dilute sulphuric acid, it is found to be warm (which means that the temperature rises during this reaction). Thus, **the chemical reaction between zinc granules and dilute sulphuric acid is also characterised by a change in temperature (which is rise in temperature).** The chemical reaction in which carbon burns in air to form carbon dioxide also releases a lot of heat (see Figure 9).

Figure 9. The chemical reaction in which carbon burns in air to form carbon dioxide releases a lot of heat.

We will now give one example of a chemical reaction in which heat energy is absorbed due to which the temperature falls. When barium hydroxide [$Ba(OH)_2$] is added to ammonium chloride (NH_4Cl) taken in a test-tube and mixed with a glass rod, then barium chloride, ammonia and water are formed. A lot of heat energy is absorbed during this reaction due to which the temperature of reaction mixture falls and the bottom of test-tube becomes very cold. Thus, **the chemical reaction between barium hydroxide and ammonium chloride to form barium chloride, ammonia and water is characterised by a change in temperature (which is fall in temperature)**. It is an *endothermic reaction* (which means ***heat absorbing reaction***).

5. Change in State

Some chemical reactions are characterised by a change in state. For example, when wax is burned (in the form of a wax candle), then water and carbon dioxide are formed (see Figure 10). Now, wax is a solid, water is a liquid whereas carbon dioxide is a gas. This means that during the combustion reaction of wax, the physical state changes from solid to liquid and gas. Thus, **the combustion reaction of candle wax is characterised by a change in state from solid to liquid and gas (because wax is a solid, water formed by the combustion of wax is a liquid at room temperature whereas carbon dioxide produced by the combustion of wax is a gas).**

Figure 10. The combustion reaction of candle wax is characterised by a change in state.

There are some chemical reactions which can show more than one characteristics. For example, the chemical reaction between zinc granules and dilute sulphuric acid shows two characteristics : evolution of a gas (hydrogen gas) and change in temperature (rise in temperature). Similarly, the chemical reaction between potassium iodide solution and lead nitrate solution shows two characteristics : formation of a precipitate (lead iodide precipitate) and change in colour (from colourless to yellow).

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 a chemical equation. Let us take one example to understand the meaning of a chemical equation clearly.

Zinc metal reacts with dilute sulphuric acid to form zinc sulphate and hydrogen gas. This reaction can be written in words as :

Zinc + Sulphuric acid ⟶ Zinc sulphate + Hydrogen

This is known as the word equation. We can change this word equation into a chemical equation by writing the symbols and formulae of the various substances in place of their names.

Now, Symbol of zinc is Zn

Formula of sulphuric acid is H_2SO_4

Formula of zinc sulphate is $ZnSO_4$

and, Formula of hydrogen is H_2

Figure 11. Zinc metal reacts with dilute sulphuric acid to form zinc sulphate solution and hydrogen gas.

So, putting the symbols and formulae of all the substances in the above word-equation, we get the following chemical equation :

$$\underbrace{Zn + H_2SO_4}_{\text{Reactants}} \longrightarrow \underbrace{ZnSO_4 + H_2}_{\text{Products}}$$

The substances which combine or react are known as reactants. Zinc and sulphuric acid are the reactants here. The reactants are always written on the left hand side in an equation with a plus sign (+) between them.

The new substances produced in a reaction are known as products. Zinc sulphate and hydrogen are the products in this case. The products are always written on the right hand side in an equation with a plus sign (+) between them.

The arrow sign (⟶) pointing towards the right hand side is put between the reactants and products. This arrow indicates that the substances written on the left hand side are combining to give the substances written on the right hand side in the equation. It should be clear by now that **a chemical equation is a short-hand method of representing a chemical reaction.**

Balanced and Unbalanced Chemical Equations

1. A balanced chemical equation has an equal number of atoms of different elements in the reactants and products. In other words, a balanced equation has an equal number of atoms of the different elements on both the sides. This point will become more clear from the following example.

Zinc reacts with dilute sulphuric acid to give zinc sulphate and hydrogen. This can be written in equation form as :

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

Let us count the number of atoms of all the elements in the reactants and products separately.

	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

We find that the reactants contain 1 zinc atom and products also contain 1 zinc atom. Reactants contain 2 hydrogen atoms and products also contain 2 hydrogen atoms. Similarly, reactants contain 1 sulphur atom and products also contain 1 sulphur atom. And finally, reactants contain 4 oxygen atoms and the products also contain 4 oxygen atoms. Thus, there is an equal number of atoms of different elements in the reactants and products, so the above chemical equation is a balanced equation. Since the number of atoms of various elements in reactants and products is equal, we can say that **a balanced chemical equation has equal masses of various elements in reactants and products.**

2. An unbalanced chemical equation has an unequal number of atoms of one or more elements in the reactants and products. In other words, an unbalanced equation has an unequal number of atoms of one or more elements on its two sides. This point will become more clear from the following example.

Hydrogen reacts with oxygen to form water. This reaction can be written in an equation form as :

$$H_2 + O_2 \longrightarrow H_2O$$

Let us count the number of hydrogen atoms and oxygen atoms in the reactants as well as product :

	In reactants	*In product*
No. of H atoms :	2	2
No. of O atoms :	2	1

In this equation, though the number of hydrogen atoms is equal in reactants and product (2 each), but the number of oxygen atoms is unequal. There are 2 oxygen atoms on the left side but only 1 oxygen atom on the right side. The above chemical equation contains an unequal number of oxygen atoms in reactants and product, so it is an unbalanced equation. Since the number of atoms of various elements in reactants and products is unequal, we can say that **an unbalanced equation has unequal masses of various elements in reactants and products.**

Figure 12. Launch of space shuttle "Discoverer". Liquid hydrogen burns in liquid oxygen to form water, providing a tremendous amount of energy to lift the shuttle out of the earth's gravitational field.

The equation : $H_2 + O_2 \rightarrow H_2O$, contains 2 oxygen atoms in the reactants but only 1 oxygen atom in the product. It appears as if 1 oxygen atom has been destroyed in this chemical reaction. This, however, cannot happen because according to the law of conservation of mass, **"matter can neither be created nor**

destroyed in a chemical reaction". This means that the total mass of all the reactants must be equal to the total mass of the products. In other words we can say that, **the number of various types of atoms in reactants must be equal to the number of same type of atoms in products.** It is obvious that we have to make the number of different types of atoms equal on both the sides of a chemical equation. To make the number of different types of atoms equal in reactants and products is known as balancing of an equation. It should be noted that **the chemical equations are balanced to satisfy the law of conservation of mass in chemical reactions.**

The reaction between hydrogen and oxygen to form water cannot be written as :

$$H_2 + O \longrightarrow H_2O$$

because oxygen occurs in the form of O_2 molecules and not as atoms O. All the substances have definite formulae which cannot be altered. So, **we should never change the formula of an element or a compound to balance an equation.** We can only multiply a symbol or a formula by figures like 2, 3, 4, etc. It will be good to note here that the elements which exist as diatomic molecules are oxygen, O_2, hydrogen, H_2, nitrogen, N_2, fluorine, F_2, chlorine, Cl_2, bromine, Br_2 and iodine, I_2. All other elements are usually considered monoatomic in equation writing and represented by their symbols. We will now learn the balancing of chemical equations.

Balancing of Chemical Equations

The process of making the number of different types of atoms equal on both the sides of an equation is called balancing of equation. The simple equations are balanced by hit and trial method. We will take one example to understand the balancing of equations by hit and trial method.

Hydrogen burns in oxygen to form water. This reaction can be written in an equation form as :

$$H_2 + O_2 \longrightarrow H_2O$$

In this reaction H_2 and O_2 are reactants whereas H_2O is the product. Let us count the number of hydrogen atoms and oxygen atoms in reactants and product.

	In reactants	*In product*
No. of H atoms :	2	2
No. of O atoms :	2	1

The number of hydrogen atoms is equal on both the sides (2 each), but the number of oxygen atoms is unequal. There are 2 oxygen atoms on the left side but only 1 oxygen atom on the right side. To have 2 oxygen atoms on the right side, we multiply H_2O by 2 and write $2H_2O$, so that :

$$H_2 + O_2 \longrightarrow 2H_2O$$

Let us count the number of various atoms on both the sides again :

	In reactants	*In product*
No. of H atoms :	2	4
No. of O atoms :	2	2

Though the number of oxygen atoms has become equal (2 on both sides), but the number of hydrogen atoms has now become unequal. There are 2 hydrogen atoms on the left side but 4 hydrogen atoms on the right side. To have 4 hydrogen atoms on the left side, we multiply H_2 by 2 and write $2H_2$, so that :

$$2H_2 + O_2 \longrightarrow 2H_2O$$

Let us count the number of various atoms on both the sides once again.

	In reactants	*In product*
No. of H atoms :	4	4
No. of O atoms :	2	2

This chemical equation contains an equal number of atoms of hydrogen and oxygen on both the sides, so this is a balanced equation.

After doing some more practice, we will find that there is no need to write so many steps to balance an equation. We will then be able to balance an equation in just one step.

To Make Equations More Informative

The equation which gives more information about the chemical reaction is known as more informative or information giving equation. **The chemical equations can be made more informative in three ways :**

1. **By indicating the "physical states" of the reactants and products.**
2. **By indicating the "heat changes" taking place in the reaction.**
3. **By indicating the "conditions" under which the reaction takes place.**

We will discuss these three points in detail one by one.

1. To Indicate the Physical States of Reactants and Products in an Equation. There can be four physical states for the reactants and products of a chemical reaction : solid, liquid, aqueous solution and gas.

Solid state is indicated by the symbol (s)

Liquid state is indicated by the symbol (l)

Aqueous solution (solution made in water) is indicated by the symbol (aq)

Gaseous state is indicated by the symbol (g)

The physical states of the reactants and products are shown by putting the above "state symbols" just after their symbols or formulae in an equation. This will become more clear from the following example.

Zinc metal reacts with dilute sulphuric acid to form zinc sulphate solution and hydrogen gas. This can be written as :

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

Here, Zinc metal is a solid, so we write Zn (s)

Dilute sulphuric acid is an aqueous solution, so we write H_2SO_4 (aq)

Zinc sulphate is also an aqueous solution, so we write $ZnSO_4$ (aq)

And, Hydrogen is a gas which is written as H_2 (g)

Figure 13. This picture shows zinc metal in the form of zinc granules. Zinc metal is a solid, so we write Zn (s) for it in a chemical equation.

The above equation can now be written as :

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

This equation is more informative because it tells us the physical states of the various substances involved in it. It tells us that zinc is in the solid state, sulphuric acid is in the form of an aqueous solution, zinc sulphate is also an aqueous solution but hydrogen is in gaseous state.

In some cases an insoluble product (called precipitate) is formed by the reaction between solutions of reactants (or a solution and a gas). Since the insoluble product (or precipitate) is a solid substance, its physical state is indicated in the equation by the symbol (s). For example, when calcium hydroxide solution (lime water) reacts with carbon dioxide gas, a white precipitate of calcium carbonate is formed alongwith water. This chemical reaction can be represented by the following chemical equation with state symbols of the reactants and products :

$$\underset{\substack{\text{Calcium hydroxide} \\ \text{(Lime water)}}}{Ca(OH)_2\,(aq)} + \underset{\text{Carbon dioxide}}{CO_2\,(g)} \longrightarrow \underset{\substack{\text{Calcium carbonate} \\ \text{(White ppt.)}}}{CaCO_3\,(s)} + \underset{\text{Water}}{H_2O\,(1)}$$

In this reaction, calcium carbonate is formed as a solid product (precipitate), so its physical state is indicated by the symbol (s). Please note that the word 'precipitate' is written in short form as 'ppt'. Since water is a liquid, so its physical state has been indicated by the symbol (l).

2. To Indicate the Heat Changes in an Equation. There are two types of reactions on the basis of heat changes involved : exothermic reactions and endothermic reactions.

(*i*) **Those reactions in which heat is evolved are known as exothermic reactions.** For example, when carbon burns in oxygen to form carbon dioxide, a lot of heat is produced in this reaction :

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

Figure 14. Carbon (in the form of coal) burns to produce heat. So, the burning of carbon is an exothermic reaction.

The burning of carbon in oxygen is an exothermic reaction because heat is evolved in this reaction. **An exothermic reaction is indicated by writing "+ Heat" or "+ Heat energy" or just "+ Energy" on the products' side of an equation** (as shown in the above equation). So, whenever we are told that a particular reaction is an exothermic reaction, we should at once write "+ Heat" or "+ Heat energy" or just "+ Energy" on the right side of the equation.

Natural gas is mainly methane (CH_4). When natural gas burns in the oxygen of air, it forms carbon dioxide and water vapour. A large amount of heat energy is also produced. This can be written as :

$$\underset{\substack{\text{Methane} \\ \text{(Natural gas)}}}{CH_4\ (g)} + \underset{\substack{\text{Oxygen} \\ \text{(From air)}}}{2O_2\ (g)} \longrightarrow \underset{\text{Carbon dioxide}}{CO_2\ (g)} + \underset{\text{Water}}{2H_2O\ (g)} + \text{Heat energy}$$

The burning of natural gas is an exothermic reaction because heat is produced in this reaction. Please note that **all the combustion reactions are exothermic reactions.** For example, combustion of fuels such as wood, coal, kerosene, petrol and diesel, are all exothermic reactions (because all these reactions produce heat energy). Even the combustion of food (like glucose) in our body during respiration is an exothermic reaction. This is discussed below.

We need energy to stay alive. We get this energy from the food we eat. During digestion, food is broken down into simpler substances. For example, the foods like *chapatti* (*roti*), bread, rice and potatoes, etc., contain mainly starch carbohydrate. During digestion, starch carbohydrate is broken down into a simple carbohydrate called glucose. **This glucose then undergoes slow combustion by combining with oxygen in the cells of our body to produce energy in a process called respiration. In addition to other functions, this energy maintains our body heat.**

During respiration, glucose combines with oxygen in the cells of our body to form carbon dioxide and water alongwith the production of energy :

$$\underset{\text{Glucose}}{C_6H_{12}O_6\ (aq)} + \underset{\text{Oxygen}}{6O_2\ (g)} \longrightarrow \underset{\text{Carbon dioxide}}{6CO_2\ (g)} + \underset{\text{Water}}{6H_2O\ (l)} + \text{Energy}$$

Respiration is an exothermic process because energy is produced during this process (as shown by the above equation).

The burning of a magnesium wire in air to form magnesium oxide is an exothermic reaction because heat and light energy are given out during this reaction. The decomposition of vegetable matter into compost is also an example of exothermic process (because heat energy is evolved during this process).

(*ii*) **Those reactions in which heat is absorbed are known as endothermic reactions.** For example, when nitrogen and oxygen are heated to a very high temperature (of about 3000°C) they combine to form nitrogen monoxide, and a lot of heat is absorbed in this reaction :

$$\underset{\text{Nitrogen}}{N_2\ (g)} + \underset{\text{Oxygen}}{O_2\ (g)} + \text{Heat} \longrightarrow \underset{\text{Nitrogen monoxide}}{2NO\ (g)}$$

The reaction between nitrogen and oxygen to form nitrogen monoxide is an endothermic reaction because heat is absorbed in this reaction. **An endothermic reaction is usually indicated by writing "+ Heat" or "+ Heat energy " or just "+ Energy" on the reactants' side of an equation** (as shown in the above equation). The reaction in which nitrogen and oxygen (of air) combine to form nitrogen monoxide takes place inside the engines of motor vehicles.

All the decomposition reactions require energy (in the form of heat, light or electricity) to take place. So, **all the decomposition reactions are endothermic reactions.** For example, when calcium carbonate is heated, it decomposes to form calcium oxide and carbon dioxide :

$$\underset{\text{Calcium carbonate}}{CaCO_3\ (s)} + \text{Heat} \longrightarrow \underset{\text{Calcium oxide}}{CaO\ (s)} + \underset{\text{Carbon dioxide}}{CO_2\ (g)}$$

The decomposition of calcium carbonate is an endothermic reaction because heat energy is absorbed in this reaction. **Photosynthesis is an endothermic reaction.** This is because sunlight energy is absorbed during the process of photosynthesis by green plants. **The electrolysis of water to form hydrogen and oxygen is also an endothermic reaction.** This is because electric energy is absorbed during this reaction. It is clear from this discussion that energy can be given out or absorbed in chemical reactions in the form of heat, light or electricity.

Figure 15. These white cliffs are made of calcium carbonate.

3. To Indicate the Conditions Under Which the Reaction Takes Place. If heat is required for a reaction to take place, then the heat sign delta (Δ) is put over the arrow of the equation. If the reaction takes place in the presence of a catalyst, then the symbol or formula of the catalyst is also written above or below the arrow sign in the equation. This will become more clear from the following example.

When potassium chlorate ($KClO_3$) is heated in the presence of manganese dioxide catalyst, it decomposes to form potassium chloride and oxygen gas. This can be written as :

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

Figure 16. This is manganese dioxide (MnO_2). It is used as a catalyst during the preparation of oxygen gas from potassium chlorate.

Here delta (Δ) stands for heat and MnO_2 is the catalyst. So, the above equation shows the conditions under which the reaction takes place. The conditions of temperature and pressure at which the reaction takes place can also be indicated in an equation by writing their values above or below the arrow sign in the equation. This will become clear from the following examples.

Methanol (or Methyl alcohol) is manufactured from carbon monoxide and hydrogen. The mixture of carbon monoxide and hydrogen gases is compressed to 300 atmospheres pressure and then passed over a catalyst consisting of a mixture of zinc oxide and chromium oxide heated to a temperature of 300°C. So, the conditions for this reaction to take place are : a pressure of 300 atmospheres (written as 300 atm), a temperature of 300°C, and a catalyst which is a mixture of zinc oxide and chromium oxide ($ZnO + CrO_3$). We can now write down a chemical equation for the reaction involved in the production of methanol alongwith conditions as follows :

$$\underset{\text{Carbon monoxide}}{CO\ (g)} + \underset{\text{Hydrogen}}{2H_2\ (g)} \xrightarrow[ZnO\ +\ CrO_3]{300\text{ atm; }300°C} \underset{\substack{\text{Methanol}\\\text{(Methyl alcohol)}}}{CH_3OH\ (l)}$$

The green plants make food by photosynthesis. During photosynthesis, carbon dioxide combines with water in the presence of 'sunlight' and the green pigment of leaves called 'chlorophyll' to make food like glucose and oxygen gas is given out. The conditions for the reaction of photosynthesis to take place are the presence of sunlight and chlorophyll. So, we can write a chemical equation for photosynthesis alongwith conditions as follows :

$$\underset{\text{Carbon dioxide}}{6CO_2\,(g)} + \underset{\text{Water}}{6H_2O\,(l)} \xrightarrow[\text{Chlorophyll}]{\text{Sunlight}} \underset{\text{Glucose}}{C_6H_{12}O_6\,(aq)} + \underset{\text{Oxygen}}{6O_2\,(g)}$$

Important Examples on Writing of Balanced Chemical Equations

We should remember the following four steps for writing equations for the chemical reactions :

First step : Write down the chemical reaction in the form of a word equation, keeping the reactants on the left side and products on the right side.

Second step : Put the symbols and formulae of all the reactants and products in the word equation.

Third step : Balance the equation by multiplying the symbols and formulae by the smallest possible figures (Do not change the formulae to balance the equation).

Fourth step : If possible, make the equation more informative by indicating the physical states of reactants and products ; by indicating the heat changes, if any, taking place in the reaction ; and by indicating the conditions under which the reaction takes place. If, however, you do not have sufficient information regarding the physical states ; heat changes and conditions of the reaction, this step may be avoided.

Keeping these points in mind, let us solve some problems now.

Sample Problem 1. Write a balanced equation for the following reaction :

Methane burns in oxygen to form carbon dioxide and water.

Solution. This reaction can be written in the form of a word equation as :

$$\text{Methane} + \text{Oxygen} \longrightarrow \text{Carbon dioxide} + \text{Water}$$

Now, Formula of methane is CH_4

Formula of oxygen is O_2

Formula of carbon dioxide is CO_2

And, Formula of water is H_2O

Writing the formulae of all the substances in the above word equation, we get :

$$CH_4 + O_2 \longrightarrow CO_2 + H_2O$$

Let us count the number of various atoms in reactants and products :

	In reactants	*In products*
No. of C atoms :	1	1
No. of H atoms :	4	2
No. of O atoms :	2	3

The number of carbon atoms is equal on both the sides (1 each) but the number of hydrogen atoms and oxygen atoms is not equal. There are 4 hydrogen atoms on the left side but only 2 hydrogen atoms on the right side. To have 4 hydrogen atoms on the right side, we multiply H_2O by 2 and write $2H_2O$. Thus,

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

Counting the number of various atoms on both the sides again, we get :

	In reactants	*In products*
No. of C atoms :	1	1
No. of H atoms :	4	4
No. of O atoms :	2	4

Only the number of oxygen atoms is unequal now. There are 2 oxygen atoms on the left side but 4 on the right side. To have 4 oxygen atoms on the left side, we multiply O_2 by 2 and write $2O_2$:

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

Let us count the number of various atoms on the two sides once again :

	In reactants	*In products*
No. of C atoms :	1	1
No. of H atoms :	4	4
No. of O atoms :	4	4

This chemical equation contains an equal number of various types of atoms in the reactants and products, so this is a balanced equation.

Discussion. The above equation can be made more informative by indicating the physical states of the reactants and products as well as the heat changes taking place in the reaction as discussed below :

Methane is a gas, so we write CH_4 (g)

Oxygen is a gas, so we write O_2 (g)

Carbon dioxide is a gas, so we write CO_2 (g)

What about the physical state of H_2O ?

Figure 17. When methane gas burns in oxygen gas (of air), it forms carbon dioxide gas and water vapour. A lot of heat is also produced.

If a reaction takes place in the aqueous medium, then H_2O is in the liquid state and we write, H_2O (l) for it. If the reaction takes place in the vapour phase, then H_2O is in the gaseous state and represented as H_2O (g). In this case, methane gas burns in oxygen gas to form carbon dioxide gas and water vapour or steam. So, water is in the gaseous state here and we write H_2O (g). If we put the physical states of all the reactants and products, then the above equation can be written as :

$$CH_4\,(g) + 2O_2\,(g) \longrightarrow CO_2\,(g) + 2H_2O\,(g)$$

We will now discuss the heat changes taking place in this reaction. When methane burns in oxygen to form carbon dioxide and water, a lot of heat is also produced, so this is an exothermic reaction. An exothermic reaction is indicated by writing "+ Heat" sign on the products' side. So, the above equation can finally be written as :

$$CH_4\,(g) + 2O_2\,(g) \longrightarrow CO_2\,(g) + 2H_2O\,(g) + \text{Heat}$$

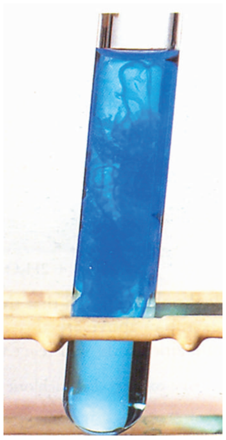


Figure 18. When sodium hydroxide solution is added to copper sulphate solution, a blue precipitate of copper hydroxide is formed alongwith sodium sulphate solution.

This equation now gives the physical states of the reactants and products as well as the heat changes taking place in the reaction, so this is a more informative equation.

Sample Problem 2. Convey the following information in the form of a balanced chemical equation :

On adding an aqueous solution of sodium hydroxide to an aqueous solution of copper sulphate, copper hydroxide is precipitated and sodium sulphate remains in solution.

Solution. In this reaction copper sulphate reacts with sodium hydroxide to form copper hydroxide and sodium sulphate. This can be written in the form of a word-equation as :

$$\text{Copper sulphate} + \text{Sodium hydroxide} \longrightarrow \text{Copper hydroxide} + \text{Sodium sulphate}$$

Now, Formula of copper sulphate is $CuSO_4$

Formula of sodium hydroxide is $NaOH$

Formula of copper hydroxide is $Cu(OH)_2$

And, Formula of sodium sulphate is Na_2SO_4

Putting these formulae in the above word-equation, we get :

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

Let us count the number of various types of atoms in reactants as well as products.

	In reactants	*In products*
No. of Cu atoms :	1	1
No. of S atoms :	1	1
No. of O atoms :	5	6
No. of Na atoms :	1	2
No. of H atoms :	1	2

We find that the number of copper atoms and sulphur atoms is equal on both the sides (1 each), but the number of oxygen atoms, sodium atoms and hydrogen atoms is not equal. Let us take the oxygen atoms first. There are 5 oxygen atoms on left side but 6 oxygen atoms on the right side. To have 6 oxygen atoms on the left side, we multiply NaOH by 2 and write 2NaOH. Thus,

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

Let us count the number of various types of atoms on both the sides once again.

	In reactants	*In products*
No. of Cu atoms :	1	1
No. of S atoms :	1	1
No. of O atoms :	6	6
No. of Na atoms :	2	2
No. of H atoms :	2	2

This equation contains an equal number of various types of atoms on both the sides, so this is a balanced equation.

We will now indicate the physical states of the reactants and products which have been given to us in this problem.

Copper sulphate is an aqueous solution, so we write $CuSO_4$ (aq)

Sodium hydroxide is also an aqueous solution, so we write NaOH (aq)

Copper hydroxide is formed as a precipitate (solid), so we write $Cu(OH)_2$ (s)

Sodium sulphate is in solution, so we write Na_2SO_4 (aq)

The above equation can now be written as :

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

Discussion. Before we answer the next question on the balancing of equations, we should know something about the various oxides of iron metal. Iron (Fe) forms two main oxides :

(*i*) **Iron (II) oxide, FeO.** This is called iron (II) oxide because the valency of iron in it is II (two). The common name of iron (two) oxide, FeO, is **ferrous oxide.**

(*ii*) **Iron (III) oxide, Fe_2O_3.** This is called iron (III) oxide because the valency of iron in it is III (three). The common name of iron (three) oxide, Fe_2O_3, is **ferric oxide.**

Figure 19. This is iron (II, III) oxide. When heated iron metal reacts with steam, it forms iron (II, III) oxide.

A third oxide of iron is Fe_3O_4. Actually, Fe_3O_4 is a mixture of iron (II) oxide FeO and iron (III) oxide, Fe_2O_3. So, **Fe_3O_4 is named as iron (II, III) oxide** ($Fe_3O_4 = FeO + Fe_2O_3$). The common name of Fe_3O_4 is **magnetic iron oxide.**

Another point to remember is that steam is the gaseous form of water, so the formula of steam is the same as that of water, which is H_2O. It will now be easy for us to understand the next question on balancing of equations.

Sample Problem 3. Write a balanced chemical equation with state symbols for the following reaction :

Heated iron metal reacts with steam to form iron (II, III) oxide, (Fe_3O_4) and hydrogen.

(NCERT Book Question)

Solution. This reaction can be written in the form of a word equation as :

Iron + Steam ⟶ Iron (II, III) oxide + Hydrogen

Now, Symbol of iron is Fe

Formula of steam is H_2O (It is the same as water)

Formula of iron (II, III) oxide is Fe_3O_4 (Given)

And, Formula of hydrogen is H_2

By writing the symbols and formulae of all the substances in the above word equation, we get the following skeletal chemical equation :

$$Fe + H_2O \longrightarrow Fe_3O_4 + H_2$$

Here Fe and H_2O are reactants whereas Fe_3O_4 and H_2 are the products. Let us count the number of atoms of various types in the reactants and products :

	In reactants	*In products*
No. of Fe atoms :	1	3
No. of H atoms :	2	2
No. of O atoms :	1	4

We can see that the number of iron atoms on the left side of the equation is only 1 but there are 3 iron atoms on the right side. Now, to have 3 iron atoms on the left side, we multiply Fe by 3 and write it as 3Fe. Thus,

$$3Fe + H_2O \longrightarrow Fe_3O_4 + H_2$$

Let us count the number of various types of atoms on both the sides again :

	In reactants	*In products*
No. of Fe atoms :	3	3
No. of H atoms :	2	2
No. of O atoms :	1	4

Now the number of iron atoms is equal on both sides (3 each) and the number of hydrogen atoms is also equal (2 each). But the number of oxygen atoms is not equal. There is only 1 oxygen atom on the left side but 4 oxygen atoms on the right side. So, to have 4 oxygen atoms on the left side, we multiply H_2O by 4 and write it as $4H_2O$. This will give us :

$$3Fe + 4H_2O \longrightarrow Fe_3O_4 + H_2$$

Let us count the number of various atoms on the two sides once again :

	In reactants	*In products*
No. of Fe atoms :	3	3
No. of H atoms :	8	2
No. of O atoms :	4	4

The number of hydrogen atoms now becomes unequal. There are 8 hydrogen atoms on the left side but only 2 hydrogen atoms on the right side. Now, to get 8 hydrogen atoms on the right side, we multiply H_2 by 4 and write it as $4H_2$. This gives us the following equation :

$$3Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2$$

Let us count the number of various types of atoms on both the sides of this equation :

	In reactants	*In products*
No. of Fe atoms :	3	3
No. of H atoms :	8	8
No. of O atoms :	4	4

The above chemical equation contains an equal number of Fe, H and O atoms in the reactants and products, so this is a balanced equation.

Iron (Fe) is a solid, steam (H_2O) is a gas, iron (II, III) oxide (Fe_3O_4) is a solid and hydrogen (H_2) is a gas. So, we can write the above chemical equation with state symbols as follows :

$$3Fe\,(s) \quad + \quad 4H_2O\,(g) \longrightarrow Fe_3O_4\,(s) \quad + \quad 4H_2\,(g)$$

Sample Problem 4. Write the balanced equation for the following chemical reaction :

Hydrogen + Chlorine ⟶ Hydrogen chloride **(NCERT Book Question)**

Solution. In this problem, hydrogen combines with chlorine to form hydrogen chloride. This has been given to us in the form of a word equation as :

Hydrogen + Chlorine ⟶ Hydrogen chloride

Now, Formula of hydrogen is H_2

Formula of chlorine is Cl_2

And, Formula of hydrogen chloride is HCl

By putting these formulae in the above word-equation, we get the following chemical equation :

$$H_2 \quad + \quad Cl_2 \longrightarrow HCl$$

Let us balance this equation now. If we look at this equation carefully, we find that there are two hydrogen atoms and two chlorine atoms on the left side but only one hydrogen atom and one chlorine atom on the right side. Now, to have two hydrogen atoms and two chlorine atoms on the right side, we have to multiply HCl by 2 and write it as 2HCl. This gives us :

$$H_2 \quad + \quad Cl_2 \longrightarrow 2HCl$$

This is a balanced equation because it contains an equal number of hydrogen atoms and chlorine atoms in the reactants and products.

Hydrogen, chlorine and hydrogen chloride, are all gases, so we can write the above equation with state symbols as follows :

$$H_2\,(g) \quad + \quad Cl_2\,(g) \longrightarrow 2HCl\,(g)$$

Sample Problem 5. Translate the following statement into chemical equation and then balance the equation :

Hydrogen gas combines with nitrogen to form ammonia.

(NCERT Book Question)

Solution. In this reaction, hydrogen combines with nitrogen to form ammonia. This can be written as :

Hydrogen + Nitrogen ⟶ Ammonia

or $H_2 \quad + \quad N_2 \longrightarrow NH_3$

This equation has two H atoms on the left side but three H atoms on the right side. So, let us multiply H_2 by 3 and NH_3 by 2 so that each side gets 6H atoms :

$$3H_2 \quad + \quad N_2 \longrightarrow 2NH_3$$

Now, this equation contains an equal number of hydrogen atoms and nitrogen atoms on both the sides, so this is a balanced chemical equation.

Hydrogen, nitrogen and ammonia, are all gases, so we can write the above equation with state symbols as follows :

$$3H_2 (g) + N_2 (g) \longrightarrow 2NH_3 (g)$$

Figure 20. This is an ammonia manufacturing plant in which hydrogen and nitrogen are made to combine to form ammonia.

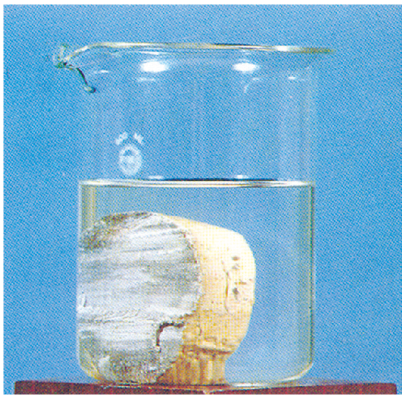


Figure 21. This beaker has a piece of sodium metal (which is stored under kerosene oil).

Figure 22. Sodium reacting with water to form sodium hydroxide and hydrogen. The heat produced during this reaction makes the hydrogen gas burn.

Sample Problem 6. Write the balanced chemical equation for the following reaction :

Sodium metal reacts with water to give sodium hydroxide and hydrogen.

(NCERT Book Question)

Solution. Here, sodium reacts with water to form sodium hydroxide and hydrogen. This can be written as :

$$\text{Sodium} + \text{Water} \longrightarrow \text{Sodium hydroxide} + \text{Hydrogen}$$

or $$Na + H_2O \longrightarrow NaOH + H_2$$

This equation has two H atoms on the left side but three H atoms on the right side. So, let us multiply H_2O by 2 and NaOH also by 2 so as to have an equal number of H atoms (4 each) on both the sides :

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

Now we have only one Na atom on left side but two Na atoms on the right side. So, let us take 2Na atoms on the left side. This gives us :

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

This equation contains an equal number of sodium, hydrogen and oxygen atoms on both the sides, so this is a balanced chemical equation.

Sodium is a solid (s), water is a liquid (l), sodium hydroxide is an aqueous solution (aq) whereas hydrogen is a gas (g). So, we can write the above chemical equation with state symbols as follows :

$$2Na (s) + 2H_2O (l) \longrightarrow 2NaOH (aq) + H_2 (g)$$

Sample Problem 7. Write a balanced chemical equation for the following chemical reaction :

Magnesium burns in oxygen to form magnesium oxide. **(NCERT Book Question)**

Solution. Magnesium burns in oxygen (of air) to form magnesium oxide. This reaction can be written in the form of a word equation as :

(*a*) This is magnesium metal (*b*) This is magnesium oxide

Figure 23. When magnesium burns in oxygen (of air), then magnesium oxide is formed.

$$\text{Magnesium} + \text{Oxygen} \longrightarrow \text{Magnesium oxide}$$

Now, Symbol of magnesium is Mg
Formula of oxygen is O_2
And, Formula of magnesium oxide is MgO

By putting these symbols and formulae in the above word equation, we get the following chemical equation :

$$Mg + O_2 \longrightarrow MgO$$

Let us balance this equation now. We can see from the above equation that there are 2 oxygen atoms on left side but only 1 oxygen atom on the right side. So, to have 2 oxygen atoms on the right side, we write 2MgO. Thus,

$$Mg + O_2 \longrightarrow 2MgO$$

Now we have 1 magnesium atom on left side but 2 magnesium atoms on the right side. To have 2 magnesium atoms on the left side, we write 2Mg. This gives us :

$$2Mg + O_2 \longrightarrow 2MgO$$

This equation contains an equal number of Mg atoms and O atoms on both the sides, so this is a balanced equation.

Magnesium is a solid, oxygen is a gas and magnesium oxide is also a solid. So, we can write the above chemical equation with state symbols as follows :

$$2Mg\ (s) + O_2\ (g) \longrightarrow 2MgO\ (s)$$

Before we go further and study the types of chemical reactions, **please answer the following questions :**

Very Short Answer Type Questions

1. Why is respiration considered an exothermic process ?

2. On what basis is a chemical equation balanced ?

3. What happens chemically when quicklime is added to water filled in a bucket ?

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

5. State whether the following statement is true or false :
A chemical equation can be balanced easily by altering the formula of a reactant or product.

6. What is wrong with the following chemical equation ?

$$Mg + O \longrightarrow MgO$$

Correct and balance it.

7. What does the symbol (aq) represent in a chemical equation ?

8. Why is photosynthesis considered an endothermic reaction ?

9. How will you indicate the following effects in a chemical equation ?
(*a*) A solution made in water
(*b*) Exothermic reaction
(*c*) Endothermic reaction

10. Translate the following statements into chemical equations and then balance the equations :
(*a*) Hydrogen sulphide gas burns in air to give water and sulphur dioxide.
(*b*) Phosphorus burns in oxygen to give phosphorus pentoxide.
(*c*) Carbon disulphide burns in air to give carbon dioxide and sulphur dioxide.
(*d*) Aluminium metal replaces iron from ferric oxide, Fe_2O_3, giving aluminium oxide and iron.
(*e*) Barium chloride reacts with zinc sulphate to give zinc chloride and barium sulphate.

11. Write the balanced chemical equations for the following reactions :
(*a*) Calcium hydroxide + Carbon dioxide $\longrightarrow$ Calcium carbonate + Water
(*b*) Aluminium + Copper chloride $\longrightarrow$ Aluminium chloride + Copper

12. Complete and balance the following equations :

(*a*) $NaOH + ... \longrightarrow Na_2SO_4 + H_2O$

(*b*) $Ca(OH)_2 + ... \longrightarrow CaCO_3 + H_2O$

13. Correct and balance the following equations :

(*i*) $Ca + H_2O \longrightarrow CaOH + H$

(*ii*) $N + H \longrightarrow NH_3$

14. Write complete balanced equations for the following reactions :

(*a*) Calcium (solid) + Water (liquid) $\longrightarrow$ Calcium hydroxide (solution) + Hydrogen (gas)

(*b*) Sulphur dioxide (gas) + Oxygen (gas) $\longrightarrow$ Sulphur trioxide (gas)

15. Balance the following equations :

(*i*) $Na + O_2 \longrightarrow Na_2O$

(*ii*) $H_2O_2 \longrightarrow H_2O + O_2$

(*iii*) $Mg(OH)_2 + HCl \longrightarrow MgCl_2 + H_2O$

(*iv*) $Fe + O_2 \longrightarrow Fe_2O_3$

(*v*) $Al(OH)_3 \longrightarrow Al_2O_3 + H_2O$

(*vi*) $NH_3 + CuO \longrightarrow Cu + N_2 + H_2O$

(*vii*) $Al_2(SO_4)_3 + NaOH \longrightarrow Al(OH)_3 + Na_2SO_4$

(*viii*) $HNO_3 + Ca(OH)_2 \longrightarrow Ca(NO_3)_2 + H_2O$

(*ix*) $NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + H_2O$

(*x*) $BaCl_2 + H_2SO_4 \longrightarrow BaSO_4 + HCl$

16. Fill in the following blanks with suitable words :

(*a*) Chemical equations are balanced to satisfy the law of ....

(*b*) A solution made in water is known as an .... solution and indicated by the symbol ....

Short Answer Type Questions

17. (*a*) Give one example of a chemical reaction.

(*b*) State two characteristics of the chemical reaction which takes place when dilute sulphuric acid is poured over zinc granules.

(*c*) Give two characteristics of the chemical reaction which occurs on adding potassium iodide solution to lead nitrate solution.

18. (*a*) What is a chemical equation ? Explain with the help of an example.

(*b*) Giving examples, state the difference between balanced and unbalanced chemical equations.

(*c*) Balance the following chemical equations :

(*i*) $NH_3 \longrightarrow N_2 + H_2$

(*ii*) $C + CO_2 \longrightarrow CO$

19. When hydrogen is passed over copper oxide, copper and steam are formed. Write a balanced equation for this reaction and state which of the chemicals are :

(*i*) elements (*ii*) compounds (*iii*) reactants

(*iv*) products (*v*) metals (*vi*) non-metals

20. (*a*) What are the various ways in which a chemical equation can be made more informative ? Give examples to illustrate your answer.

(*b*) Write balanced chemical equation from the following information :

An aqueous calcium hydroxide solution (lime water) reacts with carbon dioxide gas to produce a solid calcium carbonate precipitate and water.

21. (*a*) What is a balanced chemical equation ? Why should chemical equations be balanced ?

(*b*) Aluminium burns in chlorine to form aluminium chloride ($AlCl_3$). Write a balanced chemical equation for this reaction.

(*c*) Potassium metal reacts with water to give potassium hydroxide and hydrogen gas. Write a balanced chemical equation for this reaction.

22. (*a*) Explain, with example, how the physical states of the reactants and products can be shown in a chemical equation.

(*b*) Balance the following equation and add state symbols :

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

(*c*) Convey the following information in the form of a balanced chemical equation :

"An aqueous solution of ferrous sulphate reacts with an aqueous solution of sodium hydroxide to form a precipitate of ferrous hydroxide and sodium sulphate remains in solution."

23. Write any two observations in an activity which may suggest that a chemical reaction has taken place. Give an example in support of your answer.

24. (*a*) Aluminium hydroxide reacts with sulphuric acid to form aluminium sulphate and water. Write a balanced equation for this reaction.

(*b*) Balance the following chemical equation :

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

25. Write the balanced equations for the following reactions, and add the state symbols :

(*a*) Magnesium carbonate reacts with hydrochloric acid to produce magnesium chloride, carbon dioxide and water.

(*b*) Sodium hydroxide reacts with sulphuric acid to produce sodium sulphate and water.

26. Carbon monoxide reacts with hydrogen under certain conditions to form methanol (CH_3OH). Write a balanced chemical equation for this reaction indicating the physical states of reactants and product as well as the conditions under which this reaction takes place.

27. (*a*) Potassium chlorate ($KClO_3$) on heating forms potassium chloride and oxygen. Write a balanced equation for this reaction and indicate the evolution of gas.

(*b*) Rewrite the following information in the form of a balanced chemical equation :

Magnesium burns in carbon dioxide to form magnesium oxide and carbon.

28. (*a*) Substitute formulae for names and balance the following equation :

Calcium carbonate reacts with hydrochloric acid to produce calcium chloride, water and carbon dioxide gas.

(*b*) Write balanced chemical equation with state symbols for the following reaction :

Sodium hydroxide solution reacts with hydrochloric acid solution to produce sodium chloride solution and water.

29. Ammonia reacts with oxygen to form nitrogen and water. Write a balanced chemical equation for this reaction. Add the state symbols for all the reactants and products.

30. Write a balanced chemical equation for the process of photosynthesis giving the physical states of all the substances involved and the conditions of the reaction.

31. Translate the following statement into chemical equation and then balance it :

Barium chloride solution reacts with aluminium sulphate solution to form a precipitate of barium sulphate and aluminium chloride solution.

32. When potassium nitrate is heated, it decomposes into potassium nitrite and oxygen. Write a balanced equation for this reaction and add the state symbols of the reactants and products.

Long Answer Type Questions

33. (*a*) What is meant by a chemical reaction ? Explain with the help of an example.

(*b*) Give one example each of a chemical reaction characterised by :

(*i*) evolution of a gas
(*ii*) change in colour
(*iii*) formation of a precipitate
(*iv*) change in temperature
(*v*) change in state.

34. (*a*) State the various characteristics of chemical reactions.
(*b*) State one characteristic each of the chemical reaction which takes place when :
(*i*) dilute hydrochloric acid is added to sodium carbonate
(*ii*) lemon juice is added gradually to potassium permanganate solution
(*iii*) dilute sulphuric acid is added to barium chloride solution
(*iv*) quicklime is treated with water
(*v*) wax is burned in the form of a candle

35. (*a*) What do you understand by exothermic and endothermic reactions ?
(*b*) Give one example of an exothermic reaction and one of an endothermic reaction.
(*c*) Which of the following are endothermic reactions and which are exothermic reactions ?
(*i*) Burning of natural gas (*ii*) Photosynthesis
(*iii*) Electrolysis of water (*iv*) Respiration
(*v*) Decomposition of calcium carbonate

Multiple Choice Questions (MCQs)

36. One of the following does not happen during a chemical reaction. This is :
(*a*) Breaking of old chemical bonds and formation of new chemical bonds
(*b*) Formation of new substances with entirely different properties
(*c*) Atoms of one element change into those of another element to form new products.
(*d*) A rearrangement of atoms takes place to form new products.

37. Which of the following does not involve a chemical reaction ?
(*a*) digestion of food in our body
(*b*) process of respiration
(*c*) burning of candle wax when heated
(*d*) melting of candle wax on heating

38. You are given the solution of lead nitrate. In order to obtain a yellow precipitate you should mix with it a solution of :
(*a*) potassium chloride (*b*) potassium nitride
(*c*) potassium sulphide (*d*) potassium iodide

39. An acid which can decolourise purple coloured potassium permanganate solution is :
(*a*) sulphuric acid (*b*) citric acid
(*c*) carbonic acid (*d*) hydrochloric acid

40. The chemical reaction between two substances is characterised by a change in colour from orange to green. These two substances are most likely to be :
(*a*) potassium dichromate solution and sulphur dioxide
(*b*) potassium permanganate solution and sulphur dioxide
(*c*) potassium permanganate solution and lemon juice
(*d*) potassium dichromate solution and carbon dioxide.

41. The chemical reaction between quicklime and water is characterised by :
(*a*) evolution of hydrogen gas
(*b*) formation of slaked lime precipitate
(*c*) change in temperature of mixture
(*d*) change in colour of the product

42. One of the following is an endothermic reaction. This is :
(*a*) combination of carbon and oxygen to form carbon monoxide
(*b*) combination of nitrogen and oxygen to form nitrogen monoxide
(*c*) combination of glucose and oxygen to form carbon dioxide and water
(*d*) combination of zinc and hydrochloric acid to form zinc chloride and hydrogen

43. Which of the following is not an endothermic reaction ?
(*a*) $CaCO_3 \longrightarrow CaO + CO_2$

(*b*) $2H_2O \longrightarrow 2H_2 + O_2$
(*c*) $6CO_2 + 6H_2O \longrightarrow C_6H_{12}O_6 + 6O_2$
(*d*) $C_6H_{12}O_6 + 6O_2 \longrightarrow 6CO_2 + 6H_2O$

44. One of the following is an exothermic reaction. This is :
(*a*) electrolysis of water
(*b*) conversion of limestone into quicklime
(*c*) process of respiration
(*d*) process of photosynthesis

45. The chemical equations are balanced to satisfy one of the following laws in chemical reactions. This law is known as :
(*a*) law of conservation of momentum
(*b*) law of conservation of mass
(*c*) law of conservation of motion
(*d*) law of conservation of magnetism

Questions Based on High Order Thinking Skills (HOTS)

46. When the solution of substance X is added to a solution of potassium iodide, then a yellow solid separates out from the solution.
(*a*) What do you think substance X is likely to be ?
(*b*) Name the substance which the yellow solid consists of.
(*c*) Which characteristic of chemical reactions is illustrated by this example ?
(*d*) Write a balanced chemical equation for the reaction which takes place. Mention the physical states of all the reactants and products involved in the chemical equation.

47. When water is added gradually to a white solid X, a hissing sound is heard and a lot of heat is produced forming a product Y. A suspension of Y in water is applied to the walls of a house during white washing. A clear solution of *Y* is also used for testing carbon dioxide gas in the laboratory.
(*a*) What could be solid X ? Write its chemical formula.
(*b*) What could be product Y ? Write its chemical formula.
(*c*) What is the common name of the solution of Y which is used for testing carbon dioxide gas ?
(*d*) Write chemical equation of the reaction which takes place on adding water to solid X.
(*e*) Which characteristic of chemical reactions is illustrated by this example ?

48. When metal X is treated with a dilute acid Y, then a gas Z is evolved which burns readily by making a little explosion.
(*a*) Name any two metals which can behave like metal X.
(*b*) Name any two acids which can behave like acid Y.
(*c*) Name the gas Z.
(*d*) Is the gas Z lighter than or heavier than air ?
(*e*) Is the reaction between metal X and dilute acid Y exothermic or endothermic ?
(*f*) By taking a specific example of metal X and dilute acid Y, write a balanced chemical equation for the reaction which takes place. Also indicate physical states of all the reactants and products.

49. A solid substance P which is very hard is used in the construction of many buildings, especially flooring. When substance P is heated strongly, it decomposes to form another solid Q and a gas R is given out. Solid Q reacts with water with the release of a lot of heat to form a substance S. When gas R is passed into a clear solution of substance S, then a white precipitate of substance T is formed. The substance T has the same chemical composition as starting substance P.
(*a*) What is substance P ? Write its common name as well as chemical formula.
(*b*) What is substance Q ?
(*c*) What is gas R ?
(*d*) What is substance S ? What is its clear solution known as ?
(*e*) What is substance T ? Name any two natural forms in which substance T occurs in nature.

50. A silvery-white metal X taken in the form of ribbon, when ignited, burns in air with a dazzling white flame to form a white powder Y. When water is added to powder Y, it dissolves partially to form another substance Z.

(*a*) What could metal X be ?

(*b*) What is powder Y ?

(*c*) With which substance metal X combines to form powder Y ?

(*d*) What is substance Z ? Name one domestic use of substance Z.

(*e*) Write a balanced chemical equation of the reaction which takes place when metal X burns in air to form powder Y.

51. A metal X forms a salt XSO_4. The salt XSO_4 forms a clear solution in water which reacts with sodium hydroxide solution to form a blue precipitate Y. Metal X is used in making electric wires and alloys like brass.

(*a*) What do you think metal X could be ?

(*b*) Write the name, formula and colour of salt XSO_4.

(*c*) What is the blue precipitate Y ?

(*d*) Write a chemical equation of the reaction which takes place when salt XSO_4 reacts with sodium hydroxide solution. Give the state symbols of all the reactants and products which occur in the above equation.

52. The metal M reacts vigorously with water to form a solution S and a gas G. The solution S turns red litmus to blue whereas gas G, which is lighter than air, burns with a pop sound. Metal M has a low melting point and it is used as a coolant in nuclear reactors.

(*a*) What is metal M ?

(*b*) What is solution S ? Is it acidic or alkaline ?

(*c*) What is gas G ?

(*d*) Write a balanced chemical equation for the reaction which takes place when metal M reacts with water.

(*e*) Is this reaction exothermic or endothermic ?

53. When a mixture of gases X and Y is compressed to 300 atm pressure and then passed over a catalyst consisting of a mixture of zinc oxide and chromium oxide (heated to a temperature of 300°C), then an organic compound Z having the molecular formula CH_4O is formed. X is a highly poisonous gas which is formed in appreciable amounts when a fuel burns in a limited supply of air ; Y is a gas which can be made by the action of a dilute acid on an active metal ; and Z is a liquid organic compound which can react with sodium metal to produce hydrogen gas.

(*a*) What are X, Y and Z ?

(*b*) Write a balanced chemical equation of the reaction which takes place when X and Y combine to form Z. Indicate the conditions under which the reaction occurs.

54. The white solid compound A decomposes quite rapidly on heating in the presence of a black substance X to form a solid compound B and a gas C. When an aqueous solution of compound B is reacted with silver nitrate solution, then a white precipitate of silver chloride is obtained alongwith potassium nitrate solution. Gas C does not burn itself but helps burn other things.

(*a*) What is compound A ?

(*b*) What is compound B ?

(*c*) What is gas C ?

(*d*) What do you think is the black substance X ? What is its function ?

(*e*) What is the general name of substances like X ?

55. Gas A, which is the major cause of global warming, combines with hydrogen oxide B in nature in the presence of an environmental factor C and a green material D to form a six carbon organic compound E and a gas F. The gas F is necessary for breathing.

(*a*) What is gas A ?

(*b*) What is the common name of B ?

(*c*) What do you think could be C ?

(*d*) What is material D ? Where is it found ?

(*e*) Name the organic compound E.
(*f*) What is gas F ? Name the natural process during which it is released.

ANSWERS

5. False **6.** Oxygen should be in molecular form, O_2; $2Mg + O_2 \longrightarrow 2MgO$
10. (*a*) $2H_2S + 3O_2 \longrightarrow 2H_2O + 2SO_2$ (*b*) $P_4 + 5O_2 \longrightarrow 2P_2O_5$
(*c*) $CS_2 + 3O_2 \longrightarrow CO_2 + 2SO_2$ (*d*) $2Al + Fe_2O_3 \longrightarrow Al_2O_3 + 2Fe$
(*e*) $BaCl_2 + ZnSO_4 \longrightarrow ZnCl_2 + BaSO_4$
11. (*a*) $Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$ (*b*) $2Al + 3CuCl_2 \longrightarrow 2AlCl_3 + 3Cu$
12. (*a*) $2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2H_2O$ (*b*) $Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$
13. (*i*) $Ca + 2H_2O \longrightarrow Ca(OH)_2 + H_2$ (*ii*) $N_2 + 3H_2 \longrightarrow 2NH_3$
14. (*a*) $Ca\ (s) + 2H_2O\ (l) \longrightarrow Ca(OH)_2\ (aq) + H_2\ (g)$ (*b*) $2SO_2\ (g) + O_2\ (g) \longrightarrow 2SO_3\ (g)$
15. (*i*) $4Na + O_2 \longrightarrow 2Na_2O$ (*ii*) $2H_2O_2 \longrightarrow 2H_2O + O_2$
(*iii*) $Mg(OH)_2 + 2HCl \longrightarrow MgCl_2 + 2H_2O$. (*iv*) $4Fe + 3O_2 \longrightarrow 2Fe_2O_3$
(*v*) $2Al(OH)_3 \longrightarrow Al_2O_3 + 3H_2O$ (*vi*) $2NH_3 + 3CuO \longrightarrow 3Cu + N_2 + 3H_2O$
(*vii*) $Al_2(SO_4)_3 + 6NaOH \longrightarrow 2Al(OH)_3 + 3Na_2SO_4$
(*viii*) $2HNO_3 + Ca(OH)_2 \longrightarrow Ca(NO_3)_2 + 2H_2O$ (*ix*) $2NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2H_2O$
(*x*) $BaCl_2 + H_2SO_4 \longrightarrow BaSO_4 + 2HCl$
16 (*a*) conservation of mass (*b*) aqueous ; (aq) **18.** (*c*) (*i*) $2NH_3 \longrightarrow N_2 + 3H_2$ (*ii*) $C + CO_2 \longrightarrow 2CO$
19. $H_2 + CuO \longrightarrow Cu + H_2O$
(*i*) Elements : H_2 and Cu (*ii*) Compounds : CuO and H_2O
(*iii*) Reactants : H_2 and CuO (*iv*) Products : Cu and H_2O
(*v*) Metal : Cu (*vi*) Non-metal : H_2
20. (*b*) $Ca(OH)_2\ (aq) + CO_2\ (g) \longrightarrow CaCO_3\ (s) + H_2O\ (l)$
21. (*b*) $2Al + 3Cl_2 \longrightarrow 2AlCl_3$ (*c*) $2K + 2H_2O \longrightarrow 2KOH + H_2$
22. (*b*) $Zn\ (s) + 2HCl\ (aq) \longrightarrow ZnCl_2\ (aq) + H_2\ (g)$
(*c*) $FeSO_4\ (aq) + 2NaOH\ (aq) \longrightarrow Fe(OH)_2\ (s) + Na_2SO_4\ (aq)$
24. (*a*) $2Al(OH)_3 + 3H_2SO_4 \longrightarrow Al_2(SO_4)_3 + 6H_2O$
(*b*) $MnO_2 + 4HCl \longrightarrow MnCl_2 + Cl_2 + 2H_2O$
25. (*a*) $MgCO_3\ (s) + 2HCl\ (aq) \longrightarrow MgCl_2\ (aq) + CO_2\ (g) + H_2O\ (l)$
(*b*) $2NaOH\ (aq) + H_2SO_4\ (aq) \longrightarrow Na_2SO_4\ (aq) + 2H_2O\ (l)$
27. (*a*) $2KClO_3\ (s) \longrightarrow 2KCl\ (s) + 3O_2\ (g)$
(*b*) $2Mg + CO_2 \longrightarrow 2MgO + C$
28. (*a*) $CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$
(*b*) $NaOH\ (aq) + HCl\ (aq) \longrightarrow NaCl\ (aq) + H_2O\ (l)$
29. $4NH_3\ (g) + 3O_2\ (g) \longrightarrow 2N_2\ (g) + 6H_2O\ (l)$
31. $3BaCl_2\ (aq) + Al_2(SO_4)_3\ (aq) \longrightarrow 3BaSO_4\ (s) + 2AlCl_3\ (aq)$
32. $2KNO_3\ (s) \longrightarrow 2KNO_2\ (s) + O_2\ (g)$
35. (*c*) Endothermic reactions : Photosynthesis, Electrolysis of water, Decomposition of calcium carbonate ; Exothermic reactions : Burning of natural gas , Respiration **36.** (*c*) **37.** (*d*) **38.** (*d*) **39.** (*b*) **40.** (*a*) **41.** (*c*) **42.** (*b*) **43.** (*d*) **44.** (*c*) **45.** (*b*) **46.** (*a*) Lead nitrate (*b*) Lead iodide (*c*) Formation of a precipitate (*d*) $Pb(NO_3)_2\ (aq) + 2KI\ (aq) \rightarrow PbI_2\ (s) + 2KNO_3\ (aq)$ **47.** (*a*) Calcium oxide, CaO (*b*) Calcium hydroxide, $Ca(OH)_2$ (*c*) Lime water (*d*) $CaO + H_2O \rightarrow Ca(OH)_2$ (*e*) Change in temperature **48.** (*a*) Zinc and Iron (*b*) Dilute hydrochloric acid and Dilute sulphuric acid (*c*) Hydrogen (*d*) Lighter than air (*e*) Exothermic (*f*) Suppose metal X is zinc (Zn) and acid Y is dilute hydrochloric acid (HCl) ; $Zn\ (s) + 2HCl\ (aq) \rightarrow ZnCl_2\ (aq) + H_2\ (g)$ **49.** (*a*) Calcium carbonate, Limestone, $CaCO_3$ (*b*) Calcium oxide, CaO (*c*) Carbon dioxide, CO_2 (*d*) Calcium hydroxide, $Ca(OH)_2$; Lime water (*e*) Calcium carbonate ; Limestone and Marble **50.** (*a*) Magnesium, Mg (*b*) Magnesium oxide, MgO (*c*) Oxygen (of air), O_2 (*d*) Magnesium hydroxide, $Mg(OH)_2$; Used as antacid to relieve indigestion (*e*) $2Mg + O_2 \rightarrow 2MgO$ **51.** (*a*) Copper, Cu (*b*) Copper sulphate, $CuSO_4$, Blue colour (*c*) Copper hydroxide, $Cu(OH)_2$ (*d*) $CuSO_4\ (aq) + 2NaOH\ (aq) \longrightarrow Cu(OH)_2\ (s) + Na_2SO_4\ (aq)$

52. (*a*) Sodium, Na (*b*) Sodium hydroxide solution (NaOH solution), Alkaline (*c*) Hydrogen, H_2 (*d*) $2Na + 2H_2O \longrightarrow 2NaOH + H_2$ (*e*) Exothermic **53.** (*a*) X is carbon monoxide gas (CO); Y is hydrogen gas (H_2) ; Z is methanol (or Methyl alcohol) (CH_3OH) **Note.** The molecular formula CH_4O for compound Z tells us that one molecule of this compound contains 1 carbon atom, 4 hydrogen atoms and 1 oxygen atom ($CH_4O = CH_3OH$) (*b*) For chemical equation, see page 11 of this book **54.** (*a*) Potassium chlorate, $KClO_3$ (*b*) Potassium chloride, KCl (*c*) Oxygen, O_2 (*d*) Manganese dioxide, MnO_2 ; It acts as a catalyst in the decomposition of potassium chlorate to form oxygen gas (*e*) Catalysts **55.** (*a*) Carbon dioxide, CO_2 (*b*) Water, H_2O (*c*) Sunlight (*d*) Chlorophyll ; Green leaves of plants (*e*) Glucose, $C_6H_{12}O_6$ (*f*) Oxygen; Photosynthesis

TYPES OF CHEMICAL REACTIONS

Some of the important types of chemical reactions are :

1. **Combination reactions,**
2. **Decomposition reactions,**
3. **Displacement reactions,**
4. **Double displacement reactions,** and
5. **Oxidation and Reduction reactions.**

We will now discuss all these reactions in detail, one by one. Let us start with the combination reactions.

1. COMBINATION REACTIONS

Those reactions in which two or more substances combine to form a single substance, are called combination reactions. In a combination reaction, two or more elements can combine to form a compound ; two or more compounds can combine to form a new compound ; or an element and a compound can combine to form a new compound. We will now give some examples of combination reactions.

Example 1. Magnesium and oxygen combine, when heated, to form magnesium oxide :

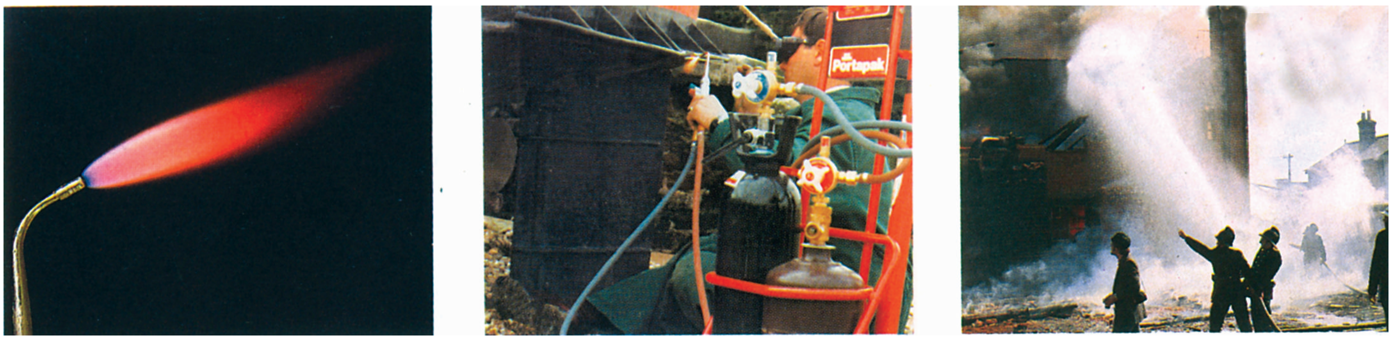
$$\underset{\text{Magnesium}}{2Mg\,(s)} + \underset{\text{Oxygen}}{O_2\,(g)} \xrightarrow{\text{Combination}} \underset{\text{Magnesium oxide}}{2MgO\,(s)}$$

In this reaction, two elements, magnesium and oxygen, are combining to form a single compound, magnesium oxide. So, this is a combination reaction. Thus, when we burn a magnesium ribbon (or magnesium wire) in air, then a combination reaction takes place with oxygen to form magnesium oxide.

Example 2. Hydrogen burns in oxygen to form water :

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

In this reaction, two elements, hydrogen and oxygen, are combining to form a single compound, water, so this is an example of a combination reaction. Thus, **the formation of water from hydrogen and oxygen is a combination reaction.**



(*a*) Hydrogen is an element. It is a gas which burns explosively.

(*b*) Oxygen is another element. It is a gas which does not burn itself but helps other things to burn.

(*c*) Hydrogen and oxygen combine to form a compound 'hydrogen oxide' which is commonly known as water. It puts out fire.

Figure 24.

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

$$\underset{\substack{\text{Carbon}\\ \text{(Coal)}}}{C\,(s)} + \underset{\substack{\text{Oxygen}\\ \text{(From air)}}}{O_2\,(g)} \xrightarrow{\text{Combination}} \underset{\text{Carbon dioxide}}{CO_2\,(g)}$$

In this reaction two elements, carbon and oxygen, are combining together to form a single compound, carbon dioxide. So, this is a combination reaction. Please note that when carbon (in the form of coal) burns in air then the carbon combines only with the oxygen present in air to form carbon dioxide gas. Thus, the burning of coal in air is an example of a combination reaction.

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

$$\underset{\text{Hydrogen}}{H_2\,(g)} + \underset{\text{Chlorine}}{Cl_2\,(g)} \xrightarrow{\text{Combination}} \underset{\text{Hydrogen chloride}}{2HCl\,(g)}$$

Here, two elements, hydrogen and chlorine, react together to form a single compound, hydrogen chloride gas. So, this is an example of combination reaction. This combination reaction is used in industry for the manufacture of hydrochloric acid (Hydrogen chloride gas on dissolving in water forms hydrochloric acid).

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

$$\underset{\text{Sodium}}{2Na\,(s)} + \underset{\text{Chlorine}}{Cl_2\,(g)} \xrightarrow{\text{Combination}} \underset{\text{Sodium chloride}}{2NaCl\,(s)}$$

In this example, two elements, sodium and chlorine, are combining together to form a single compound, sodium chloride. So, this is a combination reaction.

Figure 25. Sodium burns in chlorine to form sodium chloride.

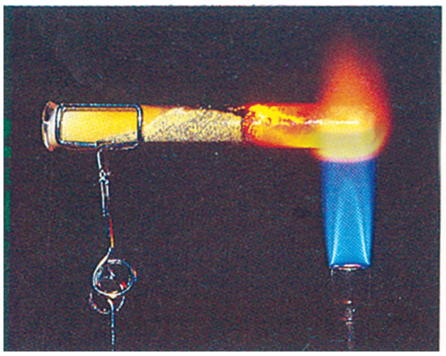


Figure 26. Iron powder and sulphur being heated together.

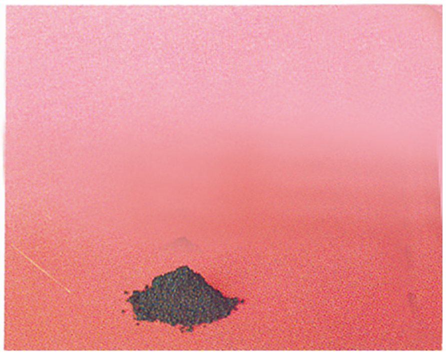


Figure 27. Iron sulphide compound is formed.

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

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

In this reaction, two elements, iron and sulphur, are reacting together to form a single compound, iron sulphide, so it is a combination reaction.

In all the above examples, **two elements combine to form a single compound.** In some combination reactions, however, **two or more compounds combine together to form a new compound.** This point will become more clear from the following example

Example 7. Calcium oxide (lime or quicklime) reacts vigorously with water to form calcium hydroxide (slaked lime) :

$$\underset{\substack{\text{Calcium oxide}\\ \text{(Lime or Quicklime)}}}{CaO\,(s)} + \underset{\text{Water}}{H_2O\,(l)} \xrightarrow{\text{Combination}} \underset{\substack{\text{Calcium hydroxide}\\ \text{(Slaked lime)}}}{Ca(OH)_2\,(s)}$$

This is a combination reaction in which two compounds, calcium oxide and water, combine to form a single compound calcium hydroxide. A large amount of heat is released when calcium oxide reacts with water to form calcium hydroxide (or slaked lime). (We have already carried out this reaction on page 5). Please note that it is solid calcium hydroxide which is known as slaked lime. Slaked lime is a white powder.

Discussion. The substance which we use for white-washing our house is lime (or quicklime) which is calcium oxide (CaO). We put calcium oxide in a drum and add water to it slowly. Calcium oxide reacts with water vigorously to form a white solid called calcium hydroxide (or slaked lime) with the evolution of heat. More water is then added to get calcium hydroxide solution. This calcium hydroxide solution is then applied to the walls of the house with a brush.

The calcium hydroxide solution, when applied to the walls, reacts slowly with the carbon dioxide gas present in air to form a thin, shining layer of calcium carbonate on the walls of the house :

$$\underset{\text{Calcium hydroxide}}{Ca(OH)_2\ (aq)} + \underset{\substack{\text{Carbon dioxide} \\ \text{(From air)}}}{CO_2\ (g)} \longrightarrow \underset{\text{Calcium carbonate}}{CaCO_3\ (s)} + \underset{\text{Water}}{H_2O\ (l)}$$

Since this process gives a white, shiny appearance to the walls of a house, it is called white-washing. The calcium carbonate is actually formed after two to three days of white-washing and gives a shiny finish to the walls.

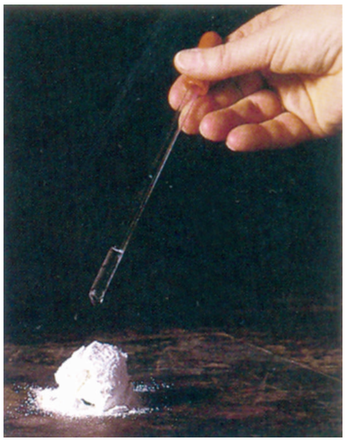


Figure 28. Calcium oxide reacts vigorously with water to form calcium hydroxide.

Figure 29. Calcium hydroxide solution reacts with carbon dioxide gas to form a white solid, calcium carbonate.

Figure 30. Ammonia combines with hydrochloric acid to form dense white fumes of ammonium chloride.

Example 8. Ammonia reacts with hydrogen chloride to form ammonium chloride. This can be written as :

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

In this reaction, two compounds, ammonia and hydrogen chloride, combine together to produce a new compound, ammonium chloride. So, this is a combination reaction.

We will now give some examples of those combination reactions in which **a compound reacts with an element to form a new compound.**

Example 9. Carbon monoxide reacts with oxygen to form carbon dioxide :

$$\underset{\text{Carbon monoxide}}{2CO\ (g)} + \underset{\text{Oxygen}}{O_2\ (g)} \xrightarrow{\text{Combination}} \underset{\text{Carbon dioxide}}{2CO_2\ (g)}$$

In this reaction, carbon monoxide compound reacts with oxygen element to form a new compound, carbon dioxide. So, this is a combination reaction.

Example 10. Sulphur dioxide reacts with oxygen to produce sulphur trioxide. This reaction can be written as :

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

In this combination reaction, a compound, sulphur dioxide, combines with an element, oxygen, to form a new compound, sulphur trioxide.

2. DECOMPOSITION REACTIONS

Those reactions in which a compound splits up into two or more simpler substances are known as decomposition reactions. The decomposition reactions are carried out by applying heat, light or electricity. Heat, light or electricity provide energy which breaks a compound into two or more simpler compounds. Please note that **a decomposition reaction is just the opposite of a combination reaction.** We will now give some examples of decomposition reactions.

Example 1. When calcium carbonate is heated, it decomposes to give calcium oxide and carbon dioxide :

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

In this reaction, one substance, calcium carbonate, is breaking up into two simpler substances, calcium oxide and carbon dioxide, so this is a decomposition reaction. Please note that calcium carbonate is also called 'limestone' and calcium oxide formed from it is called 'lime' (or quicklime). The decomposition of calcium carbonate (limestone) on heating is an important reaction used in various industries. This is because calcium oxide (lime) obtained by the decomposition of calcium carbonate has many uses in industry. For example, calcium oxide (or lime) is used on a large scale in the manufacture of cement and glass.

Figure 31. This is calcium carbonate (as limestone).

Figure 32. This is calcium oxide (or lime).

Figure 33. This is a lime kiln which decomposes calcium carbonate (limestone) into calcium oxide (or lime).

When a decomposition reaction is carried out by heating, it is called 'thermal decomposition'. ('Thermal' means 'relating to heat'). The decomposition of calcium carbonate into calcium oxide and carbon dioxide is an example of thermal decomposition (because it is carried out by heating).

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

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

This decomposition takes place in the presence of heat and catalyst. In this decomposition reaction, a single compound, potassium chlorate, is splitting up into two simpler substances, potassium chloride and oxygen. This decomposition reaction is used for preparing oxygen gas in the laboratory.

Example 3. When ferrous sulphate is heated strongly, it decomposes to form ferric oxide, sulphur dioxide and sulphur trioxide :

$$\underset{\substack{\text{Ferrous sulphate}\\ \text{(Green colour)}}}{2FeSO_4\,(s)} \xrightarrow[\text{(Decomposition)}]{\text{Heat}} \underset{\substack{\text{Ferric oxide}\\ \text{(Brown colour)}}}{Fe_2O_3\,(s)} + \underset{\substack{\text{Sulphur}\\ \text{dioxide}}}{SO_2\,(g)} + \underset{\substack{\text{Sulphur}\\ \text{trioxide}}}{SO_3\,(g)}$$

In this reaction, the green colour of ferrous sulphate changes to brown due to the formation of ferric oxide. A smell of burning sulphur is obtained due to the formation of sulphur dioxide gas. In this reaction, one substance is splitting up into three substances, so this is a decomposition reaction. It is actually a thermal decomposition reaction. Please note that ferrous sulphate is also known as iron (II) sulphate (or just iron sulphate). And ferric oxide is also known as iron (III) oxide.

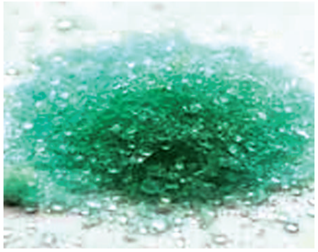


Figure 34. These are ferrous sulphate crystals.

Figure 35. This is ferric oxide (formed by the decomposition of ferrous sulphate).

The ferrous sulphate crystals which are available in a science laboratory are actually ferrous sulphate heptahydrate, $FeSO_4.7H_2O$. They contain 7 molecules of water of crystallisation. These crystals are green in colour. When the green coloured ferrous sulphate heptahydrate crystals ($FeSO_4.7H_2O$) are heated, they first lose 7 molecules of water of crystallisation to form anhydrous ferrous sulphate ($FeSO_4$) which is white in colour. And then this anhydrous ferrous sulphate decomposes to give ferric oxide, sulphur dioxide and sulphur trioxide. In the above equation, we have written ferrous sulphate crystals without water of crystallisation just to keep the equation simple.

We can carry out the decomposition reaction of ferrous sulphate in the laboratory as follows :

(*i*) Take about 2 grams of ferrous sulphate crystals in a dry boiling tube. The ferrous sulphate crystals are green in colour.

(*ii*) Heat the boiling tube over a burner (by keeping the mouth of boiling tube away from yourself and your neighbour working in the laboratory) (see Figure 36).

(*iii*) The green colour of ferrous sulphate crystals first changes to white and then a brown solid is formed (which is ferric oxide).

(*iv*) Gas having the smell of burning sulphur comes out of the boiling tube (We should smell the gas by turning it gently towards our nose with a blow of our hand and not by bringing the mouth of boiling tube under our nose).

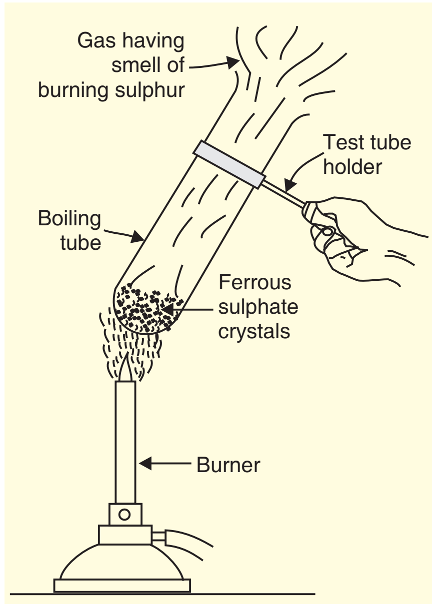


Figure 36. Decomposition reaction of ferrous sulphate.

Example 4. When lead nitrate is heated strongly, it breaks down to form simpler substances like lead monoxide, nitrogen dioxide and oxygen. This can be written as :

$$\underset{\substack{\text{Lead nitrate}\\ \text{(Colourless)}}}{2Pb(NO_3)_2\,(s)} \xrightarrow[\text{(Decomposition)}]{\text{Heat}} \underset{\substack{\text{Lead monoxide}\\ \text{(Yellow)}}}{2PbO\,(s)} + \underset{\substack{\text{Nitrogen dioxide}\\ \text{(Brown fumes)}}}{4NO_2\,(g)} + \underset{\text{Oxygen}}{O_2\,(g)}$$

In this decomposition reaction, the colourless compound lead nitrate forms a yellow compound, lead monoxide, and brown fumes of nitrogen dioxide gas are evolved. Here, one compound, lead nitrate, is breaking down to form three compounds, lead monoxide, nitrogen dioxide and oxygen, so it is a decomposition reaction. Since the decomposition of lead nitrate is brought about by heat, therefore, it is actually an

Figure 37. Lead nitrate.

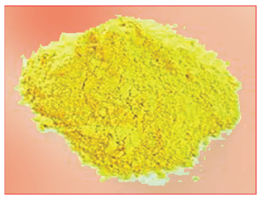


Figure 38. Lead monoxide (formed by the decomposition of lead nitrate).

example of thermal decomposition.

We can carry out the decomposition reaction of lead nitrate in the laboratory as follows :

(*i*) Take about 2 grams of lead nitrate powder in a boiling tube. Lead nitrate is a colourless compound.

(*ii*) Hold the boiling tube in a test-tube holder and heat it over a burner (see Figure 39).

(*iii*) Brown fumes of nitrogen dioxide gas are evolved which fill the boiling tube.

(*iv*) If a glowing splinter is held over the mouth of the boiling tube, it catches fire and starts burning again. This shows that oxygen gas is also evolved during this reaction.

(*v*) A yellow solid is left behind in the boiling tube. This is lead monoxide (Please note that lead monoxide is reddish-brown when hot but yellow when cold).

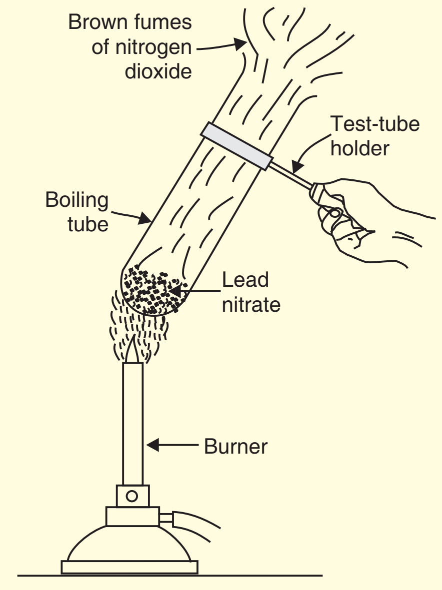


Figure 39. Decomposition reaction of lead nitrate.

All the above decomposition reactions have been carried out **by the action of heat.** We will now give some examples of those decomposition reactions which are carried out **by using electricity**.

Example 5. When electric current is passed through acidified water, it decomposes to give hydrogen gas and oxygen gas. This reaction can be represented as :

$$\underset{\text{Water}}{2H_2O\ (l)} \xrightarrow[\text{(Decomposition)}]{\text{Electricity}} \underset{\text{Hydrogen}}{2H_2\ (g)} + \underset{\text{Oxygen}}{O_2\ (g)}$$

In this decomposition reaction, a single compound, water, splits up to form two simpler substances, hydrogen and oxygen. This decomposition reaction takes place by the action of electricity. It is called electrolysis of water.

We can carry out the electrolysis of water as follows :

(*i*) Take a wide-mouthed glass bottle *B* (with bottom removed). Fix it on a stand in the inverted position as shown in Figure 40.

(*ii*) A rubber cork having two holes is fitted in the neck of the bottle. Two carbon rods (called carbon electrodes) are fixed in the two holes of the cork tightly (The 'carbon rods' are actually 'graphite rods').

(*iii*) Fill the glass bottle two-thirds with water. Add a few drops of dilute sulphuric acid to water (to make water a good conductor of electricity).

(*iv*) Two similar test-tubes filled with water are carefully inverted over the two carbon electrodes by keeping thumb over their mouth so that initially they remain completely filled with water.

(*v*) Connect the outer ends of carbon rods to the two terminals (+ and –) of a 6 volt battery by wires having a switch in them (see Figure 40). The left side carbon rod

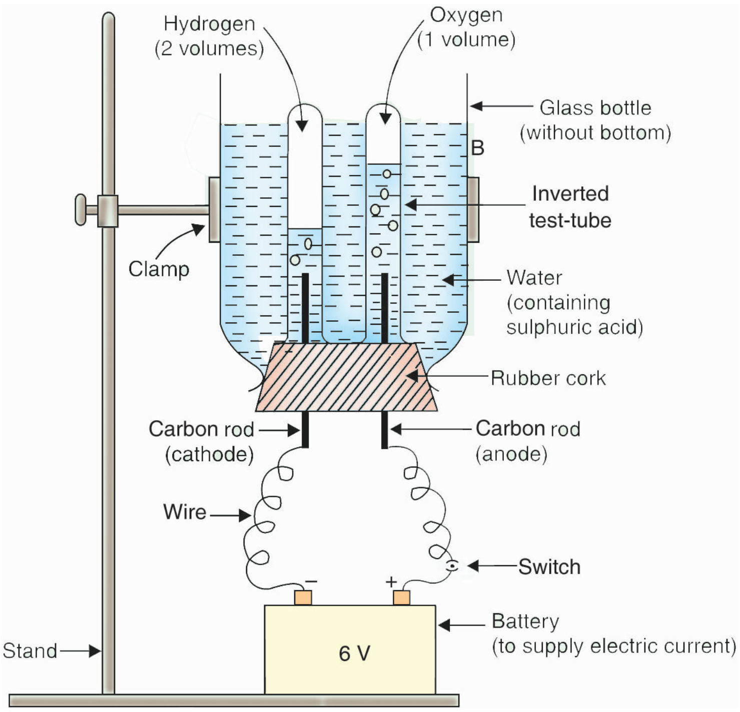


Figure 40. Experimental set-up for the electrolysis of water.

connected to the negative terminal of the battery is called cathode (negative electrode). The right side carbon rod connected to the positive terminal of the battery is called anode (positive electrode).

(*vi*) Pass in electric current through water by turning on the switch and leave the apparatus undisturbed for some time.

(*vii*) We will see the bubbles of gases being formed at both the carbon electrodes inside the test-tubes containing water (see Figure 40). These gases are formed by the decomposition of water on passing electricity.

(*viii*) The gases formed at the two electrodes go on collecting in the top parts of the inverted test-tubes (and the water level in these test-tubes falls gradually).

(*ix*) The volume of gases collected in the two test-tubes is not the same. The volume of gas collected on the negative electrode (left electrode) is double the volume of gas collected on the positive electrode (right electrode) (see Figure 40).

(*x*) Keep on passing electric current till both the test-tubes are completely filled with respective gases. Then remove the gas-filled test-tubes carefully and test them one by one by bringing a burning candle close to the mouth of each test-tube.

(*xi*) When a burning candle is brought near the mouth of left test-tube, the gas in it burns rapidly making a 'popping sound' (or 'little explosion'). We know that hydrogen gas burns with a popping sound. So, the gas collected in the left test-tube over negative electrode (which had double volume or 2 volumes) is hydrogen.

(*xii*) When the burning candle is taken near the mouth of the right side test-tube, the candle starts burning brightly. We know that oxygen gas makes things burn brightly. So, the gas collected in the right side test-tube over positive electrode (which had 1 volume) is oxygen.

Figure 41. A student carrying out the electrolysis of water in the laboratory.

Since the electrolysis of water produces 2 volumes of hydrogen gas and 1 volume of oxygen gas, we conclude that the ratio of hydrogen and oxygen elements in water is 2 : 1 by volume. In other words, electrolysis of water shows that water is a compound made up of 2 parts of hydrogen and 1 part of oxygen (by volume). So, the formula of water is H_2O.

Please note that when hydrogen burns in oxygen, water is formed (This is a combination reaction). And when water is electrolysed, then hydrogen and oxygen are formed (This is a decomposition reaction). These examples show that **a decomposition reaction is just the opposite of a combination reaction.**

Example 6. When electric current is passed through molten sodium chloride, it decomposes to give sodium metal and chlorine gas :

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

This decomposition reaction is used to obtain sodium metal from sodium chloride (common salt). It is called electrolysis of molten sodium chloride.

Example 7. When electric current is passed through molten aluminium oxide, it decomposes to give aluminium metal and oxygen gas :

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

This decomposition reaction is used to extract aluminium metal from aluminium oxide. It is called electrolysis of molten aluminium oxide.

We have just discussed those decomposition reactions which are caused by electric energy or electricity. We will now describe some decomposition reactions which are brought about **by light energy.**

Example 8. When silver chloride is exposed to light, it decomposes to form silver metal and chlorine gas :

$$\underset{\substack{\text{Silver chloride} \\ \text{(White)}}}{2AgCl\,(s)} \xrightarrow[\text{(Decomposition)}]{\text{Light}} \underset{\substack{\text{Silver} \\ \text{(Greyish white)}}}{2Ag\,(s)} + \underset{\substack{\text{Chlorine} \\ \text{(Yellowish-green)}}}{Cl_2\,(g)}$$

In this reaction, the white colour of silver chloride changes to greyish white due to the formation of silver metal. The decomposition of silver chloride is caused by light (It may be sunlight or bulb light). This reaction is used in black and white photography. We can carry out the decomposition reaction of silver chloride as follows :

(*i*) Take about 2 grams of silver chloride in a china dish. It is white in colour.

(*ii*) Place this china dish in sunlight for some time (see Figure 42).

(*iii*) We will find that white silver chloride turns greyish white (due to the formation of silver metal).

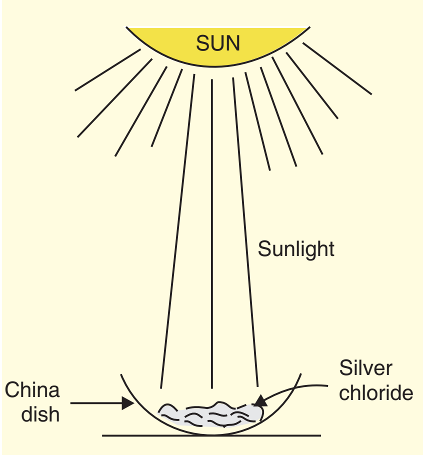


Figure 42. Decomposition of silver chloride is caused by light

Silver bromide also behaves in the same way as silver chloride with light energy. Thus, when silver bromide is exposed to light, it decomposes to form silver metal and bromine vapours :

$$\underset{\substack{\text{Silver bromide} \\ \text{(Pale yellow)}}}{2AgBr\,(s)} \xrightarrow[\text{(Decomposition)}]{\text{Light}} \underset{\substack{\text{Silver} \\ \text{(Greyish white)}}}{2Ag\,(s)} + \underset{\substack{\text{Bromine} \\ \text{(Red-brown)}}}{Br_2\,(g)}$$

In this reaction, pale yellow colour of silver bromide changes to greyish white due to the formation of silver metal. The decomposition of silver bromide is caused by light. The light may be sunlight or bulb light. This reaction of decomposition of silver bromide is also used in black and white photography.

Figure 43. Silver bromide is used in photographic film.

Uses of Decomposition Reactions. The decomposition reactions carried out by electricity are used to extract several metals from their naturally occurring compounds like chlorides or oxides. When the fused (molten) metal chloride or metal oxide is decomposed by passing electricity, then metal is produced at the cathode (negative electrode). For example, sodium metal is extracted by the electrolysis of molten sodium chloride whereas aluminium metal is extracted by the electrolysis of molten aluminium oxide (see examples 6 and 7 given on page 31).

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 the proteins decompose to form amino acids.

We will now answer some questions based on combination and decomposition reactions :

Sample Problem 1. A solution of substance *X* is used for white-washing.

(*i*) Name the substance *X* and write its formula.

(*ii*) Write the reaction of substance *X* with water. **(NCERT Book Question)**

Solution. (*i*) The substance whose solution in water we use for white-washing is calcium oxide (lime, *choona*). So, the substance *X* is calcium oxide. Its formula is CaO.

(*ii*) Write the equation for the reaction of calcium oxide with water yourself (see page 26).

Sample Problem 2. Why is double the amount of a gas collected in one of the test-tubes in the electrolysis of water experiment ? Name this gas. **(NCERT Book Question)**

Solution. The gas which is collected in double the amount in the electrolysis of water experiment is hydrogen. This is because water (H_2O) contains 2 parts of hydrogen element (as compared to only 1 part of oxygen element).

Note. Before we discuss the displacement reactions, it is very essential to know the reactivity series of metals. Because, only if we know the positions of various metals in the reactivity series that we can make out why displacement reactions take place and whether a particular displacement reaction will occur or not. Please note that the reactivity series of metals is also known as activity series of metals. Reactivity series of metals is given on page 119 of this book.

3. DISPLACEMENT REACTIONS

Those reactions in which one element takes the place of another element in a compound, are known as displacement reactions. In general, a more reactive element displaces a less reactive element from its compound. The examples of some important displacement reactions are given below.

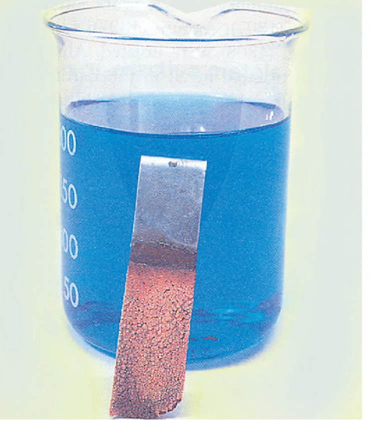


Figure 44. Zinc displaces copper from copper sulphate solution. The displaced copper forms a red-brown deposit on zinc strip.

Example 1. When a strip of zinc metal is placed in copper sulphate solution, then zinc sulphate solution and copper are obtained :

$$\underset{\substack{\text{Copper sulphate} \\ \text{(Blue solution)}}}{CuSO_4\,(aq)} + \underset{\substack{\text{Zinc} \\ \text{(Silvery-white)}}}{Zn\,(s)} \longrightarrow \underset{\substack{\text{Zinc sulphate} \\ \text{(Colourless solution)}}}{ZnSO_4\,(aq)} + \underset{\substack{\text{Copper} \\ \text{(Red-brown)}}}{Cu\,(s)}$$

In this reaction, zinc displaces copper from copper sulphate compound so that copper is set free (or liberated). The blue colour of copper sulphate solution fades due to the formation of zinc sulphate (which is colourless). A red-brown deposit of copper metal is formed on the zinc strip (see Figure 44). Please note that **this displacement reaction takes place because zinc is more reactive than copper.**

Example 2. When a piece of magnesium metal is placed in copper sulphate solution, then magnesium sulphate solution and copper metal are formed :

$$\underset{\substack{\text{Copper sulphate} \\ \text{(Blue solution)}}}{CuSO_4\,(aq)} + \underset{\substack{\text{Magnesium} \\ \text{(Silvery white)}}}{Mg\,(s)} \longrightarrow \underset{\substack{\text{Magnesium sulphate} \\ \text{(Colourless solution)}}}{MgSO_4\,(aq)} + \underset{\substack{\text{Copper} \\ \text{(Red-brown)}}}{Cu\,(s)}$$

In this reaction, magnesium displaces copper from copper sulphate solution. The blue colour of copper sulphate solution fades due to the formation of colourless solution of magnesium sulphate. A red-brown deposit of copper metal is formed on the magnesium piece. Here, **magnesium is able to displace copper from copper sulphate solution because magnesium is more reactive than copper.**

Example 3. When a piece of iron metal (say, an iron nail) is placed in copper sulphate solution, then iron sulphate solution and copper metal are formed :

$$\underset{\substack{\text{Copper sulphate} \\ \text{(Blue solution)}}}{CuSO_4\,(aq)} + \underset{\substack{\text{Iron} \\ \text{(Grey)}}}{Fe\,(s)} \longrightarrow \underset{\substack{\text{Iron sulphate} \\ \text{(Greenish solution)}}}{FeSO_4\,(aq)} + \underset{\substack{\text{Copper} \\ \text{(Red-brown)}}}{Cu\,(s)}$$

In this reaction, iron displaces copper from copper sulphate solution. The deep blue colour of copper sulphate solution fades due to the formation of light green solution of iron sulphate. A red-brown coating (or layer) of copper metal is formed on the surface of iron metal (or iron nail). Please note that **this displacement reaction occurs because iron is more reactive than copper.**

We can perform the displacement reaction between iron and copper sulphate solution as follows :

(*i*) Take about 10 mL of copper sulphate solution in a test-tube. It is deep blue in colour.

(*ii*) Take a big iron nail and clean its surface by rubbing with a sand paper.

(*iii*) Put the cleaned iron nail in the test-tube containing copper sulphate solution [see Figure 45(*a*)]. Allow the iron nail to remain in copper sulphate solution for about half an hour.

(*iv*) After half an hour, take out the iron nail from copper sulphate solution. We will find that the iron nail is covered with a red-brown layer of copper metal [see Figure 45(*b*)].

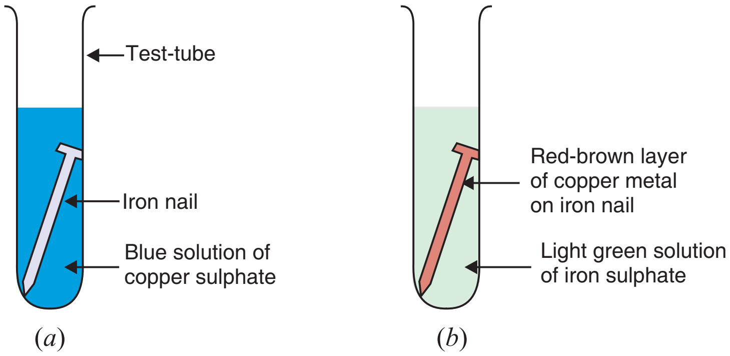


Figure 45. Displacement reaction between iron (nail) and copper sulphate solution.

(*v*) If we look at the test-tube, we find that the original deep blue colour of copper sulphate solution has faded. The solution turns light green due to the formation of iron sulphate (or ferrous sulphate).

Example 4. When a strip of lead metal is placed in a solution of copper chloride, then lead chloride solution and copper metal are formed :

$CuCl_2$ (aq)	+	Pb (s)	$\longrightarrow$	$PbCl_2$ (aq)	+	Cu (s)
Copper chloride		Lead		Lead chloride		Copper
(Green solution)		(Bluish grey)		(Colourless solution)		(Red-brown)

In this case, lead displaces copper from copper chloride solution. The green colour of copper chloride solution fades due to the formation of colourless solution of lead chloride. A red-brown layer of copper metal is deposited on the lead strip. Please note that **lead is able to displace copper from copper chloride solution because lead is more reactive than copper.** Another point to be noted is that copper chloride ($CuCl_2$) used in this reaction is actually copper (II) chloride.

Example 5. When a copper strip is placed in a solution of silver nitrate, then copper nitrate solution and silver metal are formed :

$2AgNO_3$ (aq)	+	Cu (s)	$\longrightarrow$	$Cu(NO_3)_2$ (aq)	+	2Ag (s)
Silver nitrate		Copper		Copper nitrate		Silver
(Colourless solution)		(Red-brown)		(Blue solution)		(Greyish white)

In this case copper displaces silver from silver nitrate compound. **This displacement reaction occurs because copper is more reactive than silver.** A shining greyish white deposit of silver is formed on the copper strip and the solution becomes blue due to the formation of copper nitrate.

Example 6. Iron metal reacts with dilute hydrochloric acid to form iron (II) chloride and hydrogen gas :

Fe (s)	+	2HCl (aq)	$\longrightarrow$	$FeCl_2$ (aq)	+	H_2 (g)
Iron		Hydrochloric acid		Iron (II) chloride		Hydrogen
(As iron filings)				(Ferrous chloride)		

In this reaction, iron displaces hydrogen from hydrochloric acid solution to form hydrogen gas. **This displacement reaction takes place because iron is more reactive than hydrogen.** Please note that iron metal is usually taken in the form of iron filings in this reaction.

Example 7. Magnesium metal reacts with hydrochloric acid to form magnesium chloride and hydrogen gas :

Mg (s)	+	2HCl (aq)	$\longrightarrow$	$MgCl_2$ (aq)	+	H_2 (g)
Magnesium		Hydrochloric acid		Magnesium chloride		Hydrogen

In this displacement reaction, magnesium displaces hydrogen from hydrochloric acid solution. **This displacement reaction occurs because magnesium is more reactive than hydrogen.**

Example 8. Sodium metal reacts with water to form sodium hydroxide solution and hydrogen gas :

$$\underset{\text{Sodium}}{2Na\ (s)} + \underset{\text{Water}}{2H_2O\ (l)} \longrightarrow \underset{\text{Sodium hydroxide}}{2NaOH\ (aq)} + \underset{\text{Hydrogen}}{H_2\ (g)}$$

In this displacement reaction, sodium displaces hydrogen from water. **This displacement reaction takes place because sodium is more reactive than hydrogen.**

Example 9. Chlorine gas reacts with potassium iodide solution to form potassium chloride and iodine :

$$\underset{\text{Chlorine}}{Cl_2\ (g)} + \underset{\text{Potassium iodide}}{2KI\ (aq)} \longrightarrow \underset{\text{Potassium chloride}}{2KCl\ (aq)} + \underset{\text{Iodine}}{I_2\ (s)}$$

In this displacement reaction, chlorine displaces iodine from potassium iodide. **This displacement reaction occurs because chlorine is more reactive than iodine.**

Most of the common displacement reactions occur in aqueous solutions (water solutions). There are, however, some displacement reactions which also occur between solid substances. The displacement reactions of metals with metal oxides are such reactions. Please note that **a more reactive metal displaces a less reactive metal from its oxide.** This will become more clear from the following examples.

Example 10. When copper oxide is heated with magnesium powder, then magnesium oxide and copper are formed :

$$\underset{\text{Copper oxide}}{CuO\ (s)} + \underset{\text{Magnesium}}{Mg\ (s)} \longrightarrow \underset{\text{Magnesium oxide}}{MgO\ (s)} + \underset{\text{Copper}}{Cu\ (s)}$$

This is a displacement reaction. **In this displacement reaction, a more reactive metal, magnesium, is displacing a less reactive metal, copper, from its oxide, copper oxide.**

Example 11. When iron (III) oxide is heated with aluminium powder, then aluminium oxide and iron metal are formed :

$$\underset{\substack{\text{Iron (III) oxide}\\ \text{(Ferric oxide)}}}{Fe_2O_3\ (s)} + \underset{\text{Aluminium}}{2Al\ (s)} \longrightarrow \underset{\text{Aluminium oxide}}{Al_2O_3\ (s)} + \underset{\substack{\text{Iron}\\ \text{(Molten)}}}{2Fe\ (l)}$$

This is a displacement reaction. **In this displacement reaction, a more reactive metal, aluminium, is displacing a less reactive metal, iron, from its oxide, iron (III) oxide.** Please note that so much heat is produced in this reaction that iron is obtained in the molten state (liquid state).

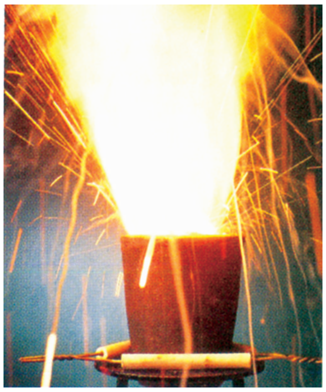


Figure 46. The displacement reaction between iron (III) oxide and powdered aluminium produces so much heat that iron metal obtained is in molten form.

All the above examples of displacement reactions are actually '**single displacement reactions**'. This is because in all these reactions only 'one element' displaces 'another element' from its compound. The single displacement reactions are, however, written as just displacement reactions. So, when we talk of a displacement reaction, it actually means a single displacement reaction. The word 'single' is usually not written with it. We will now describe another type of displacement reactions called '**double displacement reactions**'.

4. DOUBLE DISPLACEMENT REACTIONS

Those reactions in which two compounds react by an exchange of ions to form two new compounds are called double displacement reactions. A double displacement reaction usually occurs in solution and one of the products, being insoluble, precipitates out (separates as a solid). Some of the examples of double displacement reactions are given below :

Example 1. When silver nitrate solution is added to sodium chloride solution, then a white precipitate of silver chloride is formed alongwith sodium nitrate solution :

$$\underset{\text{Silver nitrate}}{AgNO_3\ (aq)} + \underset{\text{Sodium chloride}}{NaCl\ (aq)} \longrightarrow \underset{\substack{\text{Silver chloride}\\ \text{(White ppt.)}}}{AgCl\ (s)} + \underset{\text{Sodium nitrate}}{NaNO_3\ (aq)}$$

In this double displacement reaction, two compounds, silver nitrate and sodium chloride, react to form two new compounds, silver chloride and sodium nitrate. An exchange of ions takes place in this reaction. For example, the silver ions (Ag^+) of silver nitrate react with chloride ions (Cl^-) of sodium chloride to form a new compound, silver chloride (Ag^+Cl^- or AgCl). Similarly, the sodium ions (Na^+) of sodium chloride react with the nitrate ions (NO_3^-) of silver nitrate to form another new compound, sodium nitrate ($Na^+NO_3^-$ or $NaNO_3$). Please note that in the above double displacement reaction, silver chloride is formed as an insoluble white solid called a 'white precipitate'.

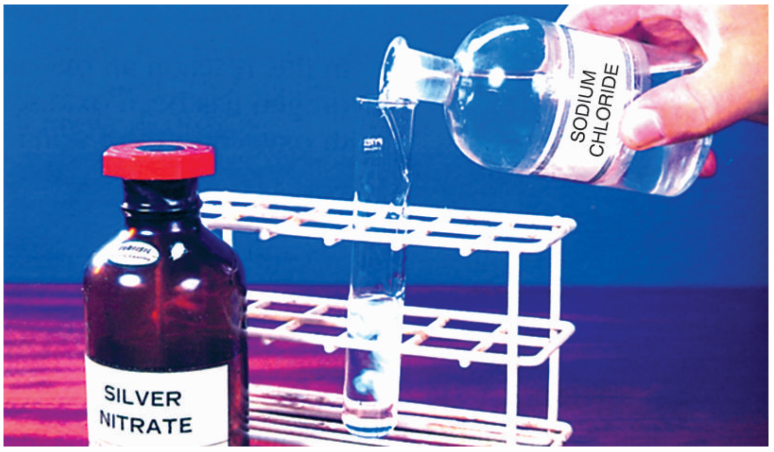


Figure 47. On mixing sodium chloride solution with silver nitrate solution, a white precipitate of silver chloride is obtained alongwith sodium nitrate solution.

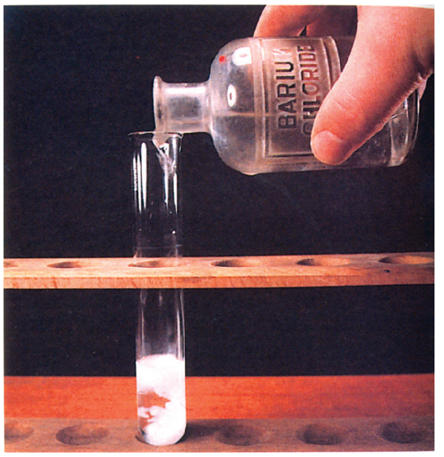


Figure 48. On mixing barium chloride solution with sodium sulphate solution, a white precipitate of barium sulphate is formed alongwith sodium chloride solution.

Example 2. When barium chloride solution is added to sodium sulphate solution, then a white precipitate of barium sulphate is formed alongwith sodium chloride solution :

$$\underset{\text{Barium chloride}}{BaCl_2\,(aq)} + \underset{\text{Sodium sulphate}}{Na_2SO_4\,(aq)} \longrightarrow \underset{\substack{\text{Barium sulphate}\\\text{(White ppt.)}}}{BaSO_4\,(s)} + \underset{\text{Sodium chloride}}{2NaCl\,(aq)}$$

In this displacement reaction, two compounds, barium chloride and sodium sulphate, react to form two new compounds, barium sulphate and sodium chloride. An exchange of ions takes place in this reaction. For example, the barium ions (Ba^{2+}) of barium chloride react with sulphate ions (SO_4^{2-}) of sodium sulphate to form barium sulphate ($Ba^{2+}SO_4^{2-}$ or $BaSO_4$). In this reaction, barium sulphate is formed as a white, insoluble solid (called precipitate) which separates out suddenly from the solution. **Any reaction in which an insoluble solid (called precipitate) is formed that separates from the solution is called a precipitation reaction.** The reaction between barium chloride solution and sodium sulphate solution to form barium sulphate precipitate (alongwith sodium chloride solution) is an example of a precipitation reaction. We can perform this precipitation

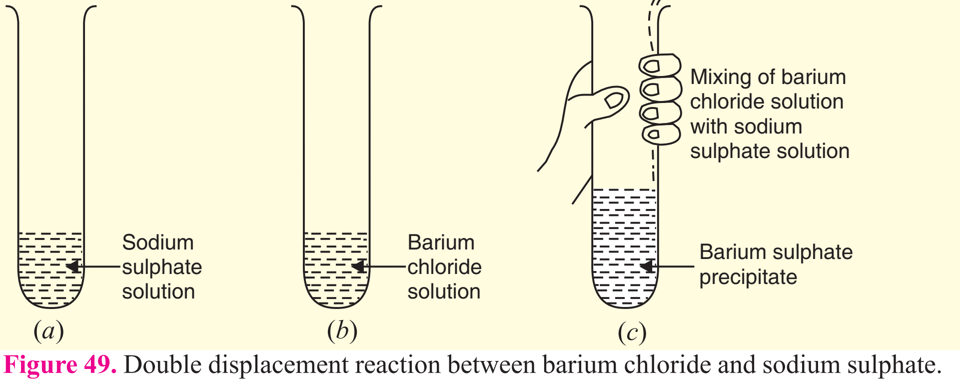


Figure 49. Double displacement reaction between barium chloride and sodium sulphate.

reaction as follows :

(*i*) Take about 3 mL of sodium sulphate solution in a test-tube [see Figure 49(*a*)].
(*ii*) In another test-tube, take 3 mL of barium chloride solution [see Figure 49(*b*)].
(*iii*) Add barium chloride solution to sodium sulphate solution [see Figure 49(*c*)].
(*iv*) A white precipitate of barium sulphate is formed at once.

Example 3. If barium chloride solution is added to copper sulphate solution, then a white precipitate of barium sulphate is produced alongwith copper chloride solution :

$$BaCl_2\,(aq) + CuSO_4\,(aq) \longrightarrow BaSO_4\,(s) + CuCl_2\,(aq)$$

Barium chloride + Copper sulphate ⟶ Barium sulphate (White ppt.) + Copper chloride

In this double displacement reaction, two compounds, barium chloride and copper sulphate, react by an exchange of their ions to form two new compounds, barium sulphate and copper chloride.

Example 4. When hydrogen sulphide gas is passed through copper sulphate solution, then a black precipitate of copper sulphide is formed alongwith sulphuric acid solution :

$$CuSO_4\,(aq) + H_2S\,(g) \longrightarrow CuS\,(s) + H_2SO_4\,(aq)$$

Copper sulphate + Hydrogen sulphide ⟶ Copper sulphide (Black ppt.) + Sulphuric acid

In this double displacement reaction, two compounds, copper sulphate and hydrogen sulphide, react by an exchange of ions to form two new compounds, copper sulphide and sulphuric acid.

Example 5. When ammonium hydroxide solution is added to aluminium chloride solution, then a white precipitate of aluminium hydroxide is formed alongwith ammonium chloride solution :

$$AlCl_3\,(aq) + 3NH_4OH\,(aq) \longrightarrow Al(OH)_3\,(s) + 3NH_4Cl\,(aq)$$

Aluminium chloride + Ammonium hydroxide ⟶ Aluminium hydroxide (White ppt.) + Ammonium chloride

In this double displacement reaction, two compounds, aluminium chloride and ammonium hydroxide, react by an exchange of their ions to form two new compounds, aluminium hydroxide and ammonium chloride.

Example 6. When potassium iodide solution is added to lead nitrate solution, then a yellow precipitate of lead iodide is produced alongwith potassium nitrate solution :

$$Pb(NO_3)_2\,(aq) + 2KI\,(aq) \longrightarrow PbI_2\,(s) + 2KNO_3\,(aq)$$

Lead nitrate + Potassium iodide ⟶ Lead iodide (Yellow ppt.) + Potassium nitrate

This is also a double displacement reaction. In this double displacement reaction, two compounds, lead nitrate and potassium iodide, react by an exchange of ions to form two new compounds, lead iodide and potassium nitrate. Please note that lead nitrate, $Pb(NO_3)_2$, is also written as lead (II) nitrate.

Example 7. The reactions between acids and bases to form salts and water are also double displacement reactions. For example, sodium hydroxide and hydrochloric acid react to form sodium chloride and water :

$$NaOH\,(aq) + HCl\,(aq) \longrightarrow NaCl\,(aq) + H_2O\,(l)$$

Sodium hydroxide + Hydrochloric acid ⟶ Sodium chloride + Water

In this double displacement reaction, two compounds, sodium hydroxide and hydrochloric acid, react by an exchange of ions to form two new compounds, sodium chloride and water. Please note that no precipitate is formed in this double displacement reaction (This is because sodium chloride is soluble in water). Let us solve some problems now.

Sample Problem 1. What happens when dilute hydrochloric acid is added to iron filings ? Tick the correct answer :

(*a*) Hydrogen gas and iron chloride are produced.

(*b*) Chlorine gas and iron hydroxide are produced.

(*c*) No reaction takes place.

(*d*) Iron salt and water are produced. **(NCERT Book Question)**

Solution. (*a*) Hydrogen gas and iron chloride are produced.

Sample Problem 2. $Fe_2O_3 + 2Al \longrightarrow Al_2O_3 + 2Fe$

The above reaction is an example of :

(*a*) combination reaction
(*b*) double displacement reaction
(*c*) decomposition reaction
(*d*) displacement reaction

Choose the correct answer. **(NCERT Book Question)**

Solution. The correct answer is : (*d*) displacement reaction.

Sample Problem 3. Write the balanced chemical equations for the following and identify the type of reaction in each case :

(*a*) Barium chloride (aq) + Potassium sulphate (aq) ⟶ Barium sulphate (s) + Potassium chloride (aq)

(*b*) Zinc carbonate (s) ⟶ Zinc oxide (s) + Carbon dioxide (g)

(*c*) Hydrogen (g) + Chlorine (g) ⟶ Hydrogen chloride (g)

(*d*) Magnesium (s) + Hydrochloric acid (aq) ⟶ Magnesium chloride (aq) + Hydrogen (g)

(NCERT Book Question)

Solution. (*a*) $BaCl_2\ (aq) + K_2SO_4\ (aq) \longrightarrow BaSO_4\ (s) + 2KCl\ (aq)$

This is a double displacement reaction.

(*b*) $ZnCO_3\ (s) \longrightarrow ZnO\ (s) + CO_2\ (g)$

This is a decomposition reaction.

(*c*) $H_2\ (g) + Cl_2\ (g) \longrightarrow 2HCl\ (g)$

This is a combination reaction.

(*d*) $Mg\ (s) + 2HCl\ (aq) \longrightarrow MgCl_2\ (aq) + H_2\ (g)$

This is a displacement reaction.

Sample Problem 4. Below are given two chemical reactions :

(*i*) $2KBr\ (aq) + Cl_2\ (aq) \longrightarrow 2KCl\ (aq) + Br_2\ (aq)$

(*ii*) $Fe\ (s) + S\ (s) \longrightarrow FeS\ (s)$

Which is combination reaction and which is displacement reaction ?

Solution. (*i*) In the first reaction, potassium bromide solution reacts with chlorine solution to form potassium chloride solution and bromine. So, in this reaction, chlorine is displacing bromine from potassium bromide to form potassium chloride and bromine is set free. Thus, it is a displacement reaction.

(*ii*) In the second reaction, iron combines with sulphur to form iron (II) sulphide. So, it is a combination reaction.

5. OXIDATION AND REDUCTION REACTIONS

The earlier concept of oxidation and reduction is based on the addition or removal of oxygen or hydrogen elements. So, in terms of oxygen or hydrogen, oxidation and reduction reactions can be defined as follows :

Oxidation : (*i*) The addition of oxygen to a substance is called oxidation.

(*ii*) The removal of hydrogen from a substance is also called oxidation.

Reduction : (*i*) The addition of hydrogen to a substance is called reduction.

(*ii*) The removal of oxygen from a substance is also called reduction.

(*a*) When phosphorus burns in air, then oxygen is added to it to form phosphorus pentoxide (P_2O_5), so phosphorus gets oxidised

(*b*) The top spoon in this picture contains edible oil (which is a liquid). When hydrogen is added to oil, it gets converted into solid fat (shown in the bottom spoon). So, oil gets reduced

(*c*) The solid fat (like margarine) is used like butter on slices of bread

Figure 50. Some examples of oxidation and reduction reactions.

It is obvious from the above definitions that **the process of reduction is just the opposite of oxidation.** Moreover, **oxidation and reduction occur together.** We will now define the oxidising agents and reducing agents.

Oxidising agent :(*i*) The substance which gives oxygen for oxidation is called an oxidising agent.

(*ii*) The substance which removes hydrogen is also called an oxidising agent.

Reducing agent :(*i*) The substance which gives hydrogen for reduction is called a reducing agent.

(*ii*) The substance which removes oxygen is also called a reducing agent.

The oxidation and reduction reactions are also called redox reactions (In the name '*redox*', the term '*red*' stands for '*reduction*' and '*ox*' stands for *oxidation*). We will now give some examples of oxidation and reduction reactions.

Example 1. When copper oxide is heated with hydrogen, then copper metal and water are formed :

$$\underset{\text{Copper oxide}}{CuO} + \underset{\text{Hydrogen}}{H_2} \xrightarrow{\text{Heat}} \underset{\text{Copper}}{Cu} + \underset{\text{Water}}{H_2O}$$

(*i*) In this reaction, CuO is changing into Cu. That is, oxygen is being removed from copper oxide. Now, by definition, removal of oxygen from a substance is called reduction, so we can say that **copper oxide is being reduced to copper**.

(*ii*) In this reaction H_2 is changing into H_2O. That is, oxygen is being added to hydrogen. Now, by definition, addition of oxygen to a substance is called oxidation, so we can say that **hydrogen is being oxidised to water**.

We find that hydrogen is being oxidised to water and at the same time copper oxide is being reduced to copper. This shows that oxidation and reduction occur together. The oxidation-reduction reaction between copper oxide and hydrogen can be shown more clearly as follows :

Removal of oxygen : Reduction

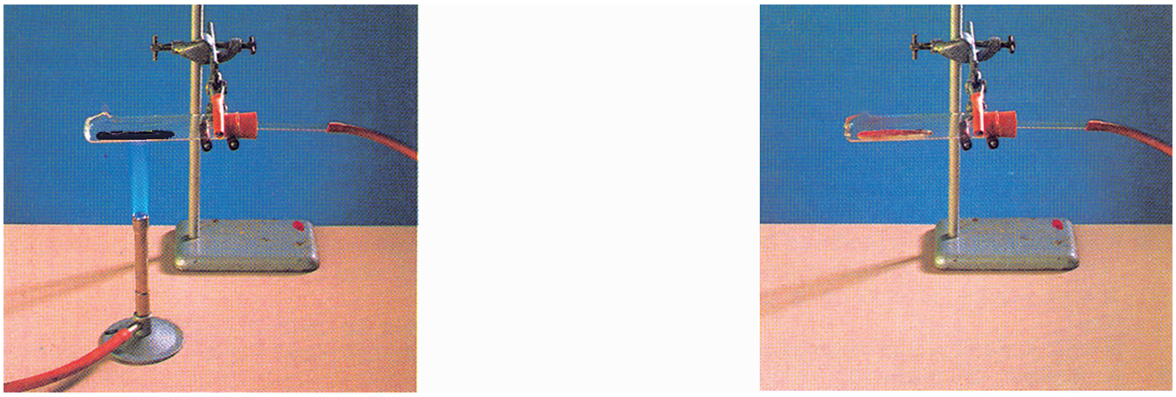
$$CuO + H_2 \xrightarrow{\text{Heat}} Cu + H_2O$$

Addition of oxygen : Oxidation

In the above reaction, copper oxide (CuO) is giving the oxygen required for the oxidation of hydrogen, therefore, **copper oxide is the oxidising agent.** Hydrogen is responsible for removing oxygen from copper oxide, therefore, **hydrogen is the reducing agent here.** This gives us the following conclusions about the above oxidation-reduction reaction :

(*i*) Substance oxidised : H_2
(*ii*) Substance reduced : CuO
(*iii*) Oxidising agent : CuO
(*iv*) Reducing agent : H_2

Please note that the substance which gets oxidised (H_2) is the reducing agent. On the other hand, the substance which gets reduced (CuO) is the oxidising agent. Another point to be noted is that the reaction between copper oxide and hydrogen to form copper and water is an **oxidation-reduction reaction** which is also a **displacement reaction**.



(*a*) Copper oxide (black solid in the above boiling tube) is being heated in a stream of hydrogen gas (coming in through the rubber pipe)

(*b*) After some time, copper metal (red-brown solid in the boiling tube) and water are formed.

Figure 51. Experiment to carry out the redox reaction between copper oxide and hydrogen.

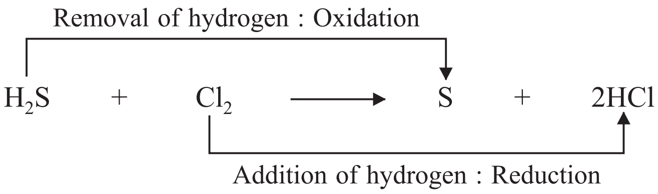
Example 2. When hydrogen sulphide reacts with chlorine, then sulphur and hydrogen chloride are formed :

$$\underset{\text{Hydrogen sulphide}}{H_2S} + \underset{\text{Chlorine}}{Cl_2} \longrightarrow \underset{\text{Sulphur}}{S} + \underset{\text{Hydrogen chloride}}{2HCl}$$

(*i*) In this reaction, H_2S is changing into S. That is, hydrogen is being removed from hydrogen sulphide. Now, by definition, the removal of hydrogen from a compound is called oxidation, so we can say that **hydrogen sulphide is being oxidised to sulphur.**

(*ii*) In this reaction, Cl_2 is changing into HCl. That is, hydrogen is being added to chlorine. Now, by definition, the addition of hydrogen to a substance is called reduction, so we can say that **chlorine is being reduced to hydrogen chloride.**

The oxidation-reduction reaction between hydrogen sulphide and chlorine can be shown more clearly as follows :



In the above reaction, chlorine is removing the hydrogen from hydrogen sulphide, therefore, **chlorine is the oxidising agent.** On the other hand, hydrogen sulphide is supplying hydrogen to chlorine for reduction, so **hydrogen sulphide is the reducing agent.** This gives us the following conclusions about the above oxidation-reduction reaction :

(*i*) Substance oxidised : H_2S
(*ii*) Substance reduced : Cl_2
(*iii*) Oxidising agent : Cl_2
(*iv*) Reducing agent : H_2S

Once again please note that the substance which gets oxidised (H_2S) acts as the reducing agent whereas the substance which gets reduced (Cl_2) acts as the oxidising agent. Thus, a very important conclusion to be remembered about the oxidation and reduction reactions is that :

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

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

Example 3. When zinc oxide is heated with carbon, then zinc metal and carbon monoxide are formed :

$$\underset{\text{Zinc oxide}}{ZnO} + \underset{\text{Carbon}}{C} \xrightarrow{\text{Heat}} \underset{\text{Zinc}}{Zn} + \underset{\text{Carbon monoxide}}{CO}$$

In this reaction, zinc oxide (ZnO) is losing oxygen, so it is being reduced to zinc (Zn). On the other hand, carbon (C) is gaining oxygen, so it is being oxidised to carbon monoxide (CO). In this reaction, zinc oxide is the oxidising agent whereas carbon is the reducing agent. This reaction is used in the production of zinc metal in industry. Carbon is used in the form of coke for the extraction of zinc metal.

Figure 52. Zinc oxide.

Figure 53. Carbon (in the form of coke).

Figure 54. Zinc metal formed by the reduction of zinc oxide with carbon (or coke).

Example 4. When manganese dioxide reacts with hydrochloric acid, then manganese dichloride, chlorine and water are formed :

$$\underset{\text{Manganese dioxide}}{MnO_2} + \underset{\text{Hydrochloric acid}}{4HCl} \longrightarrow \underset{\text{Manganese dichloride}}{MnCl_2} + \underset{\text{Chlorine}}{Cl_2} + \underset{\text{Water}}{2H_2O}$$

In this reaction, MnO_2 is losing oxygen to form $MnCl_2$, so manganese dioxide (MnO_2) is being reduced to manganese dichloride ($MnCl_2$). On the other hand, HCl is losing hydrogen to form Cl_2, so hydrochloric acid (HCl) is being oxidised to chlorine (Cl_2). In this reaction, manganese dioxide (MnO_2) is the oxidising agent whereas hydrochloric acid (HCl) is the reducing agent.

So far we have discussed oxidation and reduction in terms of oxygen and hydrogen. **There is another concept of oxidation and reduction in terms of metals and non-metals.** This is as follows :

(*i*) The addition of non-metallic element (or removal of metallic element) is called oxidation.

(*ii*) The addition of metallic element (or removal of non-metallic element) is called reduction.

This concept of oxidation and reduction will help us in understanding the following oxidation-reduction reaction. Please note that copper oxide, CuO, is also known as copper (II) oxide because the valency of copper in it is II (two).

Example 5. When copper is heated in air, it reacts with the oxygen of air to form a black compound copper oxide :

$$\underset{\substack{\text{Copper} \\ \text{(Red-brown)}}}{2Cu} + \underset{\substack{\text{Oxygen} \\ \text{(From air)}}}{O_2} \xrightarrow{\text{Heat}} \underset{\substack{\text{Copper oxide} \\ \text{(Black)}}}{2CuO}$$

In this reaction, Cu is changing into CuO. This is the addition of oxygen. But addition of oxygen is called oxidation, so copper (Cu) is oxidised to copper oxide (CuO). Now, O_2 is changing into CuO. This is the addition of copper (Cu) which is a metal. But addition of metal is called reduction, so in this reaction, oxygen (O_2) is reduced to copper oxide (CuO). Here oxygen is the oxidising agent whereas copper is the reducing agent. We can carry out this reaction as follows :

(*i*) Take about 1 gram of copper powder in a china dish. It is red-brown in colour.
(*ii*) Heat the china dish strongly over a burner (see Figure 55).
(*iii*) A black substance is formed. This black substance is copper oxide.

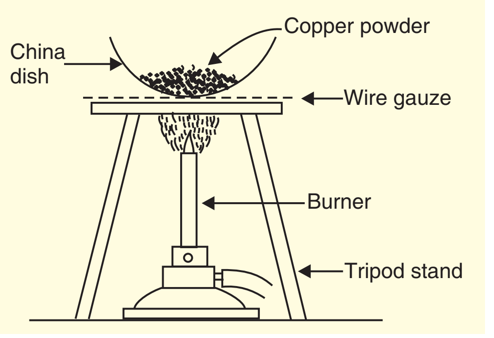


Figure 55. Oxidation of copper to copper oxide.

We have just studied that when copper metal is heated in air, it gets oxidised to form copper oxide. This reaction can be reversed by passing hydrogen gas over heated copper oxide to get back copper metal. Thus, if hydrogen gas is passed over heated copper oxide, then the black copper oxide is reduced and red-brown copper metal is obtained :

$$\underset{\substack{\text{Copper oxide} \\ \text{(Black)}}}{CuO} + \underset{\text{Hydrogen}}{H_2} \xrightarrow{\text{Heat}} \underset{\substack{\text{Copper} \\ \text{(Red-brown)}}}{Cu} + \underset{\text{Water}}{H_2O}$$

In this reaction, copper oxide is reduced to copper metal whereas hydrogen is oxidised to water.

The oxidation of magnesium is similar to the oxidation of copper. When a magnesium ribbon burns in air, it combines with the oxygen of air to form magnesium oxide. This is a combination reaction as well as an oxidation-reduction reaction. In this reaction, magnesium (Mg) is oxidised to magnesium oxide (MgO) whereas oxygen (O_2) is reduced to magnesium oxide (MgO). Oxygen is the oxidising agent whereas magnesium is the reducing agent. Let us solve some problems now.

Sample Problem 1. Name the substance oxidised and substance reduced in the following reaction :

$$\underset{\text{Sulphur dioxide}}{SO_2} + \underset{\text{Hydrogen sulphide}}{2H_2S} \longrightarrow \underset{\text{Water}}{2H_2O} + \underset{\text{Sulphur}}{3S}$$

Solution. (*i*) Here, SO_2 is changing into S. This is the removal of oxygen from SO_2. By definition, the removal of oxygen is called reduction. Thus, SO_2 is being reduced to S. So, the substance being reduced is sulphur dioxide, SO_2.

(*ii*) H_2S is changing into S. This is the removal of hydrogen from H_2S. By definition, the removal of hydrogen is known as oxidation. Thus, H_2S is being oxidised to S. So, the substance being oxidised is hydrogen sulphide, H_2S.

Sample Problem 2. Select the oxidising agent and the reducing agent from the following reaction :

$$\underset{\text{Hydrogen sulphide}}{H_2S} + \underset{\text{Iodine}}{I_2} \longrightarrow \underset{\text{Hydrogen iodide}}{2HI} + \underset{\text{Sulphur}}{S}$$

Solution. (*i*) H_2S is changing into S. This is the removal of hydrogen from H_2S. By definition, the removal of hydrogen is known as oxidation, therefore, hydrogen sulphide is being oxidised to sulphur. Iodine is removing the hydrogen from H_2S, so iodine is the oxidising agent.

(*ii*) I_2 is changing into HI. This is the addition of hydrogen to iodine. By definition, addition of hydrogen is known as reduction, therefore, iodine is being reduced to hydrogen iodide. Hydrogen sulphide is supplying the hydrogen required for reduction, so hydrogen sulphide is the reducing agent.

Sample Problem 3. Identify the substance that is oxidised and the substance that is reduced in the following reaction :

$$4Na\ (s) + O_2\ (g) \longrightarrow 2Na_2O\ (s)$$ **(NCERT Book Question)**

Solution. (*i*) Here sodium (Na) is changing into sodium oxide (Na_2O). This is the addition of oxygen to sodium. Now, addition of oxygen is called oxidation. So, the substance oxidised is sodium (Na).

(*ii*) In this reaction, oxygen (O_2) is changing into sodium oxide (Na_2O). This is the addition of a metal (sodium, Na) to oxygen. Now, by definition, the addition of a metal is called reduction. So, the substance reduced is oxygen (O_2).

Sample Problem 4. Which of the statements about the reaction below are incorrect ?

$$2PbO\ (s) + C\ (s) \longrightarrow 2Pb\ (s) + CO_2\ (g)$$

(*a*) Lead is getting reduced.
(*b*) Carbon dioxide is getting oxidised.
(*c*) Carbon is getting oxidised.
(*d*) Lead oxide is getting reduced. **(NCERT Book Question)**

Solution. The incorrect statements are :

(*a*) Lead is getting reduced.

(*b*) Carbon dioxide is getting oxidised.

Sample Probem 5. A shiny brown coloured element *X* on heating in air becomes black in colour. Name the element *X* and the black coloured compound formed. **(NCERT Book Question)**

Solution. (*i*) The shiny brown coloured element *X* is copper metal (Cu).

(*ii*) When copper metal is heated in air, it forms a black coloured compound copper oxide. So, the black coloured compound is copper oxide or copper (II) oxide, CuO.

EFFECTS OF OXIDATION REACTIONS IN EVERYDAY LIFE

Oxidation has damaging effect on metals as well as on food. 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 daily life. These are :**

1. **Corrosion of metals ,** and
2. **Rancidity of food.**

We will now describe these two effects caused by the process of oxidation in somewhat detail. Please note that the oxidation involved in the corrosion of metals as well as rancidity of food is caused naturally by the oxygen present in air. The oxidation caused by the oxygen of air is sometimes also known as 'aerial oxidation'.

Corrosion

Corrosion is the process in which metals are eaten up gradually by the action of air, moisture or a chemical (such as an acid) on their surface. Corrosion is caused mainly by the oxidation of metals by the oxygen of air. **Rusting of iron metal is the most common form of corrosion.** When an iron object is left in damp air for a considerable time, it gets covered with a red-brown flaky substance called 'rust'. This is called rusting of iron.

During the corrosion of iron (or rusting of iron), iron metal is oxidised by the oxygen of air in the presence of water (moisture) to form hydrated iron (III) oxide called rust :

$$\underset{\text{Iron}}{4Fe} + \underset{\text{Oxygen}}{3O_2} + \underset{\text{Water}}{2xH_2O} \longrightarrow \underset{\substack{\text{Hydrated iron (III) oxide} \\ \text{(Rust)}}}{2Fe_2O_3.xH_2O}$$

Please note that the number of water molecules (x) in the rust varies, it is not fixed. **The rusting of iron is a redox reaction**. Rusting involves unwanted oxidation of iron metal which occurs in nature on its own.

Figure 56. This picture shows the eating up of an iron object due to corrosion (or rusting) caused mainly by the oxidation of iron by oxygen of air.

Rust is a soft and porous substance which gradually falls off from the surface of an iron object, and then the iron below starts rusting. Thus, rusting of iron (or corrosion of iron) is a continuous process which, if not prevented in time, eats up the whole iron object (see Figure 56). **Corrosion weakens the iron and steel objects and structures such as railings, car bodies, bridges and ships, etc., and cuts short their life.** A lot of money has to be spent every year to prevent the corrosion of iron and steel objects, and to replace the damaged iron and steel structures. We will study the corrosion of metals and the methods of its prevention in detail in the discussion on metals in Chapter 3.

Rancidity

Oxidation also has damaging effect on foods containing fats and oils. When the food materials prepared in fats and oils are kept for a long time, they start giving unpleasant smell and taste. The fat and oil containing food materials which give unpleasant smell and taste are said to have become rancid (sour or stale). This happens as follows.

When the fats and oils present in food materials get oxidised by the oxygen (of air), their oxidation products have unpleasant smell and taste. Due to this the smell and taste of food materials containing fats and oils change and become very unpleasant (or obnoxious). **The condition produced by aerial oxidation of fats and oils in foods marked by unpleasant smell and taste is called rancidity.** Rancidity spoils the food materials prepared in fats and oils which have been kept for a considerable time and makes them unfit for eating. The characteristics of a rancid food are that it gives out unpleasant smell and also has an unpleasant taste. Rancidity is called '*vikritgandhita*' in Hindi.

Figure 57. These potato chips are made in oil. On keeping exposed for a long time, these potato chips start giving unpleasant smell and taste. They turn rancid.

The development of rancidity of food can be prevented or retarded (slowed down) in the following ways :

1. Rancidity can be prevented by adding anti-oxidants to foods containing fats and oils. Anti-oxidant is a substance (or chemical) which prevents oxidation. Anti-oxidants are actually reducing agents. When anti-oxidants are added to foods, then the fats and oils present in them do not get oxidised easily and hence do not turn rancid. So the foods remain good to eat for a much longer time. The two common anti-oxidants used in foods to prevent the development of rancidity are BHA (Butylated Hydroxy-Anisole) and BHT (Butylated Hydroxy-Toluene).

2. Rancidity can be prevented by packaging fat and oil containing foods in nitrogen gas. When the packed food is surrounded by an unreactive gas nitrogen, there is no oxygen (of air) to cause its oxidation and make it rancid. The manufacturers of potato chips (and other similar food products) fill the plastic bags containing chips with nitrogen gas to prevent the chips from being oxidised and turn rancid (see Figure 58).

3. Rancidity can be retarded by keeping food in a refrigerator (see Figure 59). The refrigerator has a low temperature inside it. When the food is kept in a refrigerator, the oxidation of fats and oils in it is slowed down due to low temperature. So, the development of rancidity due to oxidation is retarded.

4. Rancidity can be retarded by storing food in air-tight containers. When food is stored in air-tight containers, then there is little exposure to oxygen of air. Due to reduced exposure to oxygen, the oxidation of fats and oils present in food is slowed down and hence the development of rancidity is retarded.

Figure 58. Unreactive gas nitrogen is used in packaging some foods like potato chips to prevent their rancidity.

Figure 59. Rancidity can also be prevented by keeping the food in a refrigerator.

5. Rancidity can be retarded by storing foods away from light. In the absence of light, the oxidation of fats and oils present in food is slowed down and hence the development of rancidity is retarded.

We are now in a position to **answer the following questions :**

Very Short Answer Type Questions

1. What type of reaction is represented by the digestion of food in our body ?
2. Name the various types of chemical reactions.
3. Why does the colour of copper sulphate solution change when an iron nail is kept immersed in it ?
4. Write the balanced chemical equation for the following reaction :

   Zinc + Silver nitrate $\longrightarrow$ Zinc nitrate + Silver
5. Which term is used to indicate the development of unpleasant smell and taste in fat and oil containing foods due to aerial oxidation (when they are kept exposed for a considerable time) ?
6. What is the general name of the chemicals which are added to fat and oil containing foods to prevent the development of rancidity ?
7. State an important use of decomposition reactions.
8. What are anti-oxidants ? Why are they added to fat and oil containing foods ?
9. Explain why, food products containing fats and oils (like potato chips) are packaged in nitrogen.
10. Give one example of a decomposition reaction which is carried out :

    (*a*) with electricity

    (*b*) by applying heat
11. What type of chemical reaction is used to extract metals from their naturally occurring compounds like oxides or chlorides ?
12. Name two anti-oxidants which are usually added to fat and oil containing foods to prevent rancidity.
13. Write one equation each for the decomposition reactions where energy is supplied in the form of (*a*) heat, (*b*) light, and (*c*) electricity.
14. In the refining of silver, the recovery of silver from silver nitrate solution involved displacement by copper metal. Write down the chemical equation of the reaction involved.
15. What type of reactions are represented by the following equations ?

    (*i*) $CaCO_3 \longrightarrow CaO + CO_2$

    (*ii*) $CaO + H_2O \longrightarrow Ca(OH)_2$

    (*iii*) $2FeSO_4 \longrightarrow Fe_2O_3 + SO_2 + SO_3$

    (*iv*) $NH_4Cl \longrightarrow NH_3 + HCl$

    (*v*) $2Ca + O_2 \longrightarrow 2CaO$

16. What type of chemical reactions take place when :

(*a*) a magnesium wire is burnt in air ?

(*b*) lime-stone is heated ?

(*c*) silver bromide is exposed to sunlight ?

(*d*) electricity is passed through water ?

(*e*) ammonia and hydrogen chloride are mixed ?

17. What type of chemical reactions are represented by the following equations ?

(*i*) $A + BC \longrightarrow AC + B$

(*ii*) $A + B \longrightarrow C$

(*iii*) $X \longrightarrow Y + Z$

(*iv*) $PQ + RS \longrightarrow PS + RQ$

(*v*) $A_2O_3 + 2B \longrightarrow B_2O_3 + 2A$

18. Balance the following chemical equations :

(*a*) $FeSO_4 \xrightarrow{\text{Heat}} Fe_2O_3 + SO_2 + SO_3$

(*b*) $Pb(NO_3)_2\,(s) \xrightarrow{\text{Heat}} PbO\,(s) + NO_2\,(g) + O_2\,(g)$

19. Which of the following is a combination and which is a displacement reaction ?

(*a*) $Cl_2 + 2KI \longrightarrow 2KCl + I_2$

(*b*) $2K + Cl_2 \longrightarrow 2KCl$

20. What type of reactions are represented by the following equations ?

(*a*) $CaO + CO_2 \longrightarrow CaCO_3$

(*b*) $2Na + 2H_2O \longrightarrow 2NaOH + H_2$

(*c*) $Mg + CuSO_4 \longrightarrow MgSO_4 + Cu$

(*d*) $NH_4NO_2 \longrightarrow N_2 + 2H_2O$

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

21. In the following reaction between lead sulphide and hydrogen peroxide :

$$PbS\,(s) + 4H_2O_2\,(aq) \longrightarrow PbSO_4\,(s) + 4H_2O\,(l)$$

(*a*) Which substance is reduced ?

(*b*) Which substance is oxidised ?

22. Identify the component oxidised in the following reaction :

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

23. When SO_2 gas is passed through saturated solution of H_2S, the following reaction occurs :

$$SO_2 + 2H_2S \longrightarrow 2H_2O + 3S$$

In this reaction, which substance is oxidised and which one is reduced ?

24. Fill in the following blanks with suitable words :

(*a*) The addition of oxygen to a substance is called .... whereas removal of oxygen is called ....

(*b*) The addition of hydrogen to a substance is called .... whereas removal of hydrogen is called ....

(*c*) Anti-oxidants are often added to fat containing foods to prevent ....due to oxidation.

Short Answer Type Questions

25. What is an oxidation reaction ? Identify in the following reaction (*i*) the substance oxidised, and (*ii*) the substance reduced :

$$ZnO + C \longrightarrow Zn + CO$$

26. (*a*) What is a redox reaction ? Explain with an example.

(*b*) When a magnesium ribbon burns in air with a dazzling flame and forms a white ash, is magnesium oxidised or reduced ? Why ?

(*c*) In the reaction represented by the equation :

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

(*i*) name the substance oxidised. (*ii*) name the oxidising agent.

(*iii*) name the substance reduced. (*iv*) name the reducing agent.

27. (*a*) Define a combination reaction.
(*b*) Give one example of a combination reaction which is also exothermic.
(*c*) Give one example of a combination reaction which is also endothermic.

28. (*a*) Give an example of an oxidation reaction.
(*b*) Is oxidation an exothermic or an endothermic reaction ?
(*c*) Explain, by giving an example, how oxidation and reduction proceed side by side.

29. (*a*) What is the colour of ferrous sulphate crystals ? How does this colour change after heating ?
(*b*) Name the product formed on strongly heating ferrous sulphate crystals. What type of chemical reaction occurs in this change ?

30. What is a decomposition reaction ? Give an example of a decomposition reaction. Describe an activity to illustrate such a reaction by heating.

31. Zinc oxide reacts with carbon, on heating, to form zinc metal and carbon monoxide. Write a balanced chemical equation for this reaction. Name (*i*) oxidising agent, and (*ii*) reducing agent, in this reaction.

32. Give one example of an oxidation-reduction reaction which is also :
(*a*) a combination reaction
(*b*) a displacement reaction

33. (*a*) What is the difference between displacement and double displacement reactions ? Write equations for these reactions.
(*b*) What do you mean by a precipitation reaction ? Explain giving an example.

34. (*a*) Explain the following in terms of gain or loss of oxygen with one example each :
(*i*) oxidation (*ii*) reduction
(*b*) When copper powder is heated strongly in air, it forms copper oxide. Write a balanced chemical equation for this reaction. Name (*i*) substance oxidised, and (*ii*) substance reduced.

35. (*a*) Define the following in terms of gain or loss of hydrogen with one example each :
(*i*) oxidation (*ii*) reduction
(*b*) When a magnesium ribbon is heated, it burns in air to form magnesium oxide. Write a balanced chemical equation for this reaction. Name (*i*) substance oxidised, and (*ii*) substance reduced.

36. What is meant by (*a*) displacement reaction, and (*b*) double displacement reaction ? Explain with the help of one example each.

37. (*a*) Why are decomposition reactions called the opposite of combination reactions ? Explain with equations of these reactions.
(*b*) Express the following facts in the form of a balanced chemical equation :
"When a strip of copper metal is placed in a solution of silver nitrate, metallic silver is precipitated and a solution containing copper nitrate is formed".

38. (*a*) What happens when a piece of iron metal is placed in copper sulphate solution ? Name the type of reaction involved.
(*b*) Write balanced chemical equation with state symbols for the following reaction :
Barium chloride solution reacts with sodium sulphate solution to give insoluble barium sulphate and a solution of sodium chloride.

39. In the reaction represented by the following equation :

$$CuO\ (s) + H_2\ (g) \longrightarrow Cu\ (s) + H_2O\ (l)$$

(*a*) name the substance oxidised (*b*) name the substance reduced
(*c*) name the oxidising agent (*d*) name the reducing agent

40. What happens when silver nitrate solution is added to sodium chloride solution ?
(*a*) Write the equation for the reaction which takes place.
(*b*) Name the type of reaction involved.

41. What happens when silver chloride is exposed to sunlight ? Write a chemical equation for this reaction. Also give one use of such a reaction.

42. What happens when a zinc strip is dipped into a copper sulphate solution ?
(*a*) Write the equation for the reaction that takes place.
(*b*) Name the type of reaction involved.

Long Answer Type Questions

43. (*a*) Explain the term "corrosion" with an example. Write a chemical equation to show the process of corrosion of iron.
(*b*) What special name is given to the corrosion of iron ?
(*c*) What type of chemical reaction is involved in the corrosion of iron ?
(*d*) Name any three objects (or structures) which are gradually damaged by the corrosion of iron and steel.

44. (*a*) Explain the term "rancidity". What damage is caused by rancidity ?
(*b*) What type of chemical reaction is responsible for causing rancidity ?
(*c*) State and explain the various methods for preventing or retarding rancidity of food.

45. (*a*) What happens when an aqueous solution of sodium sulphate reacts with an aqueous solution of barium chloride ?
(*b*) Write the balanced chemical equation for the reaction which takes place.
(*c*) State the physical conditions of reactants in which the reaction will not take place.
(*d*) Name the type of chemical reaction which occurs.
(*e*) Give one example of another reaction which is of the same type as the above reaction.

Multiple Choice Questions (MCQs)

46. The removal of oxygen from a substance is called :
(*a*) oxidation (*b*) corrosion (*c*) reduction (*d*) rancidity

47. In the context of redox reactions, the removal of hydrogen from a substance is known as :
(*a*) oxidation (*b*) dehydration (*c*) reduction (*d*) dehydrogenation

48. The chemical reaction involved in the corrosion of iron metal is that of :
(*a*) oxidation as well as displacement (*b*) reduction as well as combination
(*c*) oxidation as well as combination (*d*) reduction as well as displacement

49. The term used to indicate the development of unpleasant smell and taste in fat and oil containing foods due to aerial oxidation is :
(*a*) acidity (*b*) radioactivity (*c*) rabidity (*d*) rancidity

50. In order to prevent the spoilage of potato chips, they are packed in plastic bags in an atmosphere of :
(*a*) Cl_2 (*b*) H_2 (*c*) N_2 (*d*) O_2

51. A white precipitate can be obtained by adding dilute sulphuric acid to :
(*a*) $CuSO_4$ solution (*b*) NaCl solution (*c*) $BaCl_2$ solution (*d*) Na_2SO_4 solution

52. A white precipitate will be formed if we add common salt solution to :
(*a*) $Ba(NO_3)_2$ solution (*b*) KNO_3 solution (*c*) $AgNO_3$ solution (*d*) $Mg(NO_3)_2$ solution

53. Consider the following equation of the chemical reaction of a metal M :

$$4M + 3O_2 \longrightarrow 2M_2O_3$$

This equation represents :
(*a*) combination reaction as well as reduction reaction
(*b*) decomposition reaction as well as oxidation reaction
(*c*) oxidation reaction as well as displacement reaction
(*d*) combination reaction as well as oxidation reaction

54. The process of respiration is :
(*a*) an oxidation reaction which is endothermic
(*b*) a reduction reaction which is exothermic
(*c*) a combination reaction which is endothermic
(*d*) an oxidation reaction which is exothermic

55. Which of the following can be decomposed by the action of light ?
(*a*) NaCl (*b*) KCl (*c*) AgCl (*d*) CuCl

56. Consider the reaction :

$$KBr\ (aq) + AgNO_3\ (aq) \longrightarrow KNO_3\ (aq) + AgBr\ (s)$$

This is an example of :

(*a*) decomposition reaction
(*b*) combination reaction
(c) double displacement reaction
(*d*) displacement reaction

57. You are given the following chemical equation :

$$Mg (s) + CuO (s) \longrightarrow MgO (s) + Cu (s)$$

This equation represents :

(*a*) decomposition reaction as well as displacement reaction
(*b*) combination reaction as well as double displacement reaction
(*c*) redox reaction as well as displacement reaction
(*d*) double displacement reaction as well as redox reaction

Questions Based on High Order Thinking Skills (HOTS)

58. When a green iron salt is heated strongly, its colour finally changes to brown and odour of burning sulphur is given out.

(*a*) Name the iron salt.
(*b*) Name the type of reaction that takes place during the heating of iron salt.
(*c*) Write a chemical equation for the reaction involved.

59. A colourless lead salt, when heated, produces a yellow residue and brown fumes.

(*a*) Name the lead salt.
(*b*) Name the brown fumes.
(*c*) Write a chemical equation of the reaction involved.

60. When hydrogen burns in oxygen, water is formed and when water is electrolysed, then hydrogen and oxygen are produced. What type of reaction takes place :

(*a*) in the first case ?
(*b*) in the second case ?

61. A strip of metal X is dipped in a blue coloured salt solution YSO_4. After some time, a layer of metal Y from the salt solution is formed on the surface of metal strip X. Metal X is used in galvanisation whereas metal Y is used in making electric wires. Metal X and metal Y together form an alloy Z.

(*a*) What could metal X be ?
(*b*) What could metal Y be ?
(*c*) Name the metal salt YSO_4.
(*d*) What type of chemical reaction takes place when metal X reacts with salt solution YSO_4 ? Write the equation of the chemical reaction involved.
(*e*) Name the alloy Z.

62. When a black metal compound XO is heated with a colourless gas Y_2, then metal X and another compound Y_2O are formed. Metal X is red-brown in colour which does not react with dilute acids at all. Gas Y_2 can be prepared by the action of a dilute acid on any active metal. The compound Y_2O is a liquid at room temperature which can turn anhydrous copper sulphate blue.

(*a*) What do you think is metal X ?
(*b*) What could be gas Y_2 ?
(*c*) What is compound XO ?
(*d*) What is compound Y_2O ?
(*e*) Write the chemical equation of the reaction which takes place on heating XO with Y_2.
(*f*) What type of chemical reaction is illustrated in the above equation ?

63. A metal X forms a water soluble salt XNO_3. When an aqueous solution of XNO_3 is added to common salt solution, then a white precipitate of compound Y is formed alongwith sodium nitrate solution. Metal X is said to be the best conductor of electricity and it does not evolve hydrogen when put in dilute hydrohloric acid.

(*a*) What is metal X ?
(*b*) What is salt XNO_3 ?
(*c*) Name the compound Y.

(*d*) Write the chemical equation of the reaction which takes place on reacting XNO_3 solution and common salt solution giving the physical states of all the reactants and products.

(*e*) What type of chemical reaction is illustrated by the above equation ?

64. Two metals X and Y form the salts XSO_4 and Y_2SO_4, respectively. The solution of salt XSO_4 is blue in colour whereas that of Y_2SO_4 is colourless. When barium chloride solution is added to XSO_4 solution, then a white precipitate Z is formed alongwith a salt which turns the solution green. And when barium chloride solution is added to Y_2SO_4 solution, then the same white precipitate Z is formed alongwith colourless common salt solution.

(*a*) What could the metals X and Y be ?

(*b*) Write the name and formula of salt XSO_4.

(*c*) Write the name and formula of salt Y_2SO_4.

(*d*) What is the name and formula of white precipitate Z ?

(*e*) Write the name and formula of the salt which turns the solution green in the first case.

65. A red-brown metal X forms a salt XSO_4. When hydrogen sulphide gas is passed through an aqueous solution of XSO_4, then a black precipitate of XS is formed alongwith sulphuric acid solution.

(*a*) What could the salt XSO_4 be ?

(*b*) What is the colour of salt XSO_4 ?

(*c*) Name the black precipitate XS.

(*d*) By using the formula of the salt obtained in (*a*) above, write an equation of the reaction which takes place when hydrogen sulphide gas is passed through its aqueous solution.

(*e*) What type of chemical reaction takes place in this case ?

66. When a strip of red-brown metal X is placed in a colourless salt solution YNO_3 then metal Y is set free and a blue coloured salt solution $X(NO_3)_2$ is formed. The liberated metal Y forms a shining white deposit on the strip of metal X.

(*a*) What do you think metal X is ?

(*b*) Name the salt YNO_3.

(*c*) What could be metal Y ?

(*d*) Name the salt $X(NO_3)_2$.

(*e*) What type of reaction takes place between metal X and salt solution YNO_3 ?

67. A metal salt MX when exposed to light splits up to form metal M and a gas X_2. Metal M is used in making ornaments whereas gas X_2 is used in making bleaching powder. The salt MX is itself used in black and white photography.

(*a*) What do you think metal M is ?

(*b*) What could be gas X_2 ?

(*c*) Name the metal salt MX.

(*d*) Name any two salt solutions which on mixing together can produce a precipitate of salt MX.

(*e*) What type of chemical reaction takes place when salt MX is exposed to light ? Write the equation of the reaction.

ANSWERS

1. Decomposition reaction **4.** $Zn + 2AgNO_3 \longrightarrow Zn(NO_3)_2 + 2Ag$ **5.** Rancidity **6.** Anti-oxidants **11.** Decomposition reactions (carried out by electricity) **14.** $2AgNO_3\,(aq) + Cu\,(s) \longrightarrow Cu(NO_3)_2\,(aq) + 2Ag\,(s)$ **15.** (*i*) Decomposition (*ii*) Combination (*iii*) Decomposition (*iv*) Decomposition (*v*) Combination **16.** (*a*) Combination (*b*) Decomposition (*c*) Decomposition (*d*) Decomposition (*e*) Combination **17.** (*i*) Displacement reaction (*ii*) Combination reaction (*iii*) Decomposition reaction (*iv*) Double displacement reaction (*v*) Displacement reaction **18.** (*a*) See page 29 (*b*) See page 29 **19.** (*a*) Displacement reaction (*b*) Combination reaction **20.** (*a*) Combination reaction (*b*) Displacement reaction (*c*) Displacement reaction (*d*) Decomposition reaction (*e*) Double displacement reaction 21. (*a*) H_2O_2 (*b*) PbS **22.** H_2S **23.** Substance oxidised : H_2S ; Substance reduced : SO_2 **24.** (*a*) oxidation ; reduction (*b*) reduction ; oxidation (*c*) rancidity **25.** (*i*) C (*ii*) ZnO **26.** (*b*) Magnesium is oxidised ; Addition of oxygen to magnesium takes place (*c*) (*i*) HCl (*ii*) MnO_2 (*iii*) MnO_2 (*iv*) HCl **31.** $ZnO + C \rightarrow Zn + CO$ (*i*) Zinc oxide (*ii*) Carbon **34.** (*b*) $2Cu + O_2 \longrightarrow 2CuO$ (*i*) Cu (*ii*) O_2 **35.** (*b*) $2Mg + O_2 \longrightarrow 2MgO$ (*i*) Mg (*ii*) O_2 **37.** (*b*) $Cu\,(s) + 2AgNO_3\,(aq) \longrightarrow Cu(NO_3)_2\,(aq) + 2Ag\,(s)$ **39.** (*a*) H_2 (*b*) CuO (*c*) CuO (*d*) H_2 **43** (*b*) Rusting (*c*) Oxidation **44.** (*b*) Oxidation

45. (*c*) Solid sodium sulphate and Solid barium chloride (*d*) Double displacement reaction (*e*) Double displacement reaction between silver nitrate solution and sodium chloride solution to form a white precipitate of silver chloride and sodium nitrate solution **46.** (*c*) **47.** (*a*) **48.** (*c*) **49.** (*d*) **50.** (*c*) **51.** (*c*) **52.** (*c*) **53.** (*d*) **54.** (*d*) **55.** (*c*) **56.** (*c*) **57.** (c) **58.** (*a*) Ferrous sulphate (*b*) Decomposition reaction (*c*) See page 29 **59.** (*a*) Lead nitrate (*b*) Nitrogen dioxide (*c*) See page 29 **60.** (*a*) Combination reaction (*b*) Decomposition reaction **61.** (*a*) Zinc (Zn) (*b*) Copper (Cu) (*c*) Copper sulphate ($CuSO_4$) (*d*) Displacement rection ; $CuSO_4 (aq) + Zn (s) \longrightarrow ZnSO_4 (aq) + Cu (s)$ (*e*) Brass **62.** (*a*) Copper (Cu) (*b*) Hydrogen (H_2) (*c*) Copper oxide (CuO) (*d*) Water (H_2O) (*e*) $CuO + H_2 \longrightarrow Cu + H_2O$ (*f*) Displacement reaction (which is also a redox reaction) **63.** (*a*) Silver (Ag) (*b*) Silver nitrate ($AgNO_3$) (*c*) Silver chloride (AgCl) (*d*) $AgNO_3 (aq) + NaCl (aq) \longrightarrow AgCl (s) + NaNO_3 (aq)$ (*e*) Double displacement reaction **64.** (*a*) Metal X : Copper ; Metal Y : Sodium (*b*) Copper sulphate, $CuSO_4$ (*c*) Sodium sulphate, Na_2SO_4 (*d*) Barium sulphate, $BaSO_4$ (*e*) Copper chloride, $CuCl_2$ **65.** (*a*) Copper sulphate (*b*) Blue colour (*c*) Copper sulphide (*d*) $CuSO_4 (aq) + H_2S (g) \longrightarrow CuS (s) + H_2SO_4 (aq)$ (*e*) Double displacement reaction **66.** (*a*) Copper (*b*) Silver nitrate (*c*) Silver (*d*) Copper nitrate (*e*) Displacement reaction **67.** (*a*) Silver (*b*) Chlorine (c) Silver chloride (*d*) Silver nitrate and Sodium chloride (*e*) Decomposition reaction; $2AgCl (s) \xrightarrow{\text{Light}} 2Ag (s) + Cl_2 (g)$