

5

Transformation of Substances

Whenever there is a change in the characteristics of a substance, a chemist wants to know what kind of change has occurred— **physical** or **chemical**. You have learnt earlier how to distinguish between physical and chemical changes. Let us recall the following.

	Physical change	Chemical change
Definition	A change in which no new substances are formed and which can be reversed by reversing the conditions is called a physical change.	A change in which new substances are formed and which cannot be reversed by reversing the conditions is called a chemical change.
Differences	1. A physical change is temporary. 2. A physical change can be reversed by reversing the conditions. 3. No new substances are formed after a physical change. 4. After a physical change, the mass of the substance does not change.	1. A chemical change is permanent. 2. A chemical change cannot be reversed by reversing the conditions. 3. New substances are formed after a chemical change. 4. After a chemical change, the mass of any individual substance changes.
Examples	1. A change in the state of matter, e.g., (a) the melting of ice (b) the freezing of water (c) the vaporisation of water (d) the condensation of water vapour (e) sublimation 2. The glowing of an electric heater or a bulb	1. The burning of coal or wood 2. The rusting of iron 3. The curdling of milk 4. The charring of sugar 5. Photosynthesis in plants 6. The digestion of food



Fig. 5.1 Examples of physical and chemical changes

Change in State

Change in state is an important physical change. Let us consider the salient features of such a change.

Energy Change in Change in State

Energy change is a general feature of all changes—physical or chemical. So, a change in state is also accompanied by an energy change. This will be evident from the following observations.

1. If you put some ice in a glass, water droplets will collect on the outer surface of the glass. Ice needs heat to melt, which it takes from the surroundings. On giving heat to the ice, the water vapour in the air condenses on the outer surface of the glass.



Fig. 5.2 Droplets of water collect on the outer surface of a glass containing ice.

2. If you have ever used Vicks or Iodex, you would have felt a cooling sensation on your skin. This is because a constituent of the balm evaporates, and takes away the heat required for vaporisation from your skin. The evaporation of a liquid is always accompanied by the cooling of the surface from which the liquid evaporates.
3. Water vapour condenses on a cold surface, making it hot. If some cold water, placed in a pot, is held over the water vapour coming from boiling water, the



Fig. 5.3 Vapours give out heat while condensing.

cold water becomes warmer. This is because the vapours give out heat while condensing.

Changes in State Occur at Fixed Temperatures

For a pure substance, a change in state occurs at a fixed temperature (provided we do not change the pressure). In all cases, 1 atmosphere is taken as the standard pressure.

A solid melts at a fixed temperature, called its melting point, and a liquid boils at a fixed temperature, called its boiling point. A liquid freezes at its freezing point and a gas (or vapour) liquefies at its temperature of liquefaction. Interestingly, **the melting point of a solid is the same as the freezing point of the corresponding liquid. And the boiling point of a liquid is the same as the temperature of liquefaction of the corresponding gas.**

The melting point and the boiling point of a pure substance are fixed at a particular pressure.

For example, at a pressure of 1 atmosphere, ice melts at 0°C and water freezes at 0°C , and water boils at 100°C and steam liquefies at 100°C . However, the melting point of a solid decreases and the boiling point of a liquid increases with pressure.

Determining the melting point and the boiling point of a substance

The melting point of ice can be determined by placing the bulb of a thermometer in crushed

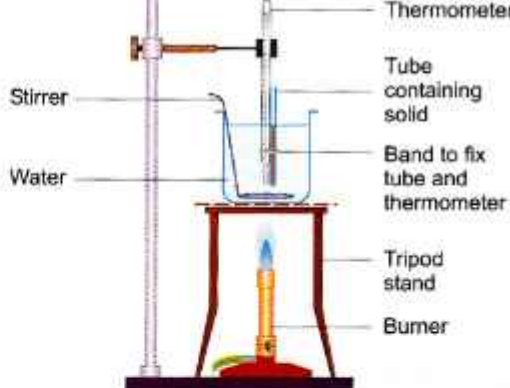
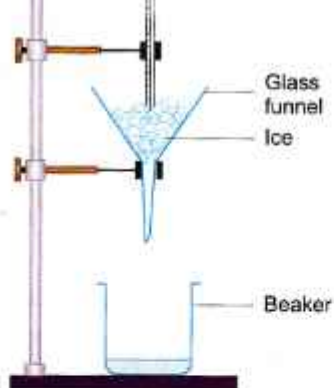


Fig. 5.4 Determining the melting point of (a) ice, and (b) any solid

ice (Figure 5.4a), and recording the temperature when it becomes constant. For other solids, another method (Figure 5.4b) is used. A small amount of the solid is placed in a narrow glass tube, called a **capillary tube**, which is tied to the bulb of the thermometer. The thermometer, along with the capillary tube, is placed in a suitable bath (a beaker or flask, containing a liquid that boils at a much higher temperature than the melting point of the solid) and heated. The temperature at which the solid melts sharply is recorded.

The boiling point of a liquid is determined using a set-up similar to the one shown in Figure 5.5. A thermometer is introduced into a flask in which the liquid is boiled. The temperature that becomes constant when the liquid keeps boiling is recorded as the boiling point of the liquid.

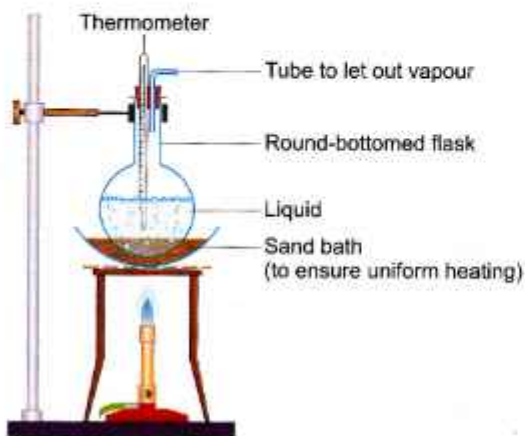


Fig. 5.5 Determining the boiling point of a liquid

The effect of impurities on the melting point and the boiling point of a substance

The melting point and the boiling point of a substance change when soluble impurities are present in it.

Melting point When there are soluble impurities in a substance, its melting point is lowered. For example, sodium chloride is soluble in water, and so the melting point of water (in this case, ice) containing sodium chloride is lower than 0°C at 1 atmosphere. Mix some common salt with ice and record the temperature—it will go to the minus side. This is the principle on which a **freezing mixture** works. A freezing mixture is a mixture of ice and common salt or calcium chloride, and it helps us attain as low a temperature as -5° , -10° or -15°C , which may be required for an experiment. Ice-cream vendors store ice cream in a freezing mixture.

Boiling point If there are soluble impurities in a substance, its boiling point is raised.

Boil some water in a beaker and place a thermometer in the boiling water (Figure 5.6). Record the temperature. Dissolve some common salt or sugar in the water, boil it again and record the temperature at which it boils. The water with common salt or sugar dissolved in it boils at a higher temperature than the plain water.

Tap water is not pure water. It contains some dissolved calcium and magnesium salts. So, the boiling point of tap water is slightly higher

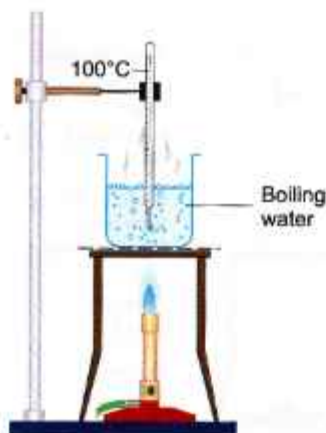


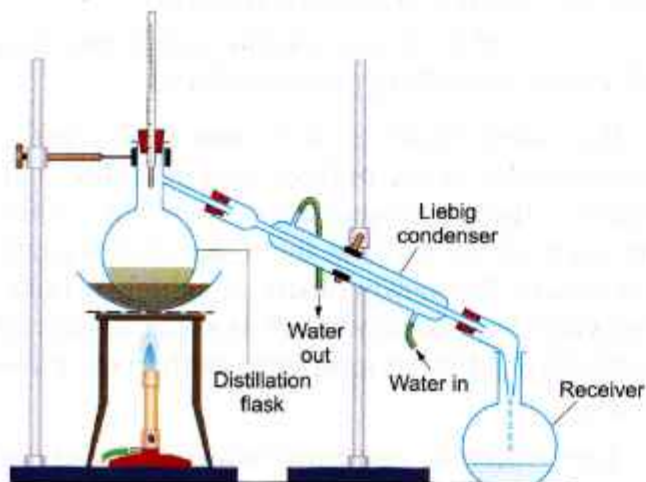
Fig. 5.6 Water boils at a temperature higher than 100°C when it contains a dissolved substance.

than that of distilled water, i.e., pure water. This can be verified by an experiment identical to the one we have just discussed. That tap water contains dissolved impurities can be easily proved as follows. Heat a sample of tap water until all the water evaporates. You will find that a white residue is left behind. Repeat the experiment with distilled water—there will be no residue.

Separating Pure Liquids by Distillation

From a mixture of a pure liquid and a solid (e.g., an aqueous solution of common salt) the liquid can be separated by **distillation**.

Distillation is performed with a set-up similar to the one shown in Figure 5.7. The mixture is boiled in a flask and the emerging vapours are cooled in a Liebig condenser and



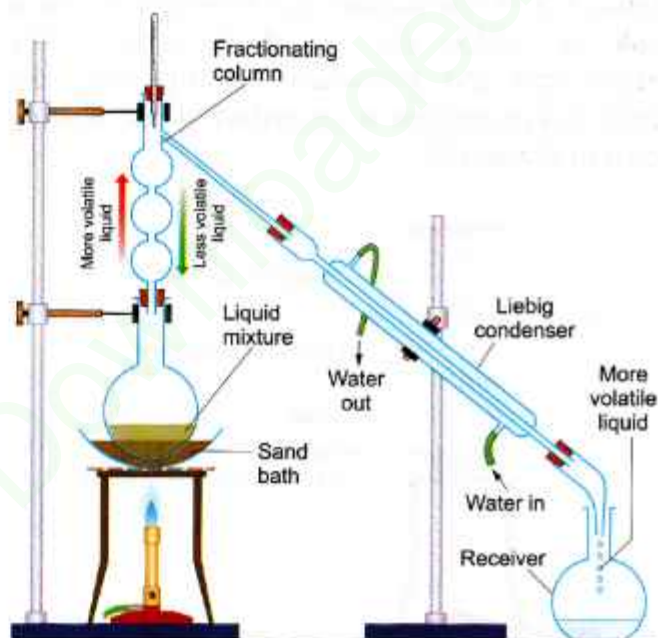
collected in a flask. A Liebig condenser is a long tube provided with a jacket, through which water is circulated for cooling the vapours in the tube.

We prepare distilled water in the laboratory by this method.

Fractional distillation

We have touched upon fractional distillation in an earlier chapter. Let us discuss this method here in a little greater detail.

Fractional distillation is used for separating liquids which differ in their boiling points by about 20°C or more. The mixture of liquids is boiled in a flask and the vapours are passed through a fractionating column. The vapours of the liquid with a lower boiling point (i.e., that which is more volatile) rise and the vapours of that which has a higher boiling point (i.e., that which is less volatile) condense in the bulblike traps. The higher-boiling liquid may also come back to the flask. The lower-boiling liquid gets into the condenser first and collects in the receiving flask. As long as the lower-boiling liquid continues to distil, the temperature remains constant. After the lower-boiling liquid has distilled off, the higher-boiling liquid begins to distil and the temperature rises. The



receiver is immediately changed and the higher-boiling liquid is collected separately. The separately collected liquids are called **fractions**.

By fractional distillation, we can separate benzene (boiling point 80°C) from toluene (boiling point 110°C), and the different fractions from crude oil.

Chemical Reactions

Energy Change in a Chemical Reaction

You have learnt earlier that the atoms of the reacting substances rearrange themselves during a chemical reaction. This can happen only if the bonds in the reacting substances are broken and new bonds are formed.

For example, hydrogen reacts with oxygen to form water in the following steps.

1. The hydrogen molecules break up into atoms and so do the oxygen molecules.
2. One atom of oxygen combines with two atoms of hydrogen to form a molecule of water.

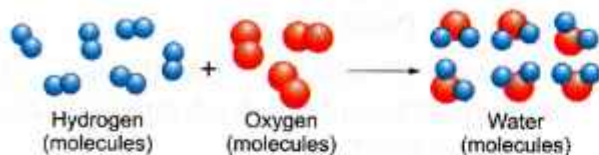


Fig. 5.9 The atoms of hydrogen and oxygen rearrange themselves to form water.

Energy is required (i.e., spent) for breaking a bond, but given out when a new bond is formed. Thus, the energy spent for breaking a bond is compensated fully or partially by the energy released in the formation of a bond. As the bonds broken and those formed are different, they involve different amounts of energy. So, the compensation is either partial or more than full. If the compensation is only partial, we will have to give energy for the reaction to occur. And, if the compensation is more than full, we will get energy from the

reaction, i.e., energy will be evolved. Thus, chemical reactions always involve an energy change. Let us look at a couple of illustrative examples.

1. In a reaction, suppose you need 500 kJ of energy to break the bonds of the reactants, but you get only 400 kJ of energy from the formation of the new bonds. So, you will have to spend 100 kJ of energy for the reaction to occur.
2. In another reaction, suppose you need 500 kJ of energy to break the bonds of the reactants, but you get 600 kJ of energy from the formation of new bonds. So, 100 kJ of energy will be released during this reaction.

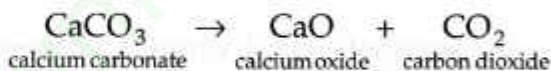
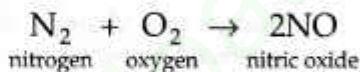
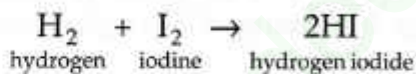
In most cases, energy is absorbed or given out in the form of heat. However, in many cases, light is also absorbed or given out.

A reaction in which heat is absorbed is called an endothermic reaction.

A reaction in which heat is evolved (i.e., given out) is called an exothermic reaction.

Examples of endothermic reactions

Endothermic reactions are few. Some examples are given below.



Examples of exothermic reactions Most chemical reactions are exothermic. Here are a few examples.

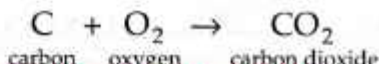
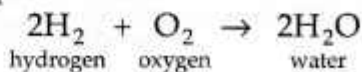
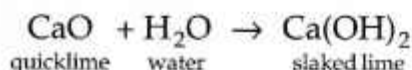




Fig. 5.10 The slaking of lime is a highly exothermic reaction.

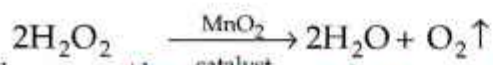
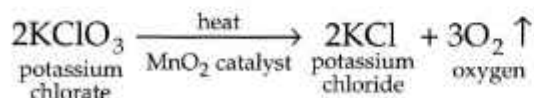


Catalysts

Some reactions are slow, whereas some others are fast. For example, rusting is slow, but the action of sodium on water is fast. So, a reaction has its own speed. But the speed (or rate) of a reaction can be changed by using what is known as a **catalyst**.

A catalyst is a substance that changes the speed (or rate) of a reaction, but itself remains unchanged.

A commonly used catalyst is manganese dioxide (MnO_2), which speeds up the evolution of oxygen from potassium chlorate or hydrogen peroxide.



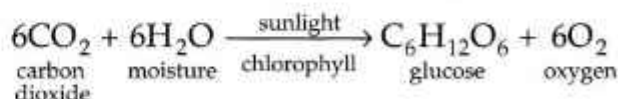
Catalysts in industries

Catalysts are used in the manufacture of many substances, e.g., ammonia, sulphuric acid, solid or hardened fats and polymers.

Iron is used as a catalyst in the manufacture of ammonia, and platinum in that of sulphuric acid. Nickel is used as a catalyst for manufacturing solid or hardened fats and very complicated substances for manufacturing polymers.

Catalysts in living organisms

Most reactions in living organisms are catalysed. For example, photosynthesis takes place in plants in the presence of chlorophyll.



In living organisms, catalysts of a special kind, known as **enzymes**, facilitate chemical reactions. Some examples are given below.

1. Carboxypeptidase, which helps digest our food
2. Nitrogenase, which helps plants take up nitrogen (the process is called nitrogen fixation in plants)
3. Catalase, which facilitates the decomposition of any hydrogen peroxide formed in our body

$$(2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2)$$
4. Carbonic anhydrase, which catalyses the hydration of carbon dioxide, forming carbonic acid ($\text{CO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{CO}_3$)

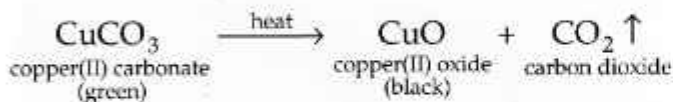
Characteristics of Chemical Reactions

Several characteristic changes accompany a chemical reaction. In many cases, you can guess what has happened by observing these changes.

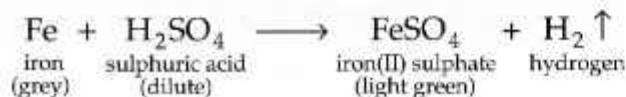
Change in colour

A change in the colour of a substance generally suggests that a chemical reaction has taken place. Let us look at some examples.

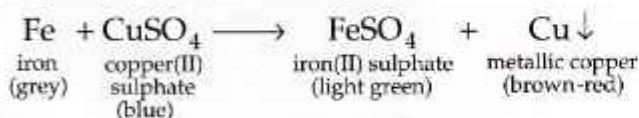
1. When heated, copper(II) carbonate (green) turns black due to the formation of copper(II) oxide.



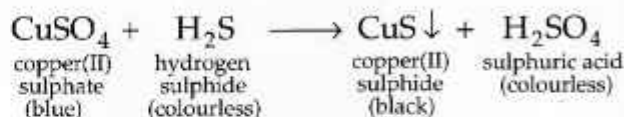
2. When dissolved in dilute sulphuric acid, iron nails (grey) give a light green solution of iron(II) sulphate.



3. When put into an aqueous solution of copper(II) sulphate (blue), an iron nail (grey) turns the solution light green. Also, there is a brown-red solid deposit over iron. You can easily guess that iron(II) sulphate (light green) has been formed and metallic copper (brown-red) is deposited over iron.

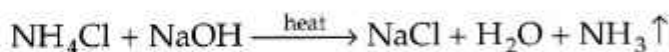
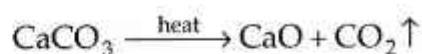


4. When treated with hydrogen sulphide gas (colourless), a solution of copper(II) sulphate (blue) gives a black precipitate of copper(II) sulphide.



The evolution of a gas

Gases evolve in many chemical reactions. Some gases have a characteristic odour, and can be easily identified. For example, sulphur dioxide has a suffocating smell, ammonia has a pungent smell and hydrogen sulphide has the smell of rotten eggs. There are chemical tests also for different gases. For example, carbon dioxide turns limewater milky, and hydrogen, when ignited, burns with a 'pop'.



When gas bubbles are formed, usually with a hissing sound, due to a reaction, the phenomenon is called effervescence.

Some examples of effervescence are given below.

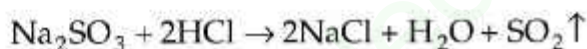


Fig. 5.11 On being heated, copper(II) carbonate (green) gives copper(II) oxide (black).

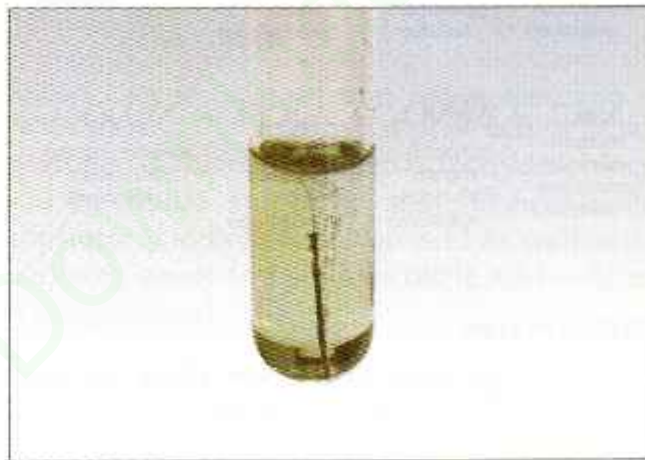


Fig. 5.12 Iron dissolves in dilute sulphuric acid to give a light green solution of iron(II) sulphate.

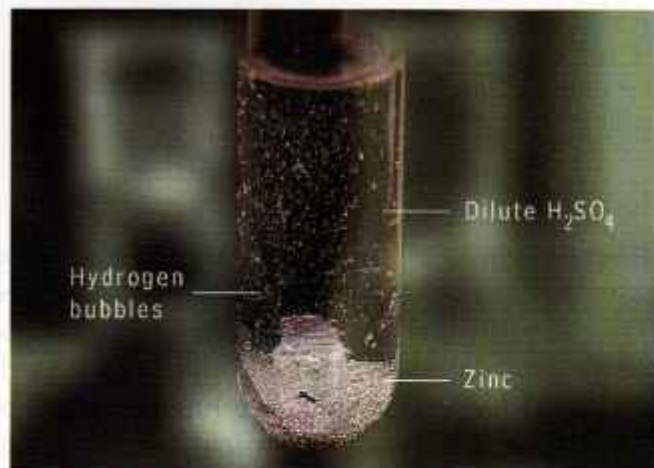
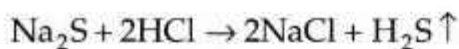
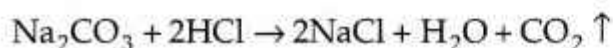
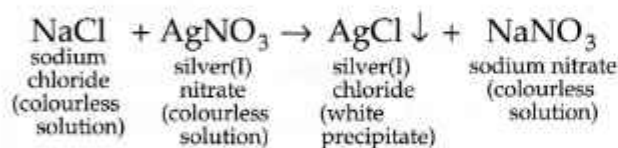
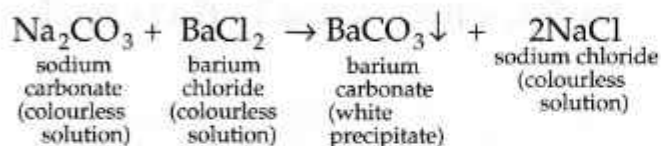


Fig. 5.13 Zinc reacts with dilute sulphuric acid and hydrogen gas is evolved.



The formation of a precipitate

The formation of a precipitate is also suggestive of a chemical reaction.

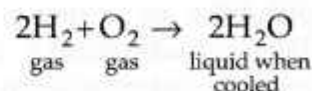


Change in state

A change in state also takes place in many reactions. For example, a gaseous mixture of hydrogen and oxygen, when ignited, gives water vapour. On being cooled, water vapour



Fig. 5.14 When an acid reacts with a carbonate or a bicarbonate, carbon dioxide is released with effervescence.



Also, ammonia gas reacts with hydrogen chloride gas to form ammonium chloride (solid).

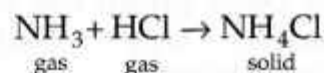


Fig. 5.15 When a glass rod dipped in NH_4OH is brought near a test tube containing HCl , NH_4Cl is formed. The solid particles of NH_4Cl give white colour to the vapours.

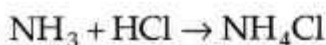
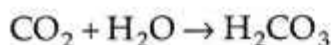
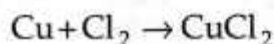
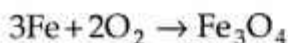
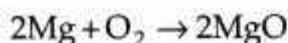
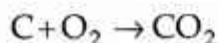
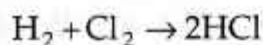
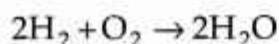
Types of Chemical Reactions

Chemical reactions are classified into several types.

Combination reactions

In a combination reaction, two or more

Some examples are given below.



Decomposition reactions

In a decomposition reaction, one substance breaks down into two or more.

Here are a few common examples.

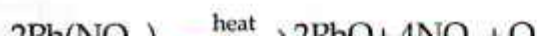
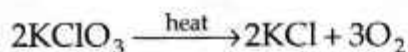
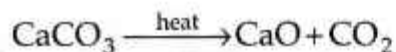
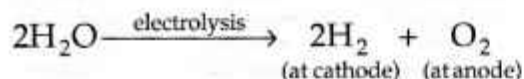


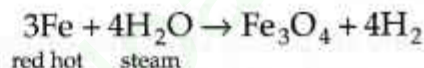
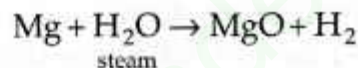
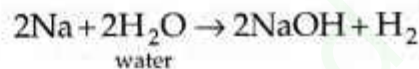
Fig. 5.16 On being heated, lead nitrate gives brown fumes of nitrogen dioxide.

Displacement reactions

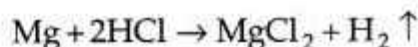
In a displacement reaction, one element displaces another from its compound and forms a new compound.

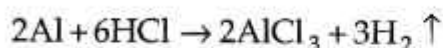
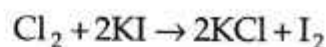
You have learnt in an earlier chapter that a more active metal displaces a less active metal or hydrogen from its compound. The following reactions fall under this category.

Displacement of hydrogen from water

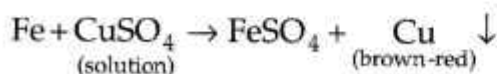
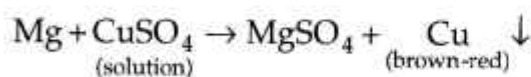
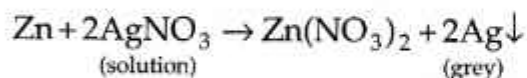


Displacement of hydrogen from hydrochloric acid or dilute sulphuric acid Metals more active than hydrogen displace hydrogen from hydrochloric acid or dilute sulphuric acid. Concentrated sulphuric acid behaves as an acid as well as an oxidising agent and so does nitric acid—dilute or concentrated.





The displacement of a metal by another metal A more active metal displaces a less active metal from a solution of its compound.

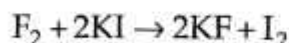
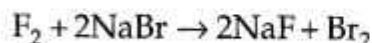
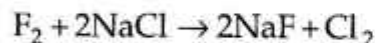


It is evident that a less active metal does not displace a more active metal from its compound. Thus, iron does not displace magnesium and aluminium, and silver does not displace zinc and iron from their compounds.

The displacement of a nonmetal by another nonmetal

Just like a metal, a nonmetal also is less active or more active than others. For example, among the chemically similar elements fluorine, chlorine, bromine and iodine—called the halogens—the activity decreases in the order $F > Cl > Br > I$. So, the following displacement reactions occur.

1. Fluorine displaces chlorine, bromine and iodine from chlorides, bromides and iodides respectively.



2. Chlorine displaces bromine and iodine from bromides and iodides respectively.



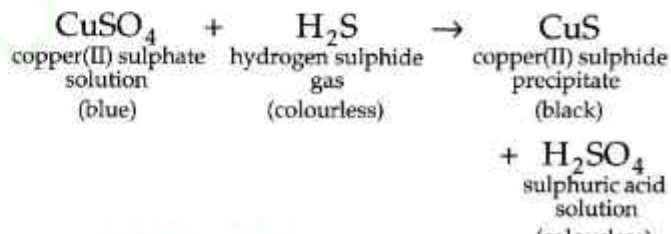
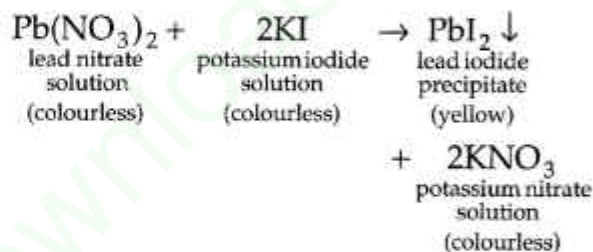
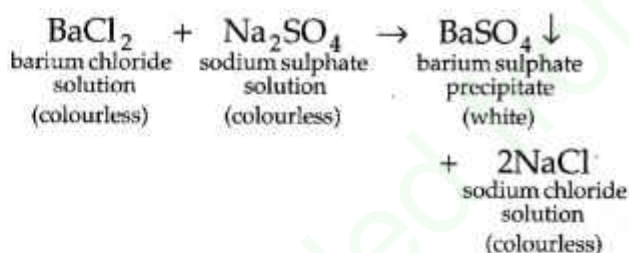
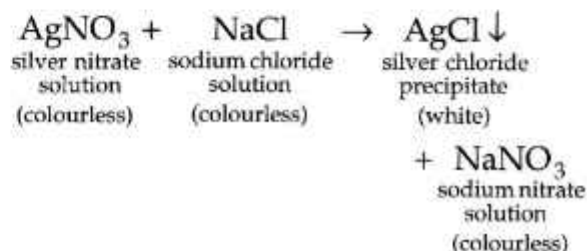
3. Bromine displaces iodine from iodides.



Double decomposition reactions

In a double decomposition reaction, the positive and negative radicals of two reactants are exchanged, leading to the precipitation of a product.

These reactions occur very fast. Some common examples are given below.



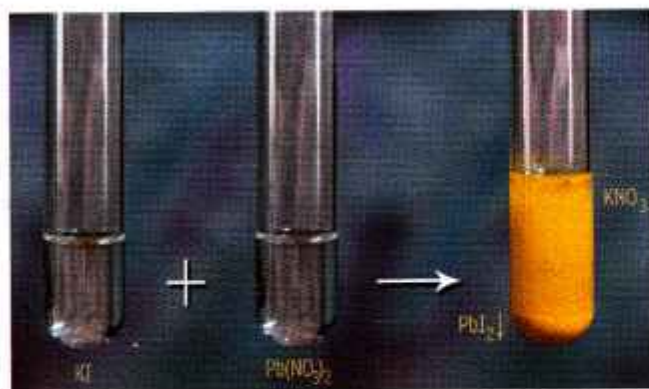
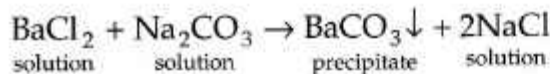


Fig. 5.17 An example of a double decomposition reaction

EXAMPLE 1 Barium carbonate (BaCO_3) is insoluble in water. What do you expect when a solution of barium chloride is mixed with one of sodium carbonate?

Solution A double decomposition reaction will take place between barium chloride and sodium carbonate. As a result, barium carbonate will precipitate.



Neutralisation reactions

You have read about acids, bases and salts in your previous class.

A reaction between an acid and a base, giving a salt and water, is called a neutralisation reaction.



Some examples are given below.

Acid	+	Base		Salt	+	Water
HCl	+	NaOH	$\rightarrow$	NaCl	+	H_2O
H_2SO_4	+	2KOH	$\rightarrow$	K_2SO_4	+	$2\text{H}_2\text{O}$
2HNO_3	+	$\text{Ca}(\text{OH})_2$	$\rightarrow$	$\text{Ca}(\text{NO}_3)_2$	+	$2\text{H}_2\text{O}$
6HCl	+	Al_2O_3	$\rightarrow$	2AlCl_3	+	$3\text{H}_2\text{O}$
2HCl	+	MgO	$\rightarrow$	MgCl_2	+	H_2O

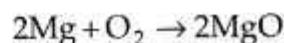
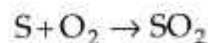
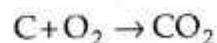
Oxidation/Reduction reactions

The addition of hydrogen to or removal of oxygen from a substance is called reduction.

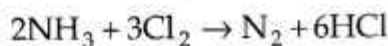
The addition of oxygen to or removal of hydrogen from a substance is called oxidation.

A substance causing reduction is called a reducing agent, and one causing oxidation is called an oxidising agent.

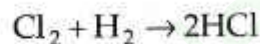
Examples of oxidation When burnt in oxygen, C is oxidised to CO_2 , S to SO_2 and Mg to MgO —these elements combine with oxygen.



NH_3 is oxidised to N_2 when it is made to react with chlorine, as it loses hydrogen.



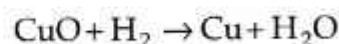
Examples of reduction When burnt in hydrogen, Cl_2 is reduced to HCl by the addition of hydrogen.



Similarly, S is reduced to H_2S when boiling sulphur combines with hydrogen.



However, CuO is reduced to Cu, by the loss of oxygen, when hydrogen is passed over hot copper(II) oxide.

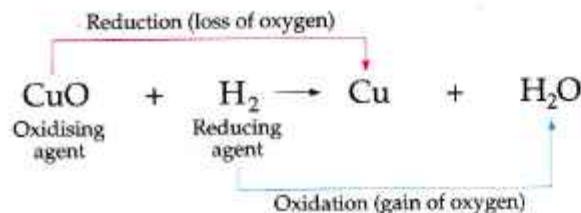


Oxidation and reduction take place simultaneously It has been found that oxidation and reduction take place simultaneously. Reactions involving them are called **redox** (reduction-oxidation) reactions.

In a redox reaction, the oxidising agent is reduced and the reducing agent is oxidised.

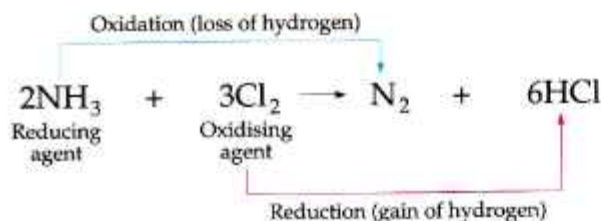
Let us examine the following reactions.

1.

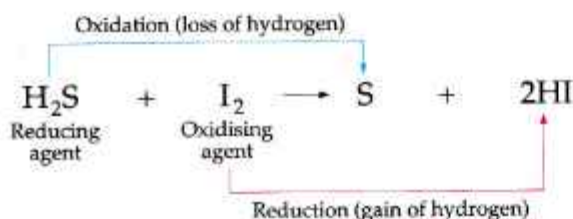


Not only is CuO reduced to Cu, but H₂ is oxidised to H₂O as well. As CuO is causing the oxidation of hydrogen, CuO is the oxidising agent in this reaction.

2.

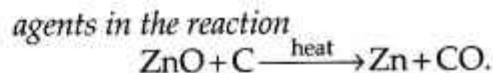
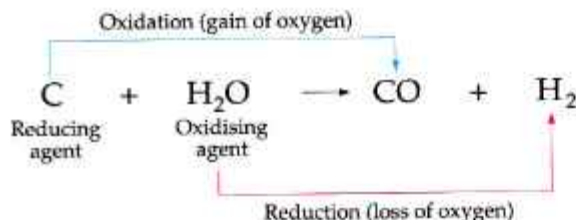


3.



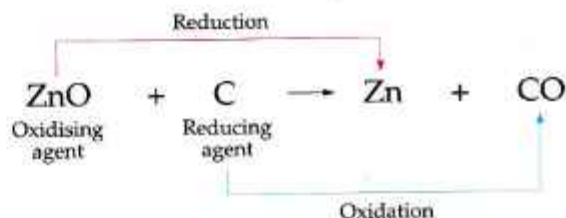
EXAMPLE 2 When steam is passed over red-hot coke, a mixture of hydrogen and carbon monoxide is formed. Schematically show that it is a redox reaction.

Solution



Solution ZnO is reduced to Zn by carbon, so carbon is the reducing agent.

C is oxidised to CO by zinc oxide, so zinc oxide is the oxidising agent.



Electrolysis

Pure water does not allow electric current to pass through it and neither do solid sodium chloride and calcium chloride. But water, when mixed with some acid, allows the passage of electric current and gets decomposed into hydrogen and oxygen. Similarly, electric current can be passed through molten sodium chloride, which then decomposes into sodium and chlorine. Electric current can also be passed through molten calcium chloride, breaking it into calcium and chlorine. This kind of decomposition, or breaking down, is called **electrolysis**.

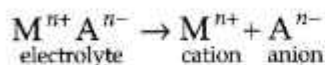
Electrolysis is a process in which a substance is decomposed, or broken down, into simpler substances, by passing electric current through it.

A substance that allows the passage of electric current in solution or in the molten state is called an **electrolyte**. All ionic substances are electrolytes.

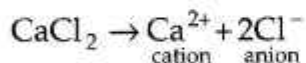
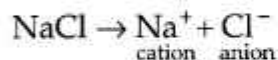
Steps in electrolysis Electrolysis may be considered to take place in the following steps.

1. Dissociation, or ionisation Ionic compounds are composed of ions. But in the solid state, these ions are held so tightly by each other that they cannot move. However, in solution or in the molten state, they can move and are said to be labile. The process is called **dissociation**. Thus, an electrolyte, $\text{M}^{n+}\text{A}^{n-}$, dissociates in solution or in the molten state into the cation M^{n+} and

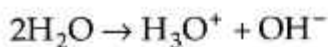
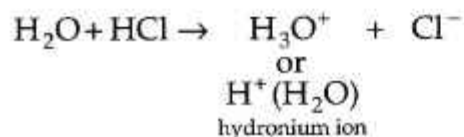
the anion A^{n-} , where $n+$ and $n-$ represent the charges on the cation and the anion respectively.



For example,



Sometimes, ions are not originally present in the substance, but they are formed by some reaction in solution. For example, water does not contain ions, but when mixed with an acid, it produces ions.

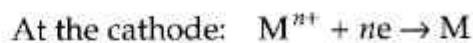


2. The movement of ions in an electric field The electrolyte (solution or melt) is poured in a vessel called an **electrolytic cell** or a **voltameter**. The cell is fitted with two electrodes connected to the two poles of a battery. The electrode connected to the negative pole is called the **cathode**, and the one connected to the positive pole, the **anode**. When the current is switched on,

- (i) the positively charged ions, i.e., the cations, move towards the negative electrode, i.e., the cathode, and
- (ii) the negatively charged ions, i.e., the anions, move towards the positive electrode, i.e., the anode.

3. Discharge of ions at the electrodes The cations take electrons from the cathode and get discharged (i.e., lose their charge). Similarly, the

anions give electrons to the anode and get discharged.



From aqueous solutions, however, highly active metals like calcium, sodium, magnesium and aluminium are not discharged at the cathode. In such cases, hydrogen is liberated at the cathode.

For all practical purposes, you can treat the ionisation of acidulated water to have taken place as follows.



The hydrogen ions are discharged at the cathode, and the hydroxide ions at the anode. The discharged hydroxide ions produce oxygen at the anode.

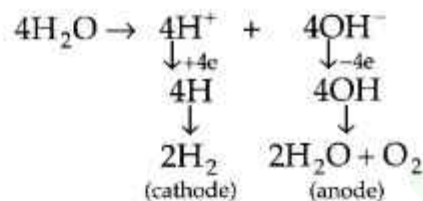


Table 5.1 Products of electrolysis

Substance	State	Discharged species	
		Cathode	Anode
Sodium chloride	molten	sodium	chlorine
Sodium chloride	aqueous solution	hydrogen	chlorine
Magnesium chloride	molten	magnesium	chlorine
Magnesium chloride	aqueous solution	hydrogen	chlorine
Aluminium oxide	solution in molten cryolite	aluminium	oxygen
Lead bromide	molten	lead	bromine
Water	acidulated	hydrogen	oxygen

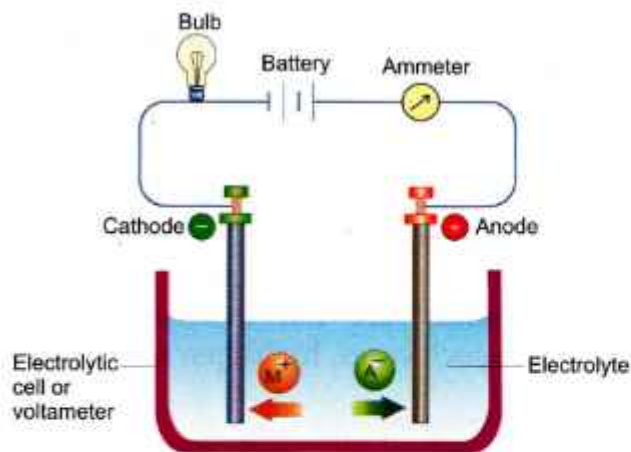


Fig. 5.18 Electrolysis

Electroplating

Depositing a layer of one metal over another by the process of electrolysis is known as electroplating.

Electroplating iron with tin, nickel or chromium protects it from rusting. Similarly, other metals also are electroplated to protect them from corrosion. Electroplating some cheaper metals or alloys with copper, silver or gold make them look like these metals. Things electroplated with copper, silver or gold are used for making cutlery or artificial jewellery. Bright cycle parts and bathroom fittings are usually made of alloys electroplated with nickel or chromium.

If something is to be electroplated, it is made the cathode, and the metal with which it is to be electroplated is made the anode. Suppose you want the electroplating to be done with a certain metal. Then a salt solution of this metal is used as the electrolyte. Upon electrolysis, the metal is deposited on the cathode, i.e., the article to be electroplated. The anions draw out metal ions from the anode, and these dissolve. This makes up for the loss of the cations.

An electrolyte should be such that the metal deposits slowly and uniformly. A solution of blue vitriol ($\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$), acidified with sulphuric acid, is used for copper plating. Similarly, a solution of nickel ammonium

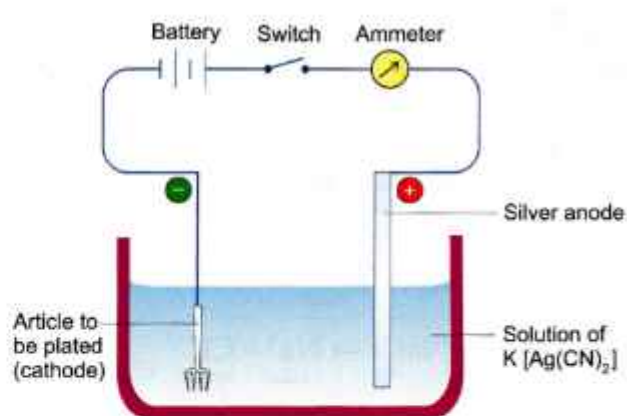


Fig. 5.19 Electroplating with silver

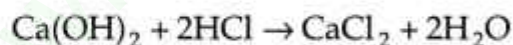
with sulphuric acid, is used for nickel plating. For silver or gold plating, a complicated salt containing cyanide ions (CN^-) is used.

What Balanced Chemical Equations Convey

A balanced chemical equation conveys the following.

1. What the reactants and products are
2. In what molecular proportion the reactants react and the products are formed
3. In what proportion by mass the reactants react and the products are formed

The reactants and products, as well as their molecular proportions, are obvious from a balanced chemical equation. For example, the equation



conveys that one molecule of calcium hydroxide reacts with two molecules of hydrogen chloride to form one molecule of calcium chloride and two molecules of water.

For understanding mass relations in chemical reactions, we will have to learn about

Molecular Mass

The molecular mass of an element or a compound is the number of times a molecule of the element or compound is heavier than an atom of hydrogen.

It is obtained by adding the atomic mass of each atom contained in a molecule of the element or compound.

For example, the molecular mass of

hydrogen (H_2) = $1 + 1 = 2$ amu,

nitrogen (N_2) = $14 + 14 = 28$ amu,

oxygen (O_2) = $16 + 16 = 32$ amu, and

chlorine (Cl_2) = $35.5 + 35.5 = 71$ amu.

Similarly, the molecular mass of

hydrogen chloride (HCl) = $1 + 35.5$
= 36.5 amu,

water (H_2O) = $2 \times 1 + 16 = 18$ amu,

ammonia (NH_3) = $14 + 3 \times 1 = 17$ amu,
and

sodium hydroxide (NaOH) = $23 + 16 + 1$
= 40 amu.

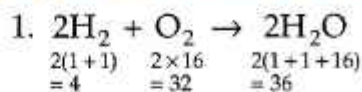
Generally, the term **formula mass** is preferred to molecular mass for ionic compounds. For example, sodium chloride (NaCl) has the formula mass of $(23 + 35.5)$ amu = 58.5 amu.

When we express these masses by the term **relative molecular mass** or **formula mass**, we simply drop the unit amu.

The Proportion of Mass Conveyed by Chemical Equations

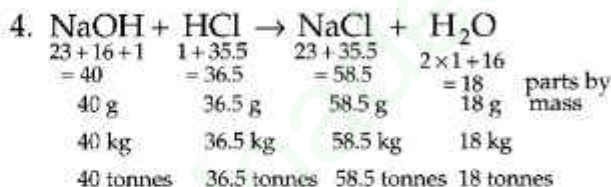
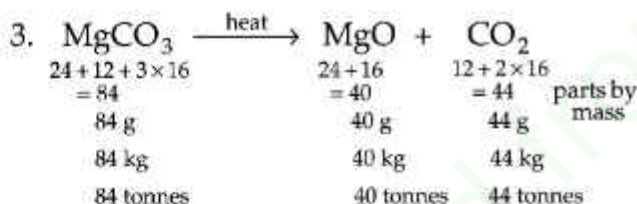
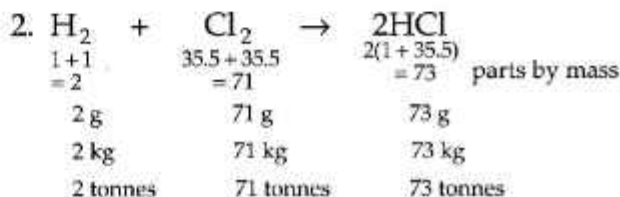
The proportion of masses of the reactants and products can be calculated from a balanced chemical equation in the manner shown below. Remember that it is only the proportion and not the actual masses of the reactants and products. We can take the actual masses in grams, kilograms, tonnes or any unit we like—the proportion will remain the same in a reaction.

Examples of calculation of mass



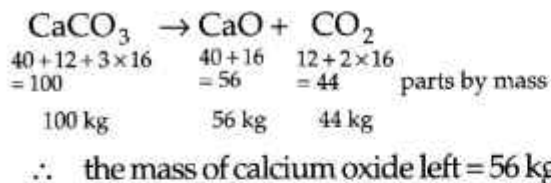
Thus, 4 parts by mass of hydrogen will react with 32 parts by mass of oxygen to give 36 parts by mass of water. In other words, 1 part by mass of hydrogen will react with 8 parts by mass of oxygen to give 9 parts by mass of water. Choosing different units,

hydrogen	+	oxygen	→	water
1 g		8 g		9 g
1 kg		8 kg		9 kg
1 tonne		8 tonnes		9 tonnes



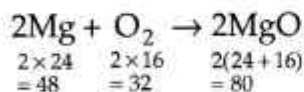
EXAMPLE 4 100 kg of limestone is so strongly heated that all the carbon dioxide formed is driven off. What is the mass of calcium oxide left as the residue? (Relative atomic masses: Ca = 40, C = 12, O = 16)

Solution The balanced chemical equation for the reaction can be expressed as follows.



EXAMPLE 5 Calculate the mass of magnesium oxide formed when 12 g of magnesium is heated in oxygen.

Solution The balanced chemical equation for the reaction is



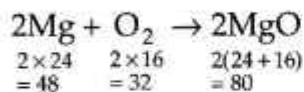
48 g of Mg gives 80 g of MgO.

$$\therefore 12 \text{ g of Mg gives } \left(\frac{80}{48} \times 12 \right) = 20 \text{ g of MgO.}$$

So, the mass of magnesium oxide formed = 20 g.

EXAMPLE 6 What will be the increase in mass if 120 g of magnesium is heated in oxygen?

Solution The balanced chemical equation for the reaction is



From the equation, it is evident that, on reaction with oxygen, 48 g of Mg will form 80 g of MgO.

Thus, the increase in mass for 48 g of Mg is 80 g – 48 g = 32 g.

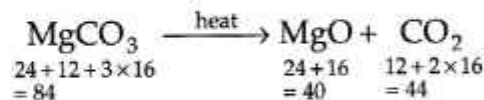
$\therefore$ the increase in mass for 120 g of Mg

$$\text{is } \left(\frac{32}{48} \times 120 \right) \text{ g} = 80 \text{ g.}$$

So, the increase in mass = 80 g.

EXAMPLE 7 What will be the loss in mass when 42 g of magnesium carbonate is strongly heated to give magnesium oxide? (Relative atomic masses: Mg = 24, C = 12, O = 16)

Solution The balanced chemical equation for the reaction can be expressed as follows.



When strongly heated,

84 g of MgCO_3 will lose 44 g of CO_2 .

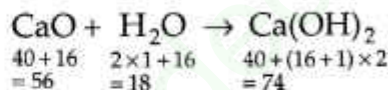
$\therefore$ 42 g of MgCO_3 will lose

$$\left(\frac{44}{84} \times 42 \right) \text{ g} = 22 \text{ g of CO}_2.$$

So, the loss in mass = 22 g.

EXAMPLE 8 How much slaked lime $[\text{Ca}(\text{OH})_2]$ will be obtained by the action of water on 168 tonnes of quicklime (CaO)? (Relative atomic masses: Ca = 40, O = 16, H = 1)

Solution The balanced chemical equation for the reaction is



When treated with water,

56 tonnes of quicklime will produce 74 tonnes of slaked lime.

$\therefore$ 168 tonnes of quicklime will produce $\left(\frac{74}{56} \times 168 \right)$ tonnes

= 222 tonnes of slaked lime.

So, the mass of slaked lime obtained = 222 tonnes.

Points to Remember

- A change in state is accompanied by an energy change.
- Changes in state of a substance occur at fixed temperatures.
- The melting point of a solid is lowered and the boiling point of a liquid is raised if they contain soluble impurities.
- Chemical reactions are accompanied by energy change.
- A reaction in which heat is absorbed is called an *endothermic reaction*.
- A reaction in which heat is evolved (i.e., given out) is called an *exothermic reaction*.
- A *catalyst* is a substance that changes the speed (or rate) of a reaction but itself remains unchanged.
- A chemical reaction is characterised by one or more of the following.
 - (i) A change in colour
 - (ii) The evolution of a gas
 - (iii) The formation of a precipitate
 - (iv) A change in state
- In a *combination reaction*, two or more reactants add up to form a product.
- In a *decomposition reaction*, one substance breaks down into two or more.
- In a *displacement reaction*, one element displaces another from its compound and forms a new compound.
- In a *double decomposition reaction*, the positive and negative radicals of two reactants are exchanged, leading to the precipitation of a product.
- A reaction between an acid and a base, giving a salt and water, is called a *neutralisation reaction*.
- The addition of hydrogen to or the removal of oxygen from a substance is called *reduction*.
- The addition of oxygen to or the removal of hydrogen from a substance is called *oxidation*.
- A substance causing reduction is called a *reducing agent*, and one causing oxidation is called an *oxidising agent*.
- Oxidation and reduction take place simultaneously, and these reactions are called *redox reactions*.
- *Electrolysis* is a process in which a substance is decomposed, or broken down, into simpler substances by passing electric current through it.
- Depositing a layer of one metal over another by the process of electrolysis is called *electroplating*.
- The *molecular mass* of an element or a compound is the number of times a molecule of the element or compound is heavier than an atom of hydrogen.

Exercise

Short-Answer Questions

1. Why do water droplets deposit on the outer walls of a glass when we place some ice in it?
2. Will the boiling point of a liquid change if the atmospheric pressure changes? Will it increase or decrease with pressure?
3. Why is the boiling point of tap water higher than that of distilled water?
4. Why are chemical reactions accompanied by energy change?
5. What are exothermic reactions? Give an example.
6. What are endothermic reactions? Give an example.

7. What is that substance called which changes the speed of a reaction, but itself remains unchanged? Give an example of such a substance.
8. What is effervescence? Give an example.
9. What are the following reactions called?
 - (a) A reaction between an acid and a base, giving a salt and water
 - (b) The addition of hydrogen to or the removal of oxygen from a substance
 - (c) The addition of oxygen to or the removal of hydrogen from a substance
10. Identify the oxidising and reducing agents in each of the following reactions.
 - (a) $2\text{NH}_3 + 3\text{Cl}_2 \rightarrow \text{N}_2 + 6\text{HCl}$
 - (b) $\text{PbO} + \text{H}_2 \rightarrow \text{Pb} + \text{H}_2\text{O}$
 - (c) $\text{H}_2\text{S} + \text{Br}_2 \rightarrow \text{S} + 2\text{HBr}$
 - (d) $\text{SnO}_2 + 2\text{C} \rightarrow \text{Sn} + 2\text{CO}$
11. What is electrolysis?
12. Give the chemical equation for the reactions taking place at the cathode and anode during the electrolysis of molten sodium chloride.
13. What is electroplating?
14. The handle of a cycle is to be electroplated with nickel. What are the substances that will be used for (i) the cathode, (ii) the anode, and (iii) the electrolyte?
15. Give balanced chemical equations for the following reactions.
 - (a) On being heated, copper(II) carbonate gives copper(II) oxide and carbon dioxide.
 - (b) Iron dissolves in dilute sulphuric acid.
 - (c) Iron displaces copper from copper(II) sulphate solution.
 - (d) On being heated, sodium hydrogencarbonate, gives sodium carbonate, water vapour and carbon dioxide.

Long-Answer Questions

1. Describe how you would determine the melting point of a solid other than ice.
2. Describe how you would determine the boiling point of a liquid.
3. How would you prove that tap water contains dissolved salts?
4. Describe, with a neat diagram, the process of distilling a liquid.

Objective Questions

Choose the correct option.

1. The melting point of ice, mixed with common salt, is
 - (a) below 0°C
 - (b) 0°C
 - (c) slightly above 0°C
 - (d) much above 0°C
2. A mixture of liquids is subjected to fractional distillation. In the fractionating column, the vapours of
 - (a) the less volatile liquid move up and those of the more volatile liquid move down
 - (b) the more volatile liquid move up and those of the less volatile liquid move down
 - (c) all liquids move up
 - (d) all liquids move down
3. Which of the following catalysts is used for manufacturing hardened fats?
 - (a) Nickel
 - (b) Iron
 - (c) Platinum
 - (d) None of these
4. Which of the following is a neutralisation reaction?
 - (a) $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$
 - (b) $\text{Fe} + \text{S} \rightarrow \text{FeS}$
 - (c) $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
 - (d) $\text{H}_2\text{SO}_4 + 2\text{KOH} \rightarrow \text{K}_2\text{SO}_4 + 2\text{H}_2\text{O}$

5. Which of the following reactions is responsible for the liberation of oxygen at the anode in the electrolysis of water acidulated with hydrochloric acid?
- (a) $4\text{OH}^- \rightarrow 2\text{H}_2\text{O} + \text{O}_2 + 4\text{e}^-$ (b) $2\text{O}^{2-} \rightarrow \text{O}_2 + 4\text{e}^-$
 (c) $2\text{Cl}_2 + 2\text{H}_2\text{O} \rightarrow 4\text{HCl} + \text{O}_2$ (d) None of these
6. On electrolysis, which of the following electrolytes will give hydrogen at the cathode?
- (a) Sodium chloride solution (b) Molten calcium chloride
 (c) Molten magnesium chloride (d) Aluminium oxide dissolved in molten cryolite

Fill in the blanks.

- The evaporation of a liquid is accompanied by the _____ of the surface from which it evaporates. (heating/cooling)
- The melting point of a solid is _____ the freezing point of the corresponding liquid. (the same as/different from)
- Enzymes act as _____ in living systems. (oxidising agents/reducing agents/catalysts)
- Pure water _____ be electrolysed, but acidulated water _____. (can/cannot)
- Complete and balance the following chemical equations.
 - $\text{Na} + \text{_____} \rightarrow \text{NaOH} + \text{_____}$
 - $\text{CuSO}_4 + \text{H}_2\text{S} \rightarrow \text{_____} + \text{H}_2\text{SO}_4$
 - $\text{Na}_2\text{S} + \text{HCl} \rightarrow \text{NaCl} + \text{_____}$
 - $\text{ZnSO}_3 \xrightarrow{\text{heat}} \text{_____} + \text{SO}_2$
 - $\text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} \downarrow + \text{_____}$
 - $\text{NaOH} + \text{H}_2\text{SO}_4 \rightarrow \text{_____} + \text{_____}$

Match the columns.

- Match the catalysts with their functions.

Catalyst

- Iron
- Carboxypeptidase
- Catalase
- Nitrogenase

Function

- Nitrogen fixation in plants
- Decomposition of hydrogen peroxide in our body
- Manufacture of ammonia
- Digestion of food

- Match the colours of the compounds mentioned below.

Compound

- Iron(II) sulphate
- Copper(II) sulphate
- Copper(II) oxide
- Silver chloride

Colour

- White
- Black
- Light green
- Blue

- Match the reactions with their types.

Reaction

- $\text{Fe} + \text{S} \rightarrow \text{FeS}$
- $2\text{Pb}(\text{NO}_3)_2 \rightarrow 2\text{PbO} + 4\text{NO}_2 + \text{O}_2$
- $\text{Cl}_2 + 2\text{KI} \rightarrow 2\text{KCl} + \text{I}_2$
- $\text{Pb}(\text{NO}_3)_2 + 2\text{KI} \rightarrow \text{PbI}_2 \downarrow + 2\text{KNO}_3$

Type

- Decomposition
- Double decomposition
- Combination
- Displacement

Indicate which of the following statements are true and which are false.

- Pure substances have fixed melting and boiling points.

3. Change in colour does not characterise a chemical reaction.
4. Ionic compounds conduct electricity in the solid state.
5. In electroplating, the anode must be made of the same metal with which the article is to be electroplated.

Numerical Problems

[Relative atomic masses: H = 1, C = 12, O = 16, S = 32, Mg = 24, Ca = 40, Fe = 56]

1. How much magnesium carbonate must be heated to obtain 60 tonnes of magnesium oxide?
(Ans. 126 tonnes)
2. On being heated at 120°C , gypsum ($\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$), gives plaster of Paris. How much pure gypsum must be heated to obtain 14.5 kg of plaster of Paris?
(Ans. 17.2 kg)
3. Iron(II) sulphide reacts with dilute sulphuric acid to give a solution of iron(II) sulphate. On crystallisation, the solution yields crystals of $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$. How many grams of these crystals will be obtained if we start with 88 g of iron(II) sulphide?
(Ans. 278 g)
4. How much magnesium must be burnt in oxygen to obtain 1 kg of magnesium oxide? (Ans. 600 g or 0.6 kg)
5. What is the volume of oxygen that will be used by 12 g of magnesium to be completely converted into magnesium oxide? Assume that the volume of 32 g of oxygen is 22.4 L.
(Ans. 5.6 L)



A Simple Experiment on Electrolysis

You can perform an interesting experiment on electrolysis at home.

Remove the plastic insulation from both ends of two copper wires. The naked ends are a shining brown-red. Use a naked end of one wire as the anode and that of the other as the cathode. Introduce the two electrodes into a glass containing distilled water. You can buy distilled water from a chemist's. Connect the other ends of the two wires respectively to the positive and negative terminals of a 3-volt battery. You will find that no change takes place. This is because distilled water does not allow an electric current to pass through it.

Now, pour a few drops of hydrochloric acid (bathroom acid) into the distilled water and make the connections as before. A gas begins to evolve at the cathode (negative electrode). The evolution becomes vigorous within a short while. The evolution of gas at the anode (positive electrode) is relatively slow. What is happening is the electrolysis of acidulated water.

Stop the electrolysis after some time. Take out the electrodes and examine them. You will find that the anode has become blackish. Why? On electrolysis, acidulated water gives hydrogen at the cathode and oxygen at the anode. The oxygen reacts with the copper (of the anode) to form copper(II) oxide, which is black.

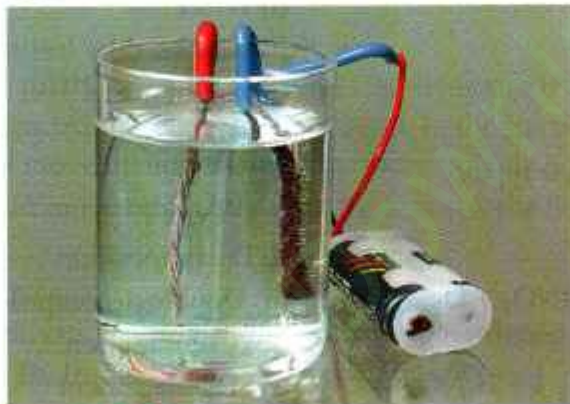


Fig. 5.20 Electrolysis of acidulated water