4

CHEMICAL REACTION AND EQUATIONS

Everyday we observe different types of changes in our surroundings. Some of these changes are very simple and are of *temporary nature*. Some others are really complex and of *permanent nature*. When ice kept in a tumbler is exposed to the atmosphere, it melts and is converted into water. When the tumbler containing this water is kept in a freezer it is converted again into ice. Thus, this is a temporary change and the substance comes to its original form. Such changes are ***physical changes***. However, milk once converted into curd can not be converted into milk again. Such changes are ***chemical changes***. These changes are of permanent nature. Both physical and chemical changes are integral part of our daily life. We can present these changes in the form of an equation.

In this lesson we shall discuss how to write and balance chemical equations.We shall also describe different types of chemical reactions.

OBJECTIVES

After completing this lesson, you will be able to:

- *write and balance simple chemical equations;*
- *describe the significance of a balanced chemical equation;*
- *explore the relationship between mole, mass and volume of various reactants and products;*
- *classify chemical reactions as combination, decomposition, displacement and double displacement reactions and*
- *define oxidation and reduction processes (redox reactions) and correlate these with corrosion and rancidity and other aspects of daily life.*

4.1 CHEMICAL EQUATIONS

You must have observed many chemical changes in nature, in your surroundings and in your daily lives. Let us perform a few activities to observe changes.

ACTIVITY 4.1

A. Take a 2 cm long magnesium ribbon. Clean it with a piece of sand paper. Hold it firmly with a pair of tongs. Heat it over a spirit lamp or a burner until it burns. Keep the ribbon as far as possible from your eyes. What do you observe? The magnesium ribbon burns with a dazzling light and liberates a lot of heat. It is soon converted into a white powdery substance.

Fig. 4.1: *Burning of magnesium ribbon*

B. Take a few zinc granules in a conical flask or in a test tube. Add dilute sulphuric or hydrochloric acid to it. What do you observe? There is evolution of gas from the test tube. If you touch the bottom of the test tube, you will find that it has become quite warm.

You can perform many more such activities in the laboratory or in the activity room.

4.1.1 How does one describe these Chemical Changes?

The two reactions mentioned above can be written in words as follows :

$$\underset{reactants}{\text{Magnesium} + \text{Oxygen}} \longrightarrow \underset{product}{\text{Magnesium oxide}} \quad ...(1)$$

$$\text{Zinc} + \text{dil Sulphuric acid} \longrightarrow \text{Zinc sulphate} + \text{Hydrogen} \quad ...(2)$$

A substance which undergoes a chemical change is called the ***reactant*** and the substance which is formed as a result of a chemical change is called the ***product***. In reaction (1) magnesium and oxygen undergo chemical change and they are the reactants. In reaction (2) zinc and dilute sulphuric acid are the reactants. Similarly in reaction (1) magnesium oxide is a new substance formed. It is the product. Can you now say what is the product in reaction (2)? Yes, it is zinc sulphate and hydrogen. In chemical reaction, the reactant (s) is (are) written on the left hand side and the product(s) is (are) written on the right hand side. The change of the reactant into the product is shown through an arrow. Use of + sign is made when there are more than one reactant or there are more than one product. Let us see if you can complete the reaction given below:

Calcium Chlorine Calcium chloride

4.1.2 Writing a Chemical Equation

Is there any other shorter way for representing a chemical change? Yes this can be done through a chemical equation. A chemical equation can be made more concise and useful if we use chemical formulae instead of words. In the previous lesson you have already studied how to represent compound with the help of a chemical formula. Now if you substitute formulae of magnesium, oxygen and magnesium oxide for the words in equation (1), we get

$$Mg + O_2 \longrightarrow MgO \quad ...(3)$$

Similarly substituting formulae for words in equation B, we get,

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2 \quad ...(4)$$

Do you remember the ***Law of conservation of mass*** studied in the previous lesson? According to it, the mass and the number of atoms present in the reactant(s) should be equal to the mass and number of atoms present in product(s). Let us count the number of atoms on both sides (left hand side and right hand side) of the chemical equations (3) and (4). We find that in equation (3), the numbers of oxygen atoms on the right hand side and the left hand side are not equal. However in (4), the number of atoms on both the sides is equal. Such chemical equations in which the number of atoms is not equal on both sides of the arrow but still represent chemical reactions are called ***skeletal chemical equations***. Skeletal chemical equations can be balanced by using suitable *coefficients* in the equation. We shall study the balancing of chemical equation in the following section.

4.2 BALANCED CHEMICAL EQUATIONS

According to the law of conservation of mass, matter can neither be created nor destroyed. Thus, ***mass of each element present in the products of a chemical reaction must be equal to its mass present in the reactants***. In other words, the number of atoms of each element remains the same before and after a chemical reaction. In a balanced chemical equation number of atoms of a particular element present in the reactants and products must be equal. If not, equation is said to be 'not balanced.' Let us reconsider the above two equations (3) and (4).

i.e. $$Mg + O_2 \longrightarrow MgO \qquad ...(3)$$

and $$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2 \qquad ...(4)$$

Which one of the above two is balanced? It is quite obvious that equation (4) is balanced, as the number of Zn, H and S (sulphur) atoms are equal on both sides of the equation. Therefore equation (4) is said to be a *balanced chemical equation*. Now what about equation (3)? By simple inspection we can see that the number of atoms of magnesium in the reactant side is equal to the number of atoms of magnesium in the product side. However, the number of atoms of oxygen on the reactant side is two (in O_2) but only one atom of oxygen is in the product side in (MgO). To make the same number of atoms of oxygen in the product side, we shall have to write 2MgO. Now the equation becomes;

$$Mg + O_2 \longrightarrow 2MgO$$

In the above equation there is a shortage of one atom of magnesium on the left hand side. For balancing the number of magnesium atoms, we need to put 2 before Mg and the equation becomes,

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

Now the number of magnesium and oxygen atoms is equal on both sides of the arrow and the chemical equation is said to be balanced. This method of balancing of a chemical equation is called the ***Hit and Trial method***.

Let us consider another reaction for writing and balancing of a chemical reaction. When steam is passed over red hot iron, hydrogen gas (H_2) is evolved and magnetic oxide of iron (Fe_3O_4) is obtained. This can be expressed as:

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

If we examine the above equation we find that the equation is not balanced. Let us try to balance it using the following steps:

Step I: Study the equation carefully and write the number of atoms of different elements in the imbalanced equation:

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

Table 4.1 Comparing number of atoms of different elements in reactants and products

Element	Number of atoms on reactants side (LHS)	Number of atoms on products side (RHS)
Fe	1	3
H	2	2
O	1	4

Step II: We should start balancing with the compound that contains the maximum number of atoms. The compound may be a reactant or a product. In the compound, select the element which has maximum number of atoms. Based on this select Fe_3O_4 in the above equation. In Fe_3O_4 the element oxygen has the maximum number of atoms. There are four oxygen atoms on the right hand side and only one oxygen atom on the left hand side of the arrow. For balancing the oxygen atoms, we can put the coefficient '4' as '$4H_2O$'. Now the equation becomes:

$$Fe + 4H_2O \longrightarrow Fe_3O_4 + H_2 \text{ (partially balanced)}$$

Step III: Here Fe and H atoms are still not balanced. Let us balance the hydrogen atoms. For this, make the number of molecules of hydrogen as four on the RHS of the arrow. The equation now becomes:

$$Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2 \text{ (partially balanced)}$$

Step IV: Now, out of the three elements, only Fe remains imbalanced. For balancing iron, we write 3 atoms of iron on left hand side and the equation becomes:

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

Step V: Finally count the number of atoms of all the three elements on both sides of the arrow. You will find that the number of atoms of oxygen, hydrogen and iron on both sides of the arrow are equal and thus the balanced equation is obtained as:

$$3Fe + 4H_2O \longrightarrow Fe_3O_4 + 4H_2 \quad \text{(balanced equation)}$$

4.2.1 How can we make a Chemical Equation more Informative?

In the balanced equation

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

we have no information about the physical states of the reactants and the products i.e. whether they are solid, liquid or gas. By using (s) for solids, (l) for liquids and (g) for gases along with reactants and products, we can make a chemical equation more informative. Thus, the above equation can be written as:

$$3Fe\ (s) + 4H_2O\ (g) \longrightarrow Fe_3O_4\ (s) + 4H_2\ (g) \qquad ...(5)$$

Here, (g) by the side of H_2O clearly indicates that water used in the reaction is in the form of steam or gas. Further, if a reactant or a product is taken as solution in water, we denote it by writing (aq). For example.

$$\underset{\text{(quick lime)}}{CaO(s)} + H_2O\ (l) \longrightarrow \underset{\text{(slacked lime)}}{Ca(OH)_2\ (aq)} \qquad ...(6)$$

Sometimes the reaction conditions such as temperature, pressure, catalyst, etc. for the reaction are also indicated above and/or below the arrow in the equation. For example,

$$CO\ (g) + 2H_2\ (g) \xrightarrow{\text{340 atm}} CH_3OH\ (l) \qquad ...(7)$$

$$6CO_2\ (aq) + 6H_2O\ (l) \xrightarrow[\text{Chlorophyl}]{\text{Sunlight}} C_6H_{12}O_{16}\ (aq) + 6O_2\ (\ g) \qquad ...(8)$$

Important Tips for balancing a chemical equation

- Use the simplest possible set of whole number coefficients to balance a chemical equation. Normally we do not write fractional coefficients in such equations as molecules are not available in fractions. We multiply the equation by an appropriate number to ensure the entire equation has whole number coefficients.
- Do not change subscripts in formulae of reactants or of products during balancing, as that may change the identity of the substances. For example, $2NO_2$ means two molecules of nitrogen dioxide but if we double the subscript we get N_2O_4 which is the formula of dinitrogen tetroxide, a completely different compound.

4.3 SIGNIFICANCE OF A BALANCED CHEMICAL EQUATION

Qualitatively a chemical equation simply describes what the reactants and products are. However, *a balanced chemical equation gives a lot of quantitative information about a chemical reaction*. A balanced chemical equation tells us:

(i) the number of atoms and molecules taking part in the reaction and the corresponding masses in atomic mass unit (amu or u).

(ii) the number of moles taking part in the reaction, with the corresponding masses in grams or in other convenient units.

(iii) relationship between the volume of the reactants and the products if all of them are in the gaseous state.

4.3.1 Mole and Mass Relationships

Let us consider a chemical reaction between nitrogen and hydrogen in the presence of a catalyst.

$$\underset{\substack{\text{(1 molecule of}\\\text{nitrogen)}}}{N_2\ (g)} + \underset{\substack{\text{(3 molecules}\\\text{of hydrogen)}}}{3H_2\ (g)} \xrightarrow{\text{catalyst}} \underset{\substack{\text{(2 molecules of}\\\text{ammonia)}}}{2NH_3\ (g)} \qquad ...(9)$$

We may multiply the entire equation by any number, say 100, we obtain

$$\underset{\text{of nitrogen}}{1 \times 100 \text{ molecules}} + \underset{\text{of hydrogen}}{3 \times 100 \text{ molecules}} \longrightarrow \underset{\text{of ammonia}}{2 \times 100 \text{ molecules}}$$

Suppose, we multiply the entire equation by 6.022×10^{23}, (Avogadro's number) we get

$$\underset{\substack{\text{molecules of}\\\text{nitrogen}}}{1 \times 6.022 \times 10^{23}} + \underset{\substack{\text{molecules of}\\\text{hydrogen}}}{3 \times 6.022 \times 10^{23}} \longrightarrow \underset{\substack{\text{molecules of}\\\text{ammonia}}}{2 \times 6.00 \times 10^{23}}$$

Since 6.022×10^{23} molecules of any substance constitute its one mole, therefore, we can write

1 mole of nitrogen + 3 moles of hydrogen $\longrightarrow$ 2 moles of ammonia

Taking molar mass into consideration, we can write

(1×28.0) g of nitrogen + $(3 \times 2.0$ g of hydrogen $\longrightarrow$ (2×17) g of ammonia

or 28.0 g of nitrogen + 6.0 g of hydrogen $\longrightarrow$ 34.0 g of ammonia

Let us write the equation (9) once again,

$$N_2\ (g) + 3H_2\ (g) \xrightarrow{\text{catalyst}} 2NH_3\ (g)$$

1 molecule of nitrogen	3 molecules of hydrogen	$\longrightarrow$	2 molecules of ammonia
1 mol of nitrogen	3 moles of hydrogen	$\longrightarrow$	2 moles of ammonia

28.0 g of nitrogen + 6.0 g of hydrogen $\longrightarrow$ 34.0 g of ammonia

Remember

Quantity of a substance consumed or produced can be determined only if we use a balanced chemical equation.

4.3.2 Volume Relationship for Reactions involving gases

The French chemist, Gay Lussac found that the volume of reactants and products in gaseous state are related to each other by small integers, provided the volumes are measured at the same temperature and pressure.

Gay Lussac's discovery of integer ratio in volume relationship is actually the *law of definite proportion by volume*. Remember, the law of definite proportion studied in lesson 3: Atoms and Molecules, was with respect to masses.

Let us take the following example

$$\underset{\text{2 volumes}}{2H_2\,(g)} + \underset{\text{1 volume}}{O_2\,(g)} \longrightarrow \underset{\text{2 volumes}}{2H_2O\,(g)}$$

$$2 \text{ mol of } H_2 + 1 \text{ mol of } O_2 \longrightarrow 2 \text{ mol of } H_2O$$

[According to Avogadro's Law]

Here, hydrogen, oxygen and water vapours are at the same temperature and pressure (say 100°C and 1 atmospheric pressure). From this basic concept we can conclude that, if we take 100 mL of hydrogen and 50 mL of oxygen, we shall get 100 mL water vapour provided all volumes are measured at the same temperature and pressure. Thus, from a balanced chemical equation, we get relationship between mole, mass and volume of the reactants and products. This quantitative relationship has been found very useful in chemical calculations.

INTEXT QUESTION 4.1

1. Write a chemical equation for each of the following reactions:

   (i) Zinc metal reacts with aqueous hydrochloric acid to produce a solution of zinc chloride and hydrogen gas.

   (ii) When solid mercury(II) oxide is heated, liquid mercury and oxygen gas are produced.

2. Balance the following chemical equations:

   (i) $H_2SO_4\,(aq) + NaOH\,(aq) \longrightarrow Na_2SO_4\,(aq) + H_2O\,(l)$

   (ii) $Al\,(s) + HCl\,(aq) \longrightarrow AlCl_3\,(aq) + H_2\,(g)$

3. What is a balanced chemical equation? Why should a chemical equation be balanced?

4.4 TYPES OF CHEMICAL REACTIONS

So far we have studied how to express a chemical change in the form of an equation. We have also studied how to balance a chemical equation in order to derive useful quantitative information. We can classify chemical reactions into the following categories (i) combination reactions, (ii) decomposition reactions (iii) displacement reactions (iv) double displacement reactions.

4.4.1 Combination Reactions

In combination reactions, as the name indicates, two or more substances (elements or compounds) simply combine to form a new substance. For example, when a substance burns it combines with oxygen present in the air. In activity 4.1 we have seen that magnesium ribbon burns with dazzling light. During burning it combines with oxygen as

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

Now try the same with carbon.

$$C\ (s) + O_2\ (g) \longrightarrow CO_2\ (g)$$

Further, let us take a few activities.

ACTIVITY 4.2

Take a small amount of calcium oxide (CaO) or quick lime in a beaker. Now slowly add water to it (Fig. 4.2). Touch the side of the beaker with your hand. Do you feel any change in the temperature? Yes, it is warm to touch. You might have seen that for white-washing we put white solid material in water and after some time it starts boiling. This white material is calcium oxide and it reacts with water to form calcium hydroxide. Temperature rises due to evolution of heat in the reaction between quick lime and water. This reaction can be expressed in the form of the following equation:

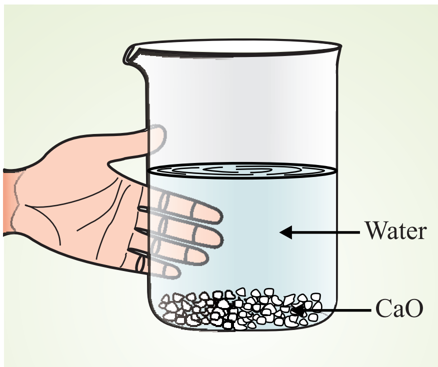

Fig. 4.2: *Reaction between quick lime and water*

$$\underset{\substack{\text{quick lime}\\\text{(Choona Patthar)}}}{CaO\ (s)} + H_2O\ (l) \longrightarrow \underset{\text{slaked lime}}{Ca(OH)_2\ (aq)} \qquad ...(10)$$

In the above reaction calcium oxide (quick lime) and water combine and form a single product-calcium hydroxide (slaked lime). Such ***a reaction in which a single product is formed from two or more reactants is known as combination reaction***.

In white washing, when slaked lime is applied on the walls it gradually reacts with carbon dioxide from the atmosphere. The bluish coloured calcium hydroxide (slaked lime) is converted into white calcium carbonate. After drying, it gives a white shiny finish. This reaction can be written as follows:

$$\underset{\text{Calcium hydroxide}}{Ca(OH)_2\ (aq)} + CO_2\ (g) \longrightarrow \underset{\text{Calcium carbonate}}{CaCO_3\ (s)} + H_2O\ (l) \quad ...(11)$$

It is interesting to note that chemical formula of marble is also $CaCO_3$.

In activities 4.1 and 4.2 you have seen that a lot of heat is evolved during the course of the reaction. Such reactions in which heat is released along with the formation of the products are called ***exothermic reactions***.

Other examples of exothermic reactions are:

(i) Burning of natural gas (CH_4) used for cooking.

$$CH_4(g) + 2O_2\ (g) \longrightarrow CO_2\ (g) + 2H_2O\ (g) \quad ...(12)$$

(ii) Respiration and digestion both are exothermic process. This heat energy comes from the food that we eat. Do you know what types of food give us energy? Food which we take in the form of rice, potatoes and bread contains ***carbohydrates***. Carbohydrates are broken down to glucose during digestion. The glucose combines with oxygen in the cells of our body and provides energy to our body.

$$C_6H_{12}O_6\ (aq) + 6O_2\ (aq) \longrightarrow 6CO_2\ (aq) + \text{energy} \quad ...(13)$$

People who do physical work, require a lot of energy and therefore, require carbohydrates in the form of sugar, potato, rice, bread, etc.

(iii) The decomposition of vegetable matter or *biomass* into compost is also an example of an exothermic reaction. If you have a compost pit in your surroundings, you can observe this yourself.

4.4.2 Decomposition Reactions

You have seen earlier that quick lime (*choona patthar*) solution is used for whitewashing of our houses. Have you ever thought how this quick lime is obtained? It is obtained by heating lime stone in a furnace (*bhatti*). Lime stone when heated gives lime and carbon dioxide.

$$\underset{\text{lime stone}}{CaCO_3\ (s)} \xrightarrow{\text{heat}} \underset{\text{quick lime}}{CaO\ (s)} + \underset{\text{carbon dioxide}}{CO_2\ (g)} \quad ...(14)$$

This reaction is an example of a decomposition reaction. **A decomposition reaction is the one in which a compound decomposes into two or more than two substances (elements or compounds).** Let us now carry out some activities.

ACTIVITY 4.3

Take about 2 g of ferrous sulphate in a hard glass test tube as shown in Fig. 4.3. Hold the test tube with a test tube holder and gently heat it over the flame. After heating for about one minute, observe the change in colour of ferrous sulphate. Smell the odour of the gas carefully. What do you observe? The green colour of the ferrous sulphate crystals gradually fades away and a smell of burning sulphur is found.

$$FeSO_4.7H_2O\ (s) \xrightarrow{\text{heat}} FeSO_4\ (s) + 7H_2O\ (g)$$

$$\underset{\text{ferrous sulphate}}{2FeSO_4\ (s)} \xrightarrow{\text{heat}} \underset{\text{ferric oxide}}{Fe_2O_3\ (s)} + \underset{\text{sulphur dioxide}}{SO_2\ (g)} + \underset{\text{sulphur trioxide}}{SO_3\ (g)} \quad ...(15)$$

Here ferrous sulphate ($FeSO_4.7H_2O$) crystal first loses water and then decomposes to SO_2 and SO_3 gases.

Another example of a decomposition reaction is given below:

$$\underset{\text{lead nitrate}}{2Pb(NO_3)_2\ (s)} \xrightarrow{\text{heat}} \underset{\text{lead oxide}}{2PbO\ (s)} + \underset{\text{nitrogen dioxide}}{4NO_2\ (g)} + O_2\ (g) \quad ...(16)$$

In the reactions given above, decomposition occurs by application of heat. Such reactions fall in the category of ***thermal decomposition***.

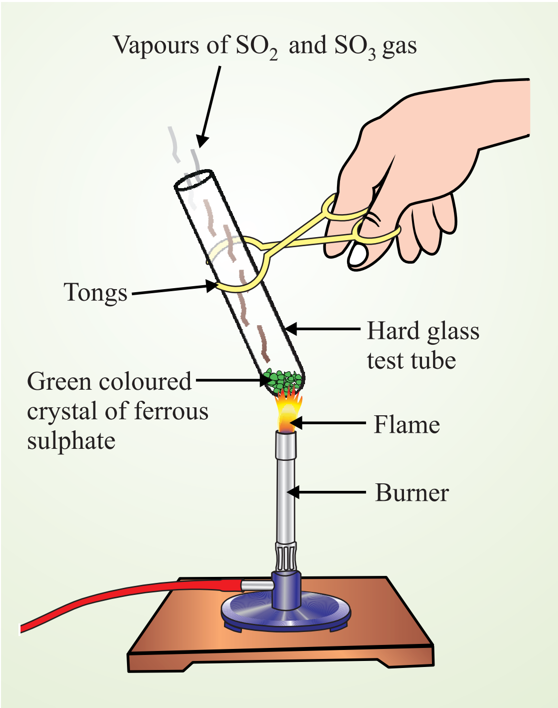

Fig. 4.3: *Thermal decomposition of ferrous sulphate*

ACTIVITY 4.4

Take a plastic mug. Drill two holes at its base and fit rubber stoppers in these holes. Insert graphite electrodes in these rubber stoppers as shown in Fig. 4.4. Connect these electrodes to a 6 volt battery.

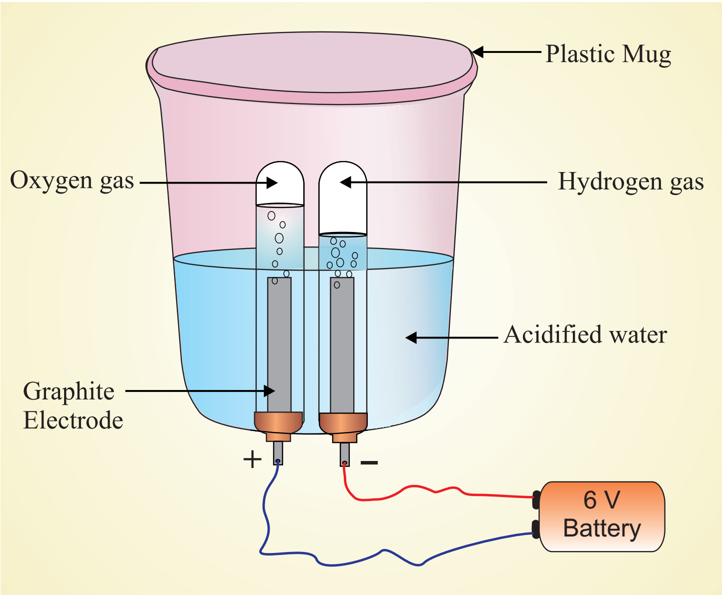

Fig. 4.4: *Electrolysis of water*

Now observe carefully what happens. You will find bubbles of gases over both the electrodes. Take two test tubes. Fill them with water and invert them over two graphite electrodes. Bubbles formed at the electrodes are found to replace water filled in the two test tubes. After sometime observe the volume of the two gases. You will find that the volume ratio of the two gases (oxygen and hydrogen) is 1:2. Carefully remove both the test tubes containing these gases one by one and test them with the help of your tutor at the study centre.

The two gases are hydrogen and oxygen and their volumes are in the ratio of 2:1 respectively (Gay Lussac's Law). The decomposition of water in this experiment takes place due to the electrical current that is passed through the water. A r***eaction in which a compound decomposes due to electrical energy into two or more than two substances (elements or compounds) is called electrolytic decomposition reaction***.

4.4.3 Displacement Reaction

For understanding this types of reaction, perform the following activity.

ACTIVITY 4.5

Take about 10 mL of dilute copper sulphate solution in each of the two test tubes and mark them as A and B. Now take two iron nails and clean them with sand paper. In test tube A, immerse one iron nail with the help of a thread as shown in Fig. 4.5. After nearly 20 minutes, observe the changes taking place on the surface of the iron nail and also in the colour of copper sulphate solutions. Compare the colour of the solution in test tube A with the colour of the solution in test tube B. What do you observe? The blue colour of copper sulphate solution fades. Similarly, compare the colour of the iron nail dipped in solution A with the other iron nail. You will see that the surface of the nail has become brownish. Do you know why the iron nail becomes brownish and the blue colour of copper sulphate solution fades?

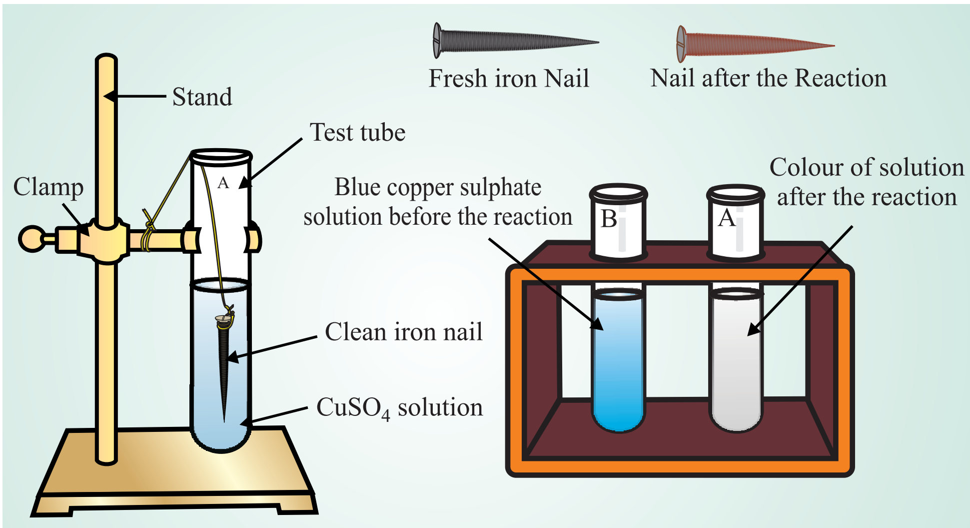

Fig. 4.5: *Reaction between iron and copper sulphate*

All this happens due to the following chemical reaction,

$Fe\ (s)$ +	$CuSO_4\ (aq)$	$\longrightarrow$	$FeSO_4\ (aq)$ +	$Cu\ (s)$	...(17)
iron	copper sulphate		ferrous sulphate	copper	

In this reaction, one element i.e. iron has displaced another element i.e. copper from copper sulphate solution. These types of reactions fall in the category of ***displacement reactions. The displacement reaction is one in which one element displaces another element from its compounds.***

Other examples of displacement reactions are:

$$Zn\ (s) + CuSO_4\ (aq) \longrightarrow ZnSO_4\ (aq) + Cu\ (s) \qquad ...(18)$$

$$Pb\ (s) + CuCl_2\ (aq) \longrightarrow PbCl_2\ (aq) + Cu\ (s) \qquad ...(19)$$

Since zinc and lead are more reactive metals than copper therefore they displace copper from its compound.

4.4.4 Double Displacement Reactions

For understanding this type of reactions, perform the following activity.

ACTIVITY 4.6

Take two test tubes and mark them A and B. In test tube A take nearly 4 mL of sodium sulphate solution and in test tube B take nearly 4 mL of barium chloride solution. Now add solution of test tube A to solution of test tube B. What do you observe? A white substance is formed which is known as a ***precipitate***. The reaction can be written as,

$$\underset{\text{sodium sulphate}}{Na_2SO_4\ (aq)} + \underset{\text{barium chloride}}{BaCl_2\ (aq)} \longrightarrow \underset{\substack{\text{barium sulphate}\\\text{(white ppt)}}}{BaSO_4\ (s)\downarrow} + \underset{\text{sodium chloride}}{2NaCl\ (aq)} \qquad ...(20)$$

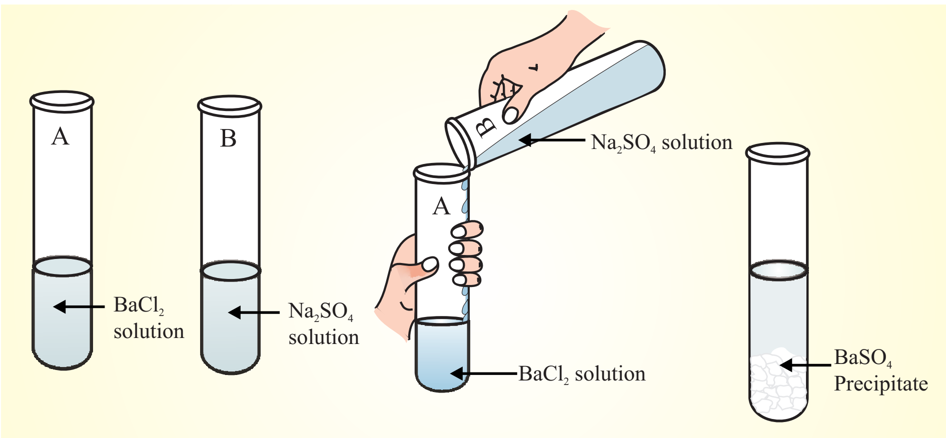

Fig. 4.6: *Precipitation reaction between sodium sulphate and barium chloride*

The white precipitate of $BaSO_4$ is formed by the reaction of Ba^{2+} ions and SO_4^{2-} ions. The other product formed is sodium chloride which remains in solution. ***Reactions in which there is an exchange of ions between the reactants, are called double displacement reactions.***

Find out different types of reaction occuring in your compounds.

4.5 OXIDATON AND REDUCTION (REDOX REACTION)

In order to understand the redox reactions, let us perform the following activity.

ACTIVITY 4.7

Take a china dish containing nearly 2 g of copper powder and heat it strongly as shown in Fig. 4.7. What do you observe? Copper powder becomes black. Why? This is because when oxygen combined with copper, copper oxide is formed which is black in colour. This reaction can be written as,

$$\underset{\text{brown}}{2Cu\ (s)} + O_2\ (g) \xrightarrow{\text{heat}} \underset{\text{black}}{2CuO\ (s)} \qquad ...(21)$$

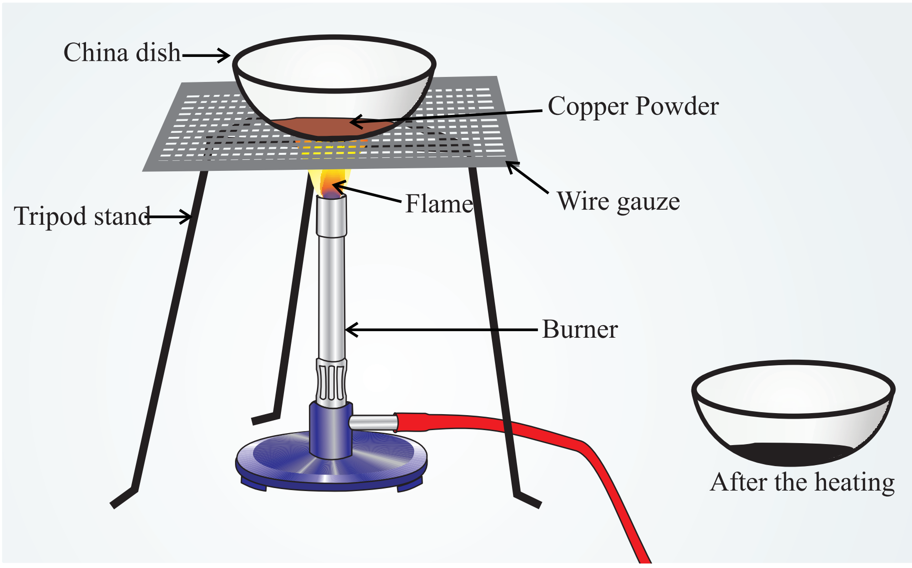

Fig. 4.7: *Heating of copper powder in air*

Now if you pass hydrogen gas over this black powder (CuO) you will observe that the surface of the black powder becomes brown, which is the original colour of the copper. This reaction can be written as,

$$\underset{\text{black}}{CuO\ (s)} + H_2\ (g) \longrightarrow \underset{\text{brown}}{Cu\ (s)} + H_2O\ (l) \qquad ...(22)$$

In reaction (21) copper gains oxygen and is said to be oxidized. In reaction (22) copper oxides loses oxygen and is said to be reduced. Hydrogen in this reaction is gaining oxygen and is thus being oxidized. ***When a substance gains oxygen during a reaction, it is said to be oxidized and when a substance loses oxygen during a reaction, it is said to be reduced.***

Thus in this reaction, during the reaction process, one reactant gets oxidized while the other gets reduced. Such reactions are called ***oxidation reduction reaction or Redox Reactions***. This can be depicted in the following way:

Reduction

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

Oxidation

In the above scheme, CuO provides oxygen and therefore is an oxidizing agent and hydrogen takes this oxygen and therefore is a reducing agent. In a redox reaction, an oxidizing agent is reduced and a reducing agent is oxidized.

Some other examples of redox reaction are :

$$ZnO\,(s) + C\,(s) \xrightarrow{\text{heat}} Zn\,(s) + CO\,(g) \qquad ...(23)$$

$$MnO_2\,(s) + 4HCl\,(aq) \longrightarrow MnCl_2\,(aq) + 2H_2O\,(l) + Cl_2\,(g) \qquad ...(24)$$

In all redox reactions, you have seen that one species is oxidized and the other is reduced. ***There is no oxidation without reduction and there is no reduction without oxidation.*** This aspect of redox reactions will be explained broadly in terms of ***electron gain*** and ***electron loss*** in the following section.

4.5.1 Redox Reactions in terms of Electron gain and Electron Loss

You just learnt oxidation and reduction in terms of gain and loss of oxygen and hydrogen. However, defining a redox reaction in this way is confined to only a few reactions.

Let us consider the reactions

$$Cu\,(s) + I_2\,(s) \xrightarrow{\text{heat}} CuI_2\,(s) \qquad ...(25)$$

$$Fe\,(s) + S\,(s) \xrightarrow{\text{heat}} FeS\,(s) \qquad ...(26)$$

These reactions do not involve any gain or loss of oxygen or hydrogen. Yet these are oxidation-reduction reactions. The reaction (25),

$$Cu\,(s) + I_2\,(s) \longrightarrow CuI_2\,(s)$$

can be written in two steps as follows:

Step (i):

$$\underset{\text{copper atom}}{Cu} \longrightarrow \underset{\text{copper ion}}{Cu^{2+}} + \underset{\text{electrons}}{2e^-}$$

Step (ii):

$$\underset{\text{iodine}}{I_2} + \underset{\text{electrons}}{2e^-} \longrightarrow \underset{\text{iodide ion}}{2I^-}$$

In step (i) one copper atom loses two electrons to become a cupric ion, Cu^{2+} and in step (ii) iodine gains two electrons and gets converted into two iodide ions. Here we say, copper is oxidized by losing electrons and iodine is reduced by gaining electrons. Thus ***a reaction in which a species loses electrons is called an oxidation reaction and a reaction in which a species gains electrons is called a reduction reaction.*** The substance which oxidizes the other substance is known as an ***oxidizing agent***. An oxidizing agent gets reduced during the reaction. Likewise, the substance which reduces the other substance is known as a ***reducing agent***. A reducing agent gets oxidized during the reaction. In reaction (25), copper acts as a reducing agent and iodine as an oxidizing agent.

Similarly, in reaction (26) iron acts as a reducing agent and sulphur as an oxidising agent.

Step (i):

$$\underset{\text{iron atom}}{Fe} \longrightarrow \underset{\text{ferrous ion}}{Fe^{2+}} + \underset{\text{electrons}}{2e^-}$$

Step (ii):

$$\underset{\text{sulphur atom}}{S} + \underset{\text{electrons}}{2e^-} \longrightarrow \underset{\text{sulphide ion}}{S^{2-}}$$

Now, you can answer the following in the space provided

(i) Reducing agent: ....

(ii) Oxidising agent: ....

(iii) element which is oxidised: ....

(iv) Element which is reduced ....

[**Hint:** Your answer should be as per rule given below]

Gain of electron is reduction and loss of electron is oxidation.

As mentioned earlier, oxidation and reduction processes occur simultaneously.

Consider the following displacement reaction

$$Zn\ (s) + CuSO_4\ (aq) \longrightarrow ZnSO_4\ (aq) + Cu\ (s)$$

or $$Zn\ (s) + Cu^{2+}\ (aq) \longrightarrow Zn^{2+}\ (aq) + Cu\ (s) \qquad ...(28)$$

Here, Zn loses electrons and gets converted into Zn^{2+} (aq). These electrons lost by Zn are gained by Cu^{2+} ion which gets converted into Cu. This broad definition of reduction-oxidation can be applied to many more reactions.

A few more examples of redox reaction are given below:

$$Fe_2O_3\ (s) + 2Al\ (s) \longrightarrow Al_2O_3\ (s) + 2Fe\ (s)$$

$$2Na\ (s) + Cl_2\ (g) \longrightarrow 2NaCl\ (s)$$

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

INTEXT QUESTIONS 4.2

1. Examine the following reaction(s) and identify which of them are **not** example(s) of a redox reaction?

   (i) $AgNO_3\ (aq) + HCl\ (aq) \longrightarrow AgCl\ (s) + HNO_3\ (aq)$

   (ii) $MnO_2\ (s) + 4HCl\ (aq) \longrightarrow MnCl_2\ (aq) + 2H_2O\ (l) + Cl_2\ (g)$

   (iii) $4Na\ (s) + O_2\ (g) \longrightarrow 2Na_2O\ (s)$

2. Identify the substances which are oxidized and the substances that are reduced in the following reactions:

   (i) $H_2\ (g) + Cl_2\ (g) \longrightarrow 2HCl\ (g)$

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

   (iii) $Zn\ (s) + 2AgNO_3\ (aq) \longrightarrow Zn(NO_3)_2\ (aq) + 2Ag\ (s)$

4.5.2 Effect of Redox Reaction in Everyday Life

We have studied different types of chemical reactions in the previous sections. Out of these reactions, redox reactions are very important in our lives. We would like to discuss corrosion in view of its economic importance. Rancidity is also important in view of its direct link with our foods and edibles. Both of these i.e. corrosion and rancidity are results of redox reactions.

- Corrosion
- Rancidity

A substance capable of destroying bacteria is called a disinfectant or a bactericide or an antiseptic. Most effective disinfectants are strong oxidizers

A bleach oxidises colored compounds to other substance which are not coloured.

Many disinfectants including chlorine which are available in different forms as solid compounds such as calcium hypochlorite, $Ca(ClO)_2$, are oxidising agents.

In an oxy-acetylene torch used for welding and cutting metals, acetylene is oxidised and produces very high temperature.

A. Corrosion

Corrosion is a destructive chemical process in which metals are oxidized in presence of air and moisture. The rusting of iron, tarnishing of silver, development of green coatings on copper, brass and bronze are a few examples of corrosion. It causes enormous damage to bridges, ships, cars and to all machines which are made of iron or steel. The damage and efforts taken to prevent it costs several crores of rupees a year. Preventing corrosion is a big challenge for an industrially developing country like ours.

ACTIVITY 4.8

Take 3 small beakers and mark them as A, B and C. In each beaker put 3 g of iron nail. In beaker A nothing is added but its mouth is covered with a watch glass. In beaker B add a few drops of water and make the iron nail wet. Leave the beaker B open, i.e. exposed to the atmosphere. In beaker C, add enough water to cover the nail completely (Fig. 4.8). Leave all the beakers for about three days. Observe the changes in all the beakers. Iron nail in beaker A are not affected, in beaker B the iron nail are rusted and in beaker C again the iron nail are not affected. Now write the condition for rusting on the basis of your findings.

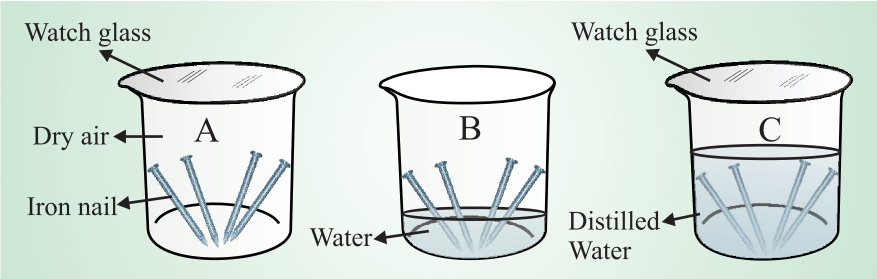

Fig. 4.8

How does one prevent corrosion?

There are several methods for protecting metals from corrosion, especially iron from rusting:

- plating the metal (iron) with a thin layer of less easily oxidized metal like nickel or chromium. This plating keeps out air (oxygen) and moisture which are main causes of corrosion.
- coating/connecting the metal with more reactive metal or with a metal which is more easily oxidized. For example, iron is connected to magnesium or coated with zinc for protecting it from corrosion. Iron rods are dipped in molten zinc to create a layer on their surface. This process of zinc coating over iron is called ***galvanization***.
- applying a protective coating such as paint.

Fig. 4.9: *Rusted nuts and bolts of iron*

B. Rancidity

You might have tasted or smelt fat/oil containing food material left for a long time. What do you find? You will find a lot of difference in the smell of fresh and stale oil or ghee. Why does this happen? This happens because fats and oils undergo oxidation and become rancid. This change is called ***rancidity***. Oxidation of fats/oils results into the formation of acids. These acids give unpleasant smell and bad taste.

Many food items which are cooked/fried in oil/fat are kept in air tight containers for sale. Keeping food items in air tight containers helps to slow down the oxidation process. Usually substances which prevent oxidation (anti-oxidants) are added to food items containing fats and oils. Do you know that the chips manufacturers usually flush bags of chips with a gas such as nitrogen to prevent oxidation of oil present in chips?

WHAT YOU HAVE LEARNT

- A chemical equation is a shorthand description of a reaction. It symbolically represents the reactants, products and their physical states.
- In a balanced chemical equation, number of atoms of each type involved in the chemical reaction is equal on the reactants and products sides of the equation.
- If charged species are involved, the sum of the charges on reactants should be equal to sum of charges on the products.
- During balancing of a chemical equation, no change in the formula of reactant(s) and product(s) is allowed.
- A balanced chemical equation obeys the law of conservation of mass and the law of constant proportions.
- In a combination reaction two or more substances combine to form a new single substance.
- In a decomposition reaction, a single substance decomposes to give two or more substances. Thus decomposition reactions are opposite to combination reactions.
- Reactions in which heat is given out during product formation are called **exothermic reactions** and reactions in which heat is absorbed during product formation are called **endothermic reactions**.
- A displacement reaction is one in which an element displaces another element from its compound.
- When two different ions are exchanged between two reactants double displacement reaction occurs.
- Precipitation reactions are the result of ion exchange between two substances, producing insoluble salts.

- Oxidation is the gain of oxygen or loss of hydrogen and reduction is loss of oxygen or gain of hydrogen. Oxidation and reduction reactions occur simultaneously and are jointly called ***redox reactions***.
- Redox reactions can broadly be defined in terms of loss and gain of electrons. Gain of electron(s) is reduction and loss of electrons is oxidation.
- Redox reactions are very important in our life situations as well as in industries.

TERMINAL EXERCISE

1. A. Write chemical equations of the following and balance them:

   (a) Carbon + oxygen $\longrightarrow$ Carbon dioxide

   (b) Hydrogen + Chlorine $\longrightarrow$ Hydrogen chloride

   (c) Barium chloride + Sodium sulphate $\longrightarrow$ Barium sulphate + sodium chloride

   B. Write balanced chemical equations with physical state symbols and necessary conditions, if any:

   (a) Nitrogen reacts with hydrogen in the presence of iron as a catalyst at 200 atmospheric pressure and 600°C temperature, and the product obtained is ammonia.

   (b) Aqueous solution of sodium hydroxide reacts with hydrochloric acid and produces sodium chloride and water.

   (c) Phosphorus burns in chlorine gas to form phosphorous pentachloride.

   C. Balance the following chemical reactions:

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

   (b) $BaCl_2\ (aq) + H_2SO_4\ (aq) \longrightarrow BaSO_4\ (s) + HCl\ (aq)$

   (c) $CuSO_4\ (aq) + Zn\ (s) \longrightarrow ZnSO_4\ (aq) + Cu\ (s)$

   (d) $H_2S\ (g) + SO_2\ (g) \longrightarrow S\ (s) + H_2O\ (l)$

   (e) $BaCl_2\ (aq) + Al_2(SO_4)_3\ (aq) \longrightarrow AlCl_3\ (aq) + BaSO_4\ (s)$

   (f) $Pb\ (NO_3)_2\ (aq) + Fe_2(SO_4)\ (aq) \longrightarrow Fe(NO_3)_3\ (aq) + PbSO_4\ (s)$

   (g) Calcium hydroxide + carbon dioxide $\longrightarrow$ Calcium carbonate + water

   (h) Aluminium + Copper (II) chloride $\longrightarrow$ Aluminium chloride + copper

   (i) Calcium carbonate + hydrochloric acid $\longrightarrow$ Calcium chloride + water + carbon dioxide

2. What is a balanced chemical equation? Write 3 characteristics of a balanced chemical equations?

3. In what way is a displacement reaction different from a double-displacement reaction? Explain with two suitable examples.

4. What happens when dilute hydrochloric acid is added to iron filings? Mark (√) at the correct answer from the following:

   (a) Hydrogen gas and iron chloride are produced and is classified as a displacement reaction.

   (b) Iron chloride and chlorine gas are produced and is classified as a decomposion reaction.

   (c) Iron hydroxide and water are produced and is classified as a combination reaction.

   (d) No reaction takes place but is classified as a double displacement reaction.

7. What do you mean by an exothermic reaction? Give a suitable example.

8. Classify each of the following reactions as combination, decomposition, displacement or double displacement reactions:

   (a) $Zn\ (s) + 2AgNO_3\ (aq) \longrightarrow Zn(NO_3)_2 + 2Ag\ (s)$

   (b) $2KNO_3\ (s) \xrightarrow{\text{heat}} 2KNO_2 + O_2\ (g)$

   (c) $Ni\ (NO_3)_2\ (aq) + 2NaOH\ (aq) \longrightarrow Ni(OH)_2\ (s) + 2NaNO_3\ (aq)$

   (d) $2KClO_3\ (s) \xrightarrow{\text{heat}} 2KCl\ (s) + 3O_2\ (g)$

   (e) $MgO\ (s) + C\ (s) \longrightarrow CO\ (g) + Mg\ (s)$

9. What is the difference between a combination and a decomposition reaction? Illustrate with suitable examples.

10. Is there any oxidation without reduction? Justify your answer.

11. 'Both combination reaction and displacement reaction fall in the category of redox reactions'. Do you agree? If so discuss this aspect with suitable examples.

12. Give two examples from everyday life situation where redox reaction takes place. How will you prove it?

13. In the following reactions name the substances which are oxidized and reduced and also mention the oxidizing and reducing agents:

   (a) $Ca\ (s) + Cl_2\ (g) \xrightarrow{\text{heat}} CaCl_2\ (s)$

Notes

(b) $3MnO_2\ (s) + 4Al\ (s) \xrightarrow{\text{heat}} 3Mn\ (l) + 2Al_2O_3\ (s)$

(c) $Fe_2O_3\ (s) + 3CO\ (g) \xrightarrow{\text{heat}} 2Fe\ (s) + 3CO_2\ (g)$

14. Explain the following in terms of electron transfer:

(a) Oxidation (b) Reduction

17. What is the law of definite proportion by volume? Explain.

ANSWERS TO INTEXT QUESTIONS

4.1

1. (i) $Zn(s) + 2HCl\ (aq) \longrightarrow ZnCl_2(aq) + H_2(g) + H_2$

   (ii) $2HgO\ (s) \longrightarrow 2Hg\ (l) + O_2$

2. (i) $H_2SO_4\ (aq) + 2NaOH\ (aq) \longrightarrow Na_2SO_4\ (aq) + 2H_2O\ (l)$

   (ii) $2Al\ (s) + 6HCl\ (aq) \longrightarrow 2AlCl_3\ (aq) + 3\ H_2\ (g)$

3. Volume of reactant and products in gaseous chemical reactions are related to each other by small integers, provided the volume are measured at the same temperature and pressure. In a balanced gaseous chemical equation we get relation between volume and between the moles of the reactants and products.

4.2

1. Following equation is not example of a redox reaction :

   (i) $AgNO_3(aq) + HCl\ (aq) \longrightarrow AgCl(s) + HNO_3(aq)$

2. (i) H_2 is oxidized and Cl_2 is reduced.

   (ii) H_2 is oxidized and CuO is reduced.

   (iii) Zn is oxidized and Ag^+ (in $AgNO_3$) is reduced.