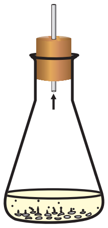
"Facts are not science — as the dictionary is not literature."
Martin H. Fischer



CHAPTER 1

Chemical Reactions and Equations

1064CH01

Consider the following situations of daily life and think what happens when –

- milk is left at room temperature during summers.
- an iron tawa/pan/nail is left exposed to humid atmosphere.
- grapes get fermented.
- food is cooked.
- food gets digested in our body.
- we respire.

In all the above situations, the nature and the identity of the initial substance have somewhat changed. We have already learnt about physical and chemical changes of matter in our previous classes. Whenever a chemical change occurs, we can say that a chemical reaction has taken place.

You may perhaps be wondering as to what is actually meant by a chemical reaction. How do we come to know that a chemical reaction has taken place? Let us perform some activities to find the answer to these questions.

Activity 1.1

CAUTION: This Activity needs the teacher's assistance. It would be better if students wear suitable eyeglasses.

- Clean a magnesium ribbon about 3-4 cm long by rubbing it with sandpaper.
- Hold it with a pair of tongs. Burn it using a spirit lamp or burner and collect the ash so formed in a watch-glass as shown in Fig. 1.1. Burn the magnesium ribbon keeping it away as far as possible from your eyes.
- What do you observe?

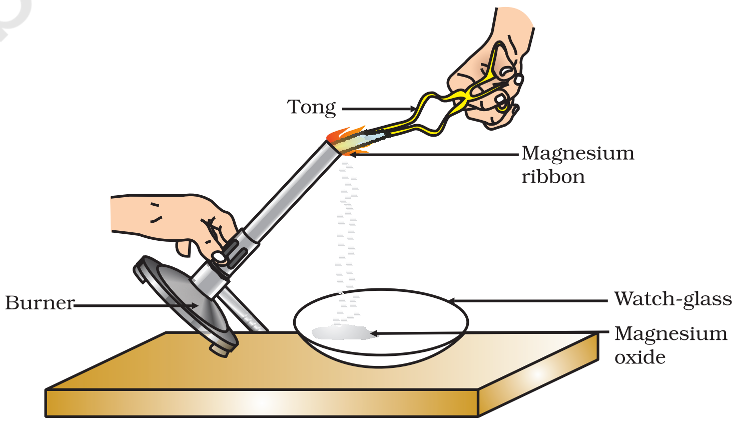


Figure 1.1
Burning of a magnesium ribbon in air and collection of magnesium oxide in a watch-glass

You must have observed that magnesium ribbon burns with a dazzling white flame and changes into a white powder. This powder is magnesium oxide. It is formed due to the reaction between magnesium and oxygen present in the air.

Activity 1.2

- Take lead nitrate solution in a test tube.
- Add potassium iodide solution to this.
- What do you observe?

Activity 1.3

- Take a few zinc granules in a conical flask or a test tube.
- Add dilute hydrochloric acid or sulphuric acid to this (Fig. 1.2).
  CAUTION: Handle the acid with care.
- Do you observe anything happening around the zinc granules?
- Touch the conical flask or test tube. Is there any change in its temperature?

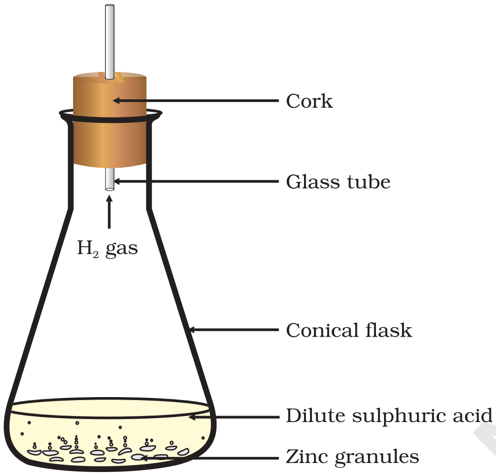


Figure 1.2
Formation of hydrogen gas by the action of dilute sulphuric acid on zinc

From the above three activities, we can say that any of the following observations helps us to determine whether a chemical reaction has taken place –

- change in state
- change in colour
- evolution of a gas
- change in temperature.

As we observe the changes around us, we can see that there is a large variety of chemical reactions taking place around us. We will study about the various types of chemical reactions and their symbolic representation in this Chapter.

1.1 CHEMICAL EQUATIONS

Activity 1.1 can be described as – when a magnesium ribbon is burnt in oxygen, it gets converted to magnesium oxide. This description of a chemical reaction in a sentence form is quite long. It can be written in a shorter form. The simplest way to do this is to write it in the form of a word-equation.
The word-equation for the above reaction would be –

$$\underset{\text{(Reactants)}}{\text{Magnesium + Oxygen}} \rightarrow \underset{\text{(Product)}}{\text{Magnesium oxide}} \qquad (1.1)$$

The substances that undergo chemical change in the reaction (1.1), magnesium and oxygen, are the reactants. The new substance is magnesium oxide, formed during the reaction, as a product.

A word-equation shows change of reactants to products through an arrow placed between them. The reactants are written on the left-hand side (LHS) with a plus sign (+) between them. Similarly, products are written on the right-hand side (RHS) with a plus sign (+) between them. The arrowhead points towards the products, and shows the direction of the reaction.

1.1.1 Writing a Chemical Equation

Is there any other shorter way for representing chemical equations? Chemical equations can be made more concise and useful if we use chemical formulae instead of words. A chemical equation represents a chemical reaction. If you recall formulae of magnesium, oxygen and magnesium oxide, the above word-equation can be written as –

$$Mg + O_2 \rightarrow MgO \qquad (1.2)$$

Count and compare the number of atoms of each element on the LHS and RHS of the arrow. Is the number of atoms of each element the same on both the sides? If yes, then the equation is balanced. If not, then the equation is unbalanced because the mass is not the same on both sides of the equation. Such a chemical equation is a skeletal chemical equation for a reaction. Equation (1.2) is a skeletal chemical equation for the burning of magnesium in air.

1.1.2 Balanced Chemical Equations

Recall the law of conservation of mass that you studied in Class IX; mass can neither be created nor destroyed in a chemical reaction. That is, the total mass of the elements present in the products of a chemical reaction has to be equal to the total mass of the elements present in the reactants.

In other words, the number of atoms of each element remains the same, before and after a chemical reaction. Hence, we need to balance a skeletal chemical equation. Is the chemical Eq. (1.2) balanced? Let us learn about balancing a chemical equation step by step.

The word-equation for Activity 1.3 may be represented as –

Zinc + Sulphuric acid → Zinc sulphate + Hydrogen

The above word-equation may be represented by the following chemical equation –

$$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2 \qquad (1.3)$$

Let us examine the number of atoms of different elements on both sides of the arrow.

Element	Number of atoms in reactants (LHS)	Number of atoms in products (RHS)
Zn	1	1
H	2	2
S	1	1
O	4	4

As the number of atoms of each element is the same on both sides of the arrow, Eq. (1.3) is a balanced chemical equation.

Let us try to balance the following chemical equation –

$$Fe + H_2O \rightarrow Fe_3O_4 + H_2 \qquad (1.4)$$

Rationalised 2023-24

Step I: To balance a chemical equation, first draw boxes around each formula. Do not change anything inside the boxes while balancing the equation.

$$\boxed{Fe} + \boxed{H_2O} \rightarrow \boxed{Fe_3O_4} + \boxed{H_2} \quad (1.5)$$

Step II: List the number of atoms of different elements present in the unbalanced equation (1.5).

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

Step III: It is often convenient to start balancing with the compound that contains the maximum number of atoms. It may be a reactant or a product. In that compound, select the element which has the maximum number of atoms. Using these criteria, we select Fe_3O_4 and the element oxygen in it. There are four oxygen atoms on the RHS and only one on the LHS.

To balance the oxygen atoms –

Atoms of oxygen	In reactants	In products
(i) Initial	1 (in H_2O)	4 (in Fe_3O_4)
(ii) To balance	1×4	4

To equalise the number of atoms, it must be remembered that we cannot alter the formulae of the compounds or elements involved in the reactions. For example, to balance oxygen atoms we can put coefficient '4' as 4 H_2O and not H_2O_4 or $(H_2O)_4$. Now the partly balanced equation becomes –

$$\boxed{Fe} + 4\boxed{H_2O} \rightarrow \boxed{Fe_3O_4} + \boxed{H_2} \quad (1.6)$$

(partly balanced equation)

Step IV: Fe and H atoms are still not balanced. Pick any of these elements to proceed further. Let us balance hydrogen atoms in the partly balanced equation.

To equalise the number of H atoms, make the number of molecules of hydrogen as four on the RHS.

Atoms of hydrogen	In reactants	In products
(i) Initial	8 (in 4 H_2O)	2 (in H_2)
(ii) To balance	8	2 × 4

The equation would be –

$$\boxed{Fe} + 4\,\boxed{H_2O} \rightarrow \boxed{Fe_3O_4} + 4\boxed{H_2} \quad (1.7)$$

(partly balanced equation)

Step V: Examine the above equation and pick up the third element which is not balanced. You find that only one element is left to be balanced, that is, iron.

Atoms of iron	In reactants	In products
(i) Initial	1 (in Fe)	3 (in Fe_3O_4)
(ii) To balance	1×3	3

To equalise Fe, we take three atoms of Fe on the LHS.

$$3\ \boxed{Fe} + 4\ \boxed{H_2O} \rightarrow \boxed{Fe_3O_4} + 4\ \boxed{H_2} \qquad (1.8)$$

Step VI: Finally, to check the correctness of the balanced equation, we count atoms of each element on both sides of the equation.

$$3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2 \qquad (1.9)$$
(balanced equation)

The numbers of atoms of elements on both sides of Eq. (1.9) are equal. This equation is now balanced. This method of balancing chemical equations is called hit-and-trial method as we make trials to balance the equation by using the smallest whole number coefficient.

Step VII: ***Writing Symbols of Physical States*** Carefully examine the above balanced Eq. (1.9). Does this equation tell us anything about the physical state of each reactant and product? No information has been given in this equation about their physical states.

To make a chemical equation more informative, the physical states of the reactants and products are mentioned along with their chemical formulae. The gaseous, liquid, aqueous and solid states of reactants and products are represented by the notations (g), (l), (aq) and (s), respectively. The word aqueous (aq) is written if the reactant or product is present as a solution in water.

The balanced Eq. (1.9) becomes

$$3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g) \qquad (1.10)$$

Note that the symbol (g) is used with H_2O to indicate that in this reaction water is used in the form of steam.

Usually physical states are not included in a chemical equation unless it is necessary to specify them.

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

$$CO(g) + 2H_2(g) \xrightarrow{340\,atm} CH_3OH(l) \qquad (1.11)$$

$$6CO_2(aq) + 12H_2O(l) \xrightarrow[Chlorophyll]{Sunlight} \underset{(Glucose)}{C_6H_{12}O_6(aq)} + 6O_2(aq) + 6H_2O(l) \qquad (1.12)$$

Using these steps, can you balance Eq. (1.2) given in the text earlier?

QUESTIONS

1. Why should a magnesium ribbon be cleaned before burning in air?
2. Write the balanced equation for the following chemical reactions.
   (i) Hydrogen + Chlorine → Hydrogen chloride
   (ii) Barium chloride + Aluminium sulphate → Barium sulphate + Aluminium chloride
   (iii) Sodium + Water → Sodium hydroxide + Hydrogen
3. Write a balanced chemical equation with state symbols for the following reactions.
   (i) Solutions of barium chloride and sodium sulphate in water react to give insoluble barium sulphate and the solution of sodium chloride.
   (ii) Sodium hydroxide solution (in water) reacts with hydrochloric acid solution (in water) to produce sodium chloride solution and water.

1.2 TYPES OF CHEMICAL REACTIONS

We have learnt in Class IX that during a chemical reaction atoms of one element do not change into those of another element. Nor do atoms disappear from the mixture or appear from elsewhere. Actually, chemical reactions involve the breaking and making of bonds between atoms to produce new substances. You will study about types of bonds formed between atoms in Chapters 3 and 4.

1.2.1 Combination Reaction

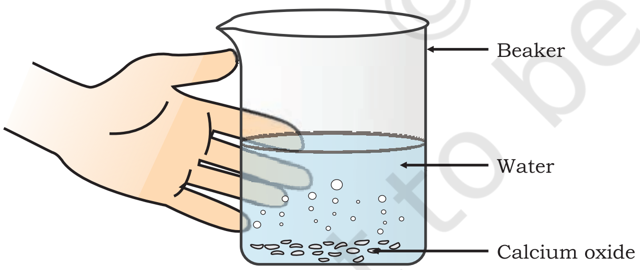


Figure 1.3
Formation of slaked lime by the reaction of calcium oxide with water

Activity 1.4

- Take a small amount of calcium oxide or quick lime in a beaker.
- Slowly add water to this.
- Touch the beaker as shown in Fig. 1.3.
- Do you feel any change in temperature?

Calcium oxide reacts vigorously with water to produce slaked lime (calcium hydroxide) releasing a large amount of heat.

$$\underset{\text{(Quick lime)}}{CaO(s)} + H_2O(l) \rightarrow \underset{\text{(Slaked lime)}}{Ca(OH)_2(aq)} + \text{Heat} \qquad (1.13)$$

In this reaction, calcium oxide and water combine to form a single product, calcium hydroxide. Such a reaction in which a single product is formed from two or more reactants is known as a combination reaction.

Do You Know?

A solution of slaked lime produced by the reaction 1.13 is used for whitewashing walls. Calcium hydroxide reacts slowly with the carbon dioxide in air to form a thin layer of calcium carbonate on the walls. Calcium carbonate is formed after two to three days of whitewashing and gives a shiny finish to the walls. It is interesting to note that the chemical formula for marble is also $CaCO_3$.

$$\underset{\text{(Calcium hydroxide)}}{Ca(OH)_2(aq)} + CO_2(g) \rightarrow \underset{\text{(Calcium carbonate)}}{CaCO_3(s)} + H_2O(l) \qquad (1.14)$$

Let us discuss some more examples of combination reactions.

(i) Burning of coal

$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad (1.15)$$

(ii) Formation of water from $H_2(g)$ and $O_2(g)$

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l) \qquad (1.16)$$

In simple language we can say that when two or more substances (elements or compounds) combine to form a single product, the reactions are called combination reactions.

In Activity 1.4, we also observed that a large amount of heat is evolved. This makes the reaction mixture warm. Reactions in which heat is released along with the formation of products are called exothermic chemical reactions.

Other examples of exothermic reactions are –

(i) Burning of natural gas

$$CH_4(g) + 2O_2\ (g) \rightarrow CO_2\ (g) + 2H_2O\ (g) \qquad (1.17)$$

(ii) Do you know that respiration is an exothermic process?

We all know that we need energy to stay alive. We get this energy from the food we eat. During digestion, food is broken down into simpler substances. For example, rice, potatoes and bread contain carbohydrates. These carbohydrates are broken down to form glucose. This glucose combines with oxygen in the cells of our body and provides energy. The special name of this reaction is respiration, the process of which you will study in Chapter 6.

$$\underset{\text{(Glucose)}}{C_6H_{12}O_6(aq)} + 6O_2(aq) \rightarrow 6CO_2(aq) + 6H_2O(l) + \text{energy} \qquad (1.18)$$

(iii) The decomposition of vegetable matter into compost is also an example of an exothermic reaction.

Identify the type of the reaction taking place in Activity 1.1, where heat is given out along with the formation of a single product.

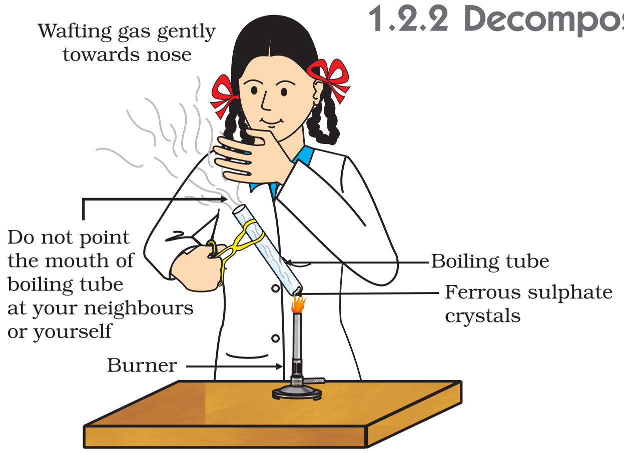


Figure 1.4
Correct way of heating the boiling tube containing crystals of ferrous sulphate and of smelling the odour

1.2.2 Decomposition Reaction

Activity 1.5

- Take about 2 g ferrous sulphate crystals in a dry boiling tube.
- Note the colour of the ferrous sulphate crystals.
- Heat the boiling tube over the flame of a burner or spirit lamp as shown in Fig. 1.4.
- Observe the colour of the crystals after heating.

Have you noticed that the green colour of the ferrous sulphate crystals has changed? You can also smell the characteristic odour of burning sulphur.

$$\underset{\text{(Ferrous sulphate)}}{2FeSO_4(s)} \xrightarrow{\text{Heat}} \underset{\text{(Ferric oxide)}}{Fe_2O_3(s)} + SO_2(g) + SO_3(g) \qquad (1.19)$$

In this reaction you can observe that a single reactant breaks down to give simpler products. This is a decomposition reaction. Ferrous sulphate crystals ($FeSO_4$, $7H_2O$) lose water when heated and the colour of the crystals changes. It then decomposes to ferric oxide (Fe_2O_3), sulphur dioxide (SO_2) and sulphur trioxide (SO_3). Ferric oxide is a solid, while SO_2 and SO_3 are gases.

Decomposition of calcium carbonate to calcium oxide and carbon dioxide on heating is an important decomposition reaction used in various industries. Calcium oxide is called lime or quick lime. It has many uses – one is in the manufacture of cement. When a decomposition reaction is carried out by heating, it is called thermal decomposition.

$$\underset{\text{(Limestone)}}{CaCO_3(s)} \xrightarrow{\text{Heat}} \underset{\text{(Quick lime)}}{CaO(s)} + CO_2(g) \qquad (1.20)$$

Another example of a thermal decomposition reaction is given in Activity 1.6.

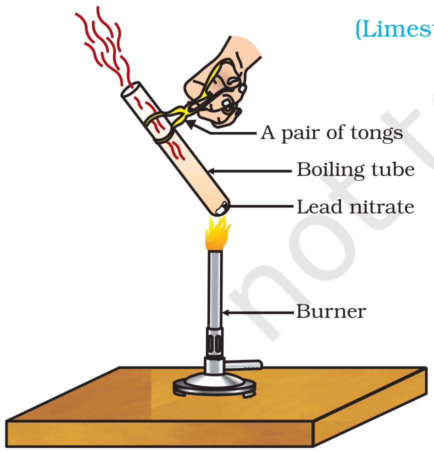


Figure 1.5
Heating of lead nitrate and emission of nitrogen dioxide

Activity 1.6

- Take about 2 g lead nitrate powder in a boiling tube.
- Hold the boiling tube with a pair of tongs and heat it over a flame, as shown in Fig. 1.5.
- What do you observe? Note down the change, if any.

You will observe the emission of brown fumes. These fumes are of nitrogen dioxide (NO_2). The reaction that takes place is –

$$2Pb(NO_3)_2(s) \xrightarrow{\text{Heat}} 2PbO(s) + 4NO_2(g) + O_2(g) \quad (1.21)$$

(Lead nitrate) (Lead oxide) (Nitrogen dioxide) (Oxygen)

Let us perform some more decomposition reactions as given in Activities 1.7 and 1.8.

Activity 1.7

- Take a plastic mug. Drill two holes at its base and fit rubber stoppers in these holes. Insert carbon electrodes in these rubber stoppers as shown in Fig. 1.6.
- Connect these electrodes to a 6 volt battery.
- Fill the mug with water such that the electrodes are immersed. Add a few drops of dilute sulphuric acid to the water.
- Take two test tubes filled with water and invert them over the two carbon electrodes.
- Switch on the current and leave the apparatus undisturbed for some time.
- You will observe the formation of bubbles at both the electrodes. These bubbles displace water in the test tubes.
- Is the volume of the gas collected the same in both the test tubes?
- Once the test tubes are filled with the respective gases, remove them carefully.
- Test these gases one by one by bringing a burning candle close to the mouth of the test tubes.

  CAUTION: This step must be performed carefully by the teacher.
- What happens in each case?
- Which gas is present in each test tube?

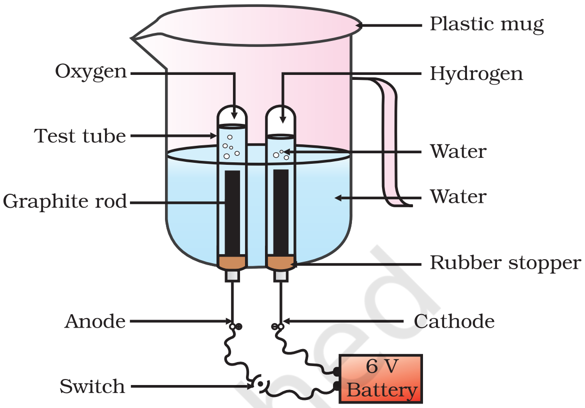


Figure 1.6
Electrolysis of water

Activity 1.8

- Take about 2 g silver chloride in a china dish.
- What is its colour?
- Place this china dish in sunlight for some time (Fig. 1.7).
- Observe the colour of the silver chloride after some time.

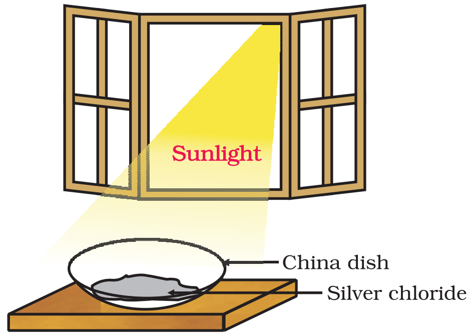


Figure 1.7
Silver chloride turns grey in sunlight to form silver metal

You will see that white silver chloride turns grey in sunlight. This is due to the decomposition of silver chloride into silver and chlorine by light.

$$2AgCl(s) \xrightarrow{\text{Sunlight}} 2Ag(s) + Cl_2(g) \quad (1.22)$$

Silver bromide also behaves in the same way.

$$2AgBr(s) \xrightarrow{\text{Sunlight}} 2Ag(s) + Br_2(g) \quad (1.23)$$

The above reactions are used in black and white photography. What form of energy is causing these decomposition reactions?

We have seen that the decomposition reactions require energy either in the form of heat, light or electricity for breaking down the reactants. Reactions in which energy is absorbed are known as endothermic reactions.

Carry out the following Activity

Take about 2 g barium hydroxide in a test tube. Add 1 g of ammonium chloride and mix with the help of a glass rod. Touch the bottom of the test tube with your palm. What do you feel? Is this an exothermic or endothermic reaction?

QUESTIONS

1. A solution of a substance 'X' is used for whitewashing.
   (i) Name the substance 'X' and write its formula.
   (ii) Write the reaction of the substance 'X' named in (i) above with water.
2. Why is the amount of gas collected in one of the test tubes in Activity 1.7 double of the amount collected in the other? Name this gas.

1.2.3 Displacement Reaction

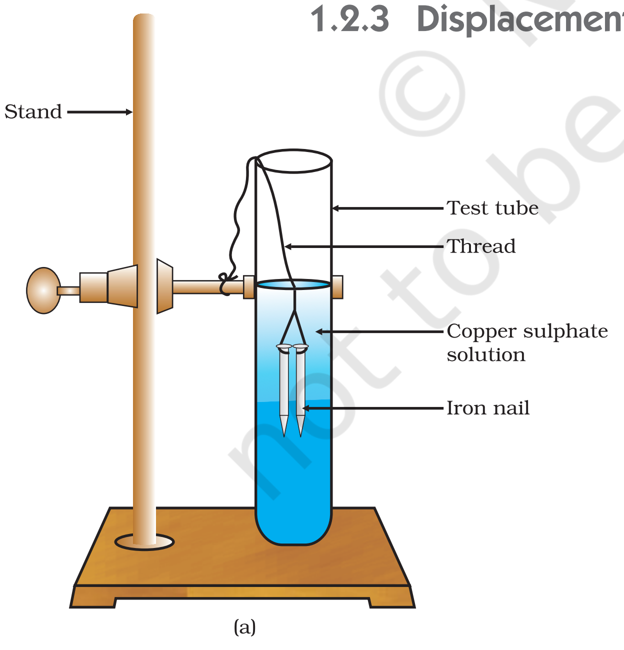


Figure 1.8
(a) Iron nails dipped in copper sulphate solution

Activity 1.9

- Take three iron nails and clean them by rubbing with sand paper.
- Take two test tubes marked as (A) and (B). In each test tube, take about 10 mL copper sulphate solution.
- Tie two iron nails with a thread and immerse them carefully in the copper sulphate solution in test tube B for about 20 minutes [Fig. 1.8 (a)]. Keep one iron nail aside for comparison.
- After 20 minutes, take out the iron nails from the copper sulphate solution.
- Compare the intensity of the blue colour of copper sulphate solutions in test tubes (A) and (B) [Fig. 1.8 (b)].
- Also, compare the colour of the iron nails dipped in the copper sulphate solution with the one kept aside [Fig. 1.8 (b)].

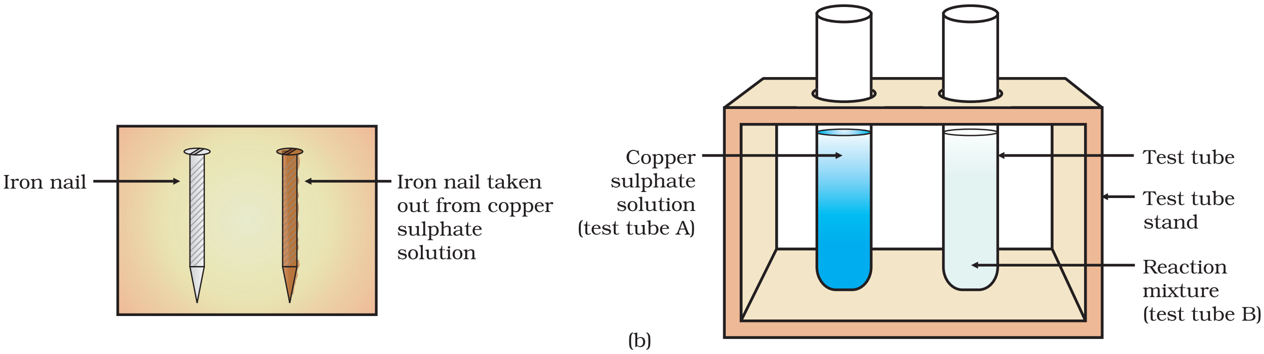


Figure 1.8 *(b) Iron nails and copper sulphate solutions compared before and after the experiment*

Why does the iron nail become brownish in colour and the blue colour of copper sulphate solution fades?

The following chemical reaction takes place in this Activity–

$$\underset{\text{(Copper sulphate)}}{Fe(s) + CuSO_4(aq)} \rightarrow \underset{\text{(Iron sulphate)}}{FeSO_4(aq)} + Cu(s) \qquad (1.24)$$

In this reaction, iron has displaced or removed another element, copper, from copper sulphate solution. This reaction is known as displacement reaction.

Other examples of displacement reactions are

$$Zn(s) + \underset{\text{(Copper sulphate)}}{CuSO_4(aq)} \rightarrow \underset{\text{(Zinc sulphate)}}{ZnSO_4(aq)} + Cu(s) \qquad (1.25)$$

$$Pb(s) + \underset{\text{(Copper chloride)}}{CuCl_2(aq)} \rightarrow \underset{\text{(Lead chloride)}}{PbCl_2(aq)} + Cu(s) \qquad (1.26)$$

Zinc and lead are more reactive elements than copper. They displace copper from its compounds.

1.2.4 Double Displacement Reaction

Activity 1.10

- Take about 3 mL of sodium sulphate solution in a test tube.
- In another test tube, take about 3 mL of barium chloride solution.
- Mix the two solutions (Fig. 1.9).
- What do you observe?

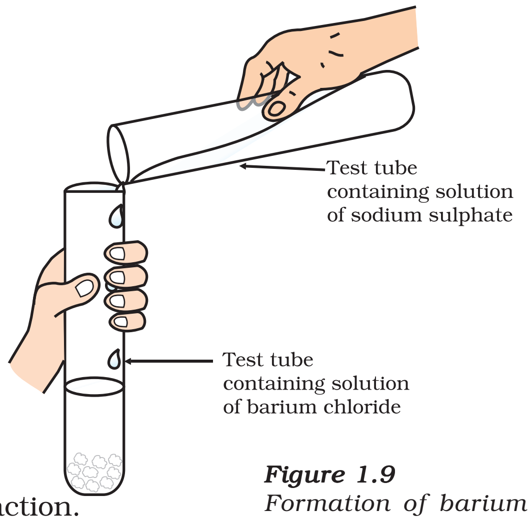


Figure 1.9 *Formation of barium sulphate and sodium chloride*

You will observe that a white substance, which is insoluble in water, is formed. This insoluble substance formed is known as a precipitate. Any reaction that produces a precipitate can be called a precipitation reaction.

$$\underset{\text{(Sodium sulphate)}}{Na_2SO_4(aq)} + \underset{\text{(Barium chloride)}}{BaCl_2(aq)} \rightarrow \underset{\text{(Barium sulphate)}}{BaSO_4(s)} + \underset{\text{(Sodium chloride)}}{2NaCl(aq)} \qquad (1.27)$$

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

Recall Activity 1.2, where you have mixed the solutions of lead(II) nitrate and potassium iodide.
(i) What was the colour of the precipitate formed? Can you name the compound precipitated?
(ii) Write the balanced chemical equation for this reaction.
(iii) Is this also a double displacement reaction?

1.2.5 Oxidation and Reduction

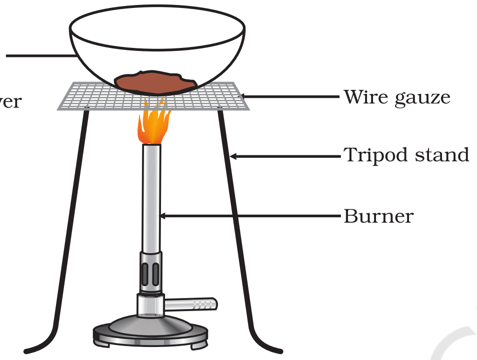


Figure 1.10
Oxidation of copper to copper oxide

Activity 1.11

- Heat a china dish containing about 1 g copper powder (Fig. 1.10).
- What do you observe?

The surface of copper powder becomes coated with black copper(II) oxide. Why has this black substance formed?
This is because oxygen is added to copper and copper oxide is formed.

$$2Cu + O_2 \xrightarrow{\text{Heat}} 2CuO \qquad (1.28)$$

If hydrogen gas is passed over this heated material (CuO), the black coating on the surface turns brown as the reverse reaction takes place and copper is obtained.

$$CuO + H_2 \xrightarrow{\text{Heat}} Cu + H_2O \qquad (1.29)$$

If a substance gains oxygen during a reaction, it is said to be oxidised. If a substance loses oxygen during a reaction, it is said to be reduced.

During this reaction (1.29), the copper(II) oxide is losing oxygen and is being reduced. The hydrogen is gaining oxygen and is being oxidised. In other words, one reactant gets oxidised while the other gets reduced during a reaction. Such reactions are called oxidation-reduction reactions or redox reactions.

Oxidation

$$CuO + H_2 \xrightarrow{\text{Heat}} Cu + H_2O \qquad (1.30)$$

Reduction

Some other examples of redox reactions are:

$$ZnO + C \rightarrow Zn + CO \qquad (1.31)$$

$$MnO_2 + 4HCl \rightarrow MnCl_2 + 2H_2O + Cl_2 \qquad (1.32)$$

In reaction (1.31) carbon is oxidised to CO and ZnO is reduced to Zn. In reaction (1.32) HCl is oxidised to Cl_2 whereas MnO_2 is reduced to $MnCl_2$.

From the above examples we can say that if a substance gains oxygen or loses hydrogen during a reaction, it is oxidised. If a substance loses oxygen or gains hydrogen during a reaction, it is reduced.

Recall Activity 1.1, where a magnesium ribbon burns with a dazzling flame in air (oxygen) and changes into a white substance, magnesium oxide. Is magnesium being oxidised or reduced in this reaction?

1.3 HAVE YOU OBSERVED THE EFFECTS OF OXIDATION REACTIONS IN EVERYDAY LIFE?

1.3.1 Corrosion

You must have observed that iron articles are shiny when new, but get coated with a reddish brown powder when left for some time. This process is commonly known as rusting of iron. Some other metals also get tarnished in this manner. Have you noticed the colour of the coating formed on copper and silver? When a metal is attacked by substances around it such as moisture, acids, etc., it is said to corrode and this process is called corrosion. The black coating on silver and the green coating on copper are other examples of corrosion.

Corrosion causes damage to car bodies, bridges, iron railings, ships and to all objects made of metals, specially those of iron. Corrosion of iron is a serious problem. Every year an enormous amount of money is spent to replace damaged iron. You will learn more about corrosion in Chapter 3.

1.3.2 Rancidity

Have you ever tasted or smelt the fat/oil containing food materials left for a long time?

When fats and oils are oxidised, they become rancid and their smell and taste change. Usually substances which prevent oxidation (antioxidants) are added to foods containing fats and oil. Keeping food in air tight containers helps to slow down oxidation. Do you know that chips manufacturers usually flush bags of chips with gas such as nitrogen to prevent the chips from getting oxidised ?

Q U E S T I O N S

1. Why does the colour of copper sulphate solution change when an iron nail is dipped in it?
2. Give an example of a double displacement reaction other than the one given in Activity 1.10.
3. Identify the substances that are oxidised and the substances that are reduced in the following reactions.
   (i) $4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$
   (ii) $CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$

What you have learnt

- A complete chemical equation represents the reactants, products and their physical states symbolically.
- A chemical equation is balanced so that the numbers of atoms of each type involved in a chemical reaction are the same on the reactant and product sides of the equation. Equations must always be balanced.
- In a combination reaction two or more substances combine to form a new single substance.
- Decomposition reactions are opposite to combination reactions. In a decomposition reaction, a single substance decomposes to give two or more substances.
- Reactions in which heat is given out along with the products are called exothermic reactions.
- Reactions in which energy is absorbed are known as endothermic reactions.
- When an element displaces another element from its compound, a displacement reaction occurs.
- Two different atoms or groups of atoms (ions) are exchanged in double displacement reactions.
- Precipitation reactions produce insoluble salts.
- Reactions also involve the gain or loss of oxygen or hydrogen by substances. Oxidation is the gain of oxygen or loss of hydrogen. Reduction is the loss of oxygen or gain of hydrogen.

EXERCISES

1. Which of the statements about the reaction below are incorrect?

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

   (a) Lead is getting reduced.
   (b) Carbon dioxide is getting oxidised.
   (c) Carbon is getting oxidised.
   (d) Lead oxide is getting reduced.

   (i) (a) and (b)
   (ii) (a) and (c)
   (iii) (a), (b) and (c)
   (iv) all

2. $Fe_2O_3 + 2Al \rightarrow Al_2O_3 + 2Fe$

   The above reaction is an example of a

   (a) combination reaction.
   (b) double displacement reaction.

(c) decomposition reaction.

(d) displacement reaction.

3. What happens when dilute hydrochloric acid is added to iron fillings? Tick the correct answer.

(a) Hydrogen gas and iron chloride are produced.

(b) Chlorine gas and iron hydroxide are produced.

(c) No reaction takes place.

(d) Iron salt and water are produced.

4. What is a balanced chemical equation? Why should chemical equations be balanced?

5. Translate the following statements into chemical equations and then balance them.

(a) Hydrogen gas combines with nitrogen to form ammonia.

(b) Hydrogen sulphide gas burns in air to give water and sulpur dioxide.

(c) Barium chloride reacts with aluminium sulphate to give aluminium chloride and a precipitate of barium sulphate.

(d) Potassium metal reacts with water to give potassium hydroxide and hydrogen gas.

6. Balance the following chemical equations.

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

(b) $NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$

(c) $NaCl + AgNO_3 \rightarrow AgCl + NaNO_3$

(d) $BaCl_2 + H_2SO_4 \rightarrow BaSO_4 + HCl$

7. Write the balanced chemical equations for the following reactions.

(a) Calcium hydroxide + Carbon dioxide $\rightarrow$ Calcium carbonate + Water

(b) Zinc + Silver nitrate $\rightarrow$ Zinc nitrate + Silver

(c) Aluminium + Copper chloride $\rightarrow$ Aluminium chloride + Copper

(d) Barium chloride + Potassium sulphate $\rightarrow$ Barium sulphate + Potassium chloride

8. Write the balanced chemical equation for the following and identify the type of reaction in each case.

(a) Potassium bromide(aq) + Barium iodide(aq) $\rightarrow$ Potassium iodide(aq) + Barium bromide(s)

(b) Zinc carbonate(s) $\rightarrow$ Zinc oxide(s) + Carbon dioxide(g)

(c) Hydrogen(g) + Chlorine(g) $\rightarrow$ Hydrogen chloride(g)

(d) Magnesium(s) + Hydrochloric acid(aq) $\rightarrow$ Magnesium chloride(aq) + Hydrogen(g)

9. What does one mean by exothermic and endothermic reactions? Give examples.

10. Why is respiration considered an exothermic reaction? Explain.

11. Why are decomposition reactions called the opposite of combination reactions? Write equations for these reactions.

12. Write one equation each for decomposition reactions where energy is supplied in the form of heat, light or electricity.
13. What is the difference between displacement and double displacement reactions? Write equations for these reactions.
14. In the refining of silver, the recovery of silver from silver nitrate solution involved displacement by copper metal. Write down the reaction involved.
15. What do you mean by a precipitation reaction? Explain by giving examples.
16. Explain the following in terms of gain or loss of oxygen with two examples each.
    (a) Oxidation
    (b) Reduction
17. A shiny brown coloured element 'X' on heating in air becomes black in colour. Name the element 'X' and the black coloured compound formed.
18. Why do we apply paint on iron articles?
19. Oil and fat containing food items are flushed with nitrogen. Why?
20. Explain the following terms with one example each.
    (a) Corrosion
    (b) Rancidity

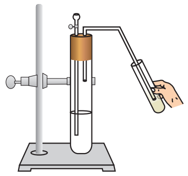
Group Activity

Perform the following activity.

- Take four beakers and label them as A, B, C and D.
- Put 25 mL of water in A, B and C beakers and copper sulphate solution in beaker D.
- Measure and record the temperature of each liquid contained in the beakers above.
- Add two spatulas of potassium sulphate, ammonium nitrate, anhydrous copper sulphate and fine iron fillings to beakers A, B, C and D respectively and stir.
- Finally measure and record the temperature of each of the mixture above.

Find out which reactions are exothermic and which ones are endothermic in nature.

CHAPTER 2



Acids, Bases and Salts

1064CH02

You have learnt in your previous classes that the sour and bitter tastes of food are due to acids and bases, respectively, present in them.

If someone in the family is suffering from a problem of acidity after overeating, which of the following would you suggest as a remedy– lemon juice, vinegar or baking soda solution?

- Which property did you think of while choosing the remedy? Surely you must have used your knowledge about the ability of acids and bases to nullify each other's effect.
- Recall how we tested sour and bitter substances without tasting them.

You already know that acids are sour in taste and change the colour of blue litmus to red, whereas, bases are bitter and change the colour of the red litmus to blue. Litmus is a natural indicator, turmeric is another such indicator. Have you noticed that a stain of curry on a white cloth becomes reddish-brown when soap, which is basic in nature, is scrubbed on it? It turns yellow again when the cloth is washed with plenty of water. You can also use synthetic indicators such as methyl orange and phenolphthalein to test for acids and bases.

In this Chapter, we will study the reactions of acids and bases, how acids and bases cancel out each other's effects and many more interesting things that we use and see in our day-to-day life.

Do You Know?

Litmus solution is a purple dye, which is extracted from lichen, a plant belonging to the division Thallophyta, and is commonly used as an indicator. When the litmus solution is neither acidic nor basic, its colour is purple. There are many other natural materials like red cabbage leaves, turmeric, coloured petals of some flowers such as *Hydrangea*, *Petunia* and *Geranium*, which indicate the presence of acid or base in a solution. These are called acid-base indicators or sometimes simply indicators.

QUESTION

1. You have been provided with three test tubes. One of them contains distilled water and the other two contain an acidic solution and a basic solution, respectively. If you are given only red litmus paper, how will you identify the contents of each test tube?

2.1 UNDERSTANDING THE CHEMICAL PROPERTIES OF ACIDS AND BASES

2.1.1 Acids and Bases in the Laboratory

Activity 2.1

- Collect the following solutions from the science laboratory– hydrochloric acid (HCl), sulphuric acid (H_2SO_4), nitric acid (HNO_3), acetic acid (CH_3COOH), sodium hydroxide (NaOH), calcium hydroxide [$Ca(OH)_2$], potassium hydroxide (KOH), magnesium hydroxide [$Mg(OH)_2$], and ammonium hydroxide (NH_4OH).
- Put a drop of each of the above solutions on a watch-glass one by one and test with a drop of the indicators shown in Table 2.1.
- What change in colour did you observe with red litmus, blue litmus, phenolphthalein and methyl orange solutions for each of the solutions taken?
- Tabulate your observations in Table 2.1.

Table 2.1

Sample solution	Red litmus solution	Blue litmus solution	Phenolph--thalein solution	Methyl orange solution

These indicators tell us whether a substance is acidic or basic by change in colour. There are some substances whose odour changes in acidic or basic media. These are called olfactory indicators. Let us try out some of these indicators.

Activity 2.2

- Take some finely chopped onions in a plastic bag along with some strips of clean cloth. Tie up the bag tightly and leave overnight in the fridge. The cloth strips can now be used to test for acids and bases.
- Take two of these cloth strips and check their odour.
- Keep them on a clean surface and put a few drops of dilute HCl solution on one strip and a few drops of dilute NaOH solution on the other.

- Rinse both cloth strips with water and again check their odour.
- Note your observations.
- Now take some dilute vanilla essence and clove oil and check their odour.
- Take some dilute HCl solution in one test tube and dilute NaOH solution in another. Add a few drops of dilute vanilla essence to both test tubes and shake well. Check the odour once again and record changes in odour, if any.
- Similarly, test the change in the odour of clove oil with dilute HCl and dilute NaOH solutions and record your observations.

Which of these – vanilla, onion and clove, can be used as olfactory indicators on the basis of your observations?

Let us do some more activities to understand the chemical properties of acids and bases.

2.1.2 How do Acids and Bases React with Metals?

Activity 2.3

CAUTION: This activity needs the teacher's assistance.

- Set the apparatus as shown in Fig. 2.1.
- Take about 5 mL of dilute sulphuric acid in a test tube and add a few pieces of zinc granules to it.
- What do you observe on the surface of zinc granules?
- Pass the gas being evolved through the soap solution.
- Why are bubbles formed in the soap solution?
- Take a burning candle near a gas filled bubble.
- What do you observe?
- Repeat this Activity with some more acids like HCl, HNO_3 and CH_3COOH.
- Are the observations in all the cases the same or different?

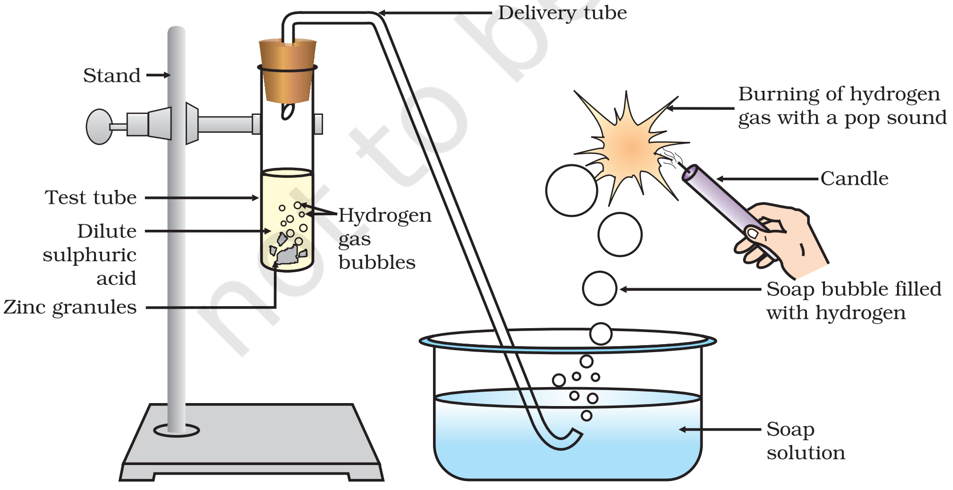


Figure 2.1 *Reaction of zinc granules with dilute sulphuric acid and testing hydrogen gas by burning*

Note that the metal in the above reactions displaces hydrogen atoms from the acids as hydrogen gas and forms a compound called a salt. Thus, the reaction of a metal with an acid can be summarised as –

Acid + Metal → Salt + Hydrogen gas

Can you now write the equations for the reactions you have observed?

Activity 2.4

- Place a few pieces of granulated zinc metal in a test tube.
- Add 2 mL of sodium hydroxide solution and warm the contents of the test tube.
- Repeat the rest of the steps as in Activity 2.3 and record your observations.

The reaction that takes place can be written as follows.

$$2NaOH(aq) + Zn(s) \rightarrow \underset{\text{(Sodium zincate)}}{Na_2ZnO_2(s)} + H_2(g)$$

You find again that hydrogen is formed in the reaction. However, such reactions are not possible with all metals.

2.1.3 How do Metal Carbonates and Metal Hydrogencarbonates React with Acids?

Activity 2.5

- Take two test tubes, label them as A and B.
- Take about 0.5 g of sodium carbonate (Na_2CO_3) in test tube A and about 0.5 g of sodium hydrogencarbonate ($NaHCO_3$) in test tube B.
- Add about 2 mL of dilute HCl to both the test tubes.
- What do you observe?
- Pass the gas produced in each case through lime water (calcium hydroxide solution) as shown in Fig. 2.2 and record your observations.

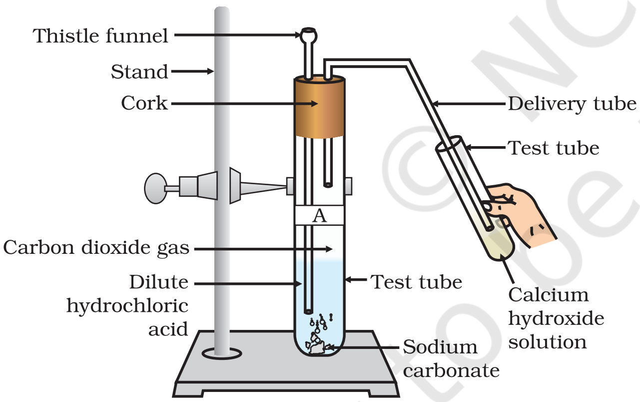


Figure 2.2
Passing carbon dioxide gas through calcium hydroxide solution

The reactions occurring in the above Activity are written as –

Test tube A: $Na_2CO_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$

Test tube B: $NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l) + CO_2(g)$

On passing the carbon dioxide gas evolved through lime water,

$$\underset{\text{(Lime water)}}{Ca(OH)_2(aq)} + CO_2(g) \rightarrow \underset{\text{(White precipitate)}}{CaCO_3(s)} + H_2O(l)$$

On passing excess carbon dioxide the following reaction takes place:

$$CaCO_3(s) + H_2O(l) + CO_2(g) \rightarrow Ca(HCO_3)_2(aq)$$

(Soluble in water)

Limestone, chalk and marble are different forms of calcium carbonate. All metal carbonates and hydrogencarbonates react with acids to give a corresponding salt, carbon dioxide and water.

Thus, the reaction can be summarised as –

Metal carbonate/Metal hydrogencarbonate + Acid → Salt + Carbon dioxide + Water

2.1.4 How do Acids and Bases React with each other?

Activity 2.6

- Take about 2 mL of dilute NaOH solution in a test tube and add two drops of phenolphthalein solution.
- What is the colour of the solution?
- Add dilute HCl solution to the above solution drop by drop.
- Is there any colour change for the reaction mixture?
- Why did the colour of phenolphthalein change after the addition of an acid?
- Now add a few drops of NaOH to the above mixture.
- Does the pink colour of phenolphthalein reappear?
- Why do you think this has happened?

In the above Activity, we have observed that the effect of a base is nullified by an acid and vice-versa. The reaction taking place is written as –

$$NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$$

The reaction between an acid and a base to give a salt and water is known as a neutralisation reaction. In general, a neutralisation reaction can be written as –

Base + Acid → Salt + Water

2.1.5 Reaction of Metallic Oxides with Acids

Activity 2.7

- Take a small amount of copper oxide in a beaker and add dilute hydrochloric acid slowly while stirring.
- Note the colour of the solution. What has happened to the copper oxide?

You will notice that the colour of the solution becomes blue-green and the copper oxide dissolves. The blue-green colour of the solution is due to the formation of copper(II) chloride in the reaction. The general reaction between a metal oxide and an acid can be written as –

Metal oxide + Acid → Salt + Water

Now write and balance the equation for the above reaction. Since metallic oxides react with acids to give salts and water, similar to the reaction of a base with an acid, metallic oxides are said to be basic oxides.

2.1.6 Reaction of a Non-metallic Oxide with Base

You saw the reaction between carbon dioxide and calcium hydroxide (lime water) in Activity 2.5. Calcium hydroxide, which is a base, reacts with carbon dioxide to produce a salt and water. Since this is similar to the reaction between a base and an acid, we can conclude that non-metallic oxides are acidic in nature.

QUESTIONS

1. Why should curd and sour substances not be kept in brass and copper vessels?
2. Which gas is usually liberated when an acid reacts with a metal? Illustrate with an example. How will you test for the presence of this gas?
3. Metal compound A reacts with dilute hydrochloric acid to produce effervescence. The gas evolved extinguishes a burning candle. Write a balanced chemical equation for the reaction if one of the compounds formed is calcium chloride.

2.2 WHAT DO ALL ACIDS AND ALL BASES HAVE IN COMMON?

In Section 2.1 we have seen that all acids have similar chemical properties. What leads to this similarity in properties? We saw in Activity 2.3 that all acids generate hydrogen gas on reacting with metals, so hydrogen seems to be common to all acids. Let us perform an Activity to investigate whether all compounds containing hydrogen are acidic.

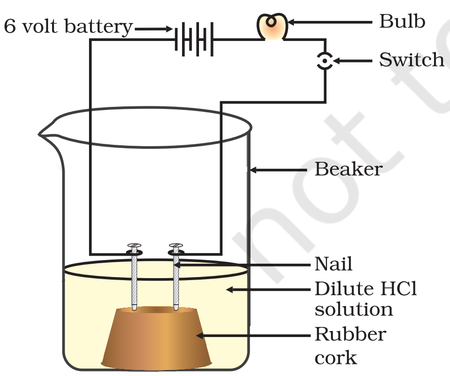


Figure 2.3
Acid solution in water conducts electricity

Activity 2.8

- Take solutions of glucose, alcohol, hydrochloric acid, sulphuric acid, etc.
- Fix two nails on a cork, and place the cork in a 100 mL beaker.
- Connect the nails to the two terminals of a 6 volt battery through a bulb and a switch, as shown in Fig. 2.3.
- Now pour some dilute HCl in the beaker and switch on the current.
- Repeat with dilute sulphuric acid.
- What do you observe?
- Repeat the experiment separately with glucose and alcohol solutions. What do you observe now?
- Does the bulb glow in all cases?

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The bulb will start glowing in the case of acids, as shown in Fig. 2.3. But you will observe that glucose and alcohol solutions do not conduct electricity. Glowing of the bulb indicates that there is a flow of electric current through the solution. The electric current is carried through the acidic solution by ions.

Acids contain H^+ ion as cation and anion such as Cl^- in HCl, NO_3^- in HNO_3, SO_4^{2-} in H_2SO_4, CH_3COO^- in CH_3COOH. Since the cation present in acids is H^+, this suggests that acids produce hydrogen ions, $H^+(aq)$, in solution, which are responsible for their acidic properties.

Repeat the same Activity using alkalis such as sodium hydroxide, calcium hydroxide, etc. What can you conclude from the results of this Activity?

2.2.1 What Happens to an Acid or a Base in a Water Solution?

Do acids produce ions only in aqueous solution? Let us test this.

Activity 2.9

- Take about 1g solid NaCl in a clean and dry test tube and set up the apparatus as shown in Fig. 2.4.
- Add some concentrated sulphuric acid to the test tube.
- What do you observe? Is there a gas coming out of the delivery tube?
- Test the gas evolved successively with dry and wet blue litmus paper.
- In which case does the litmus paper change colour?
- On the basis of the above Activity, what do you infer about the acidic character of:
  (i) dry HCl gas
  (ii) HCl solution?

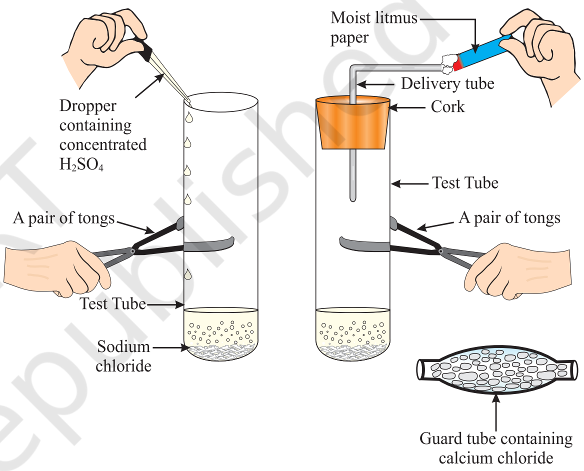

Figure 2.4 *Preparation of* HCl *gas*

Note to teachers: If the climate is very humid, you will have to pass the gas produced through a guard tube (drying tube) containing calcium chloride to dry the gas.

This experiment suggests that hydrogen ions in HCl are produced in the presence of water. The separation of H^+ ion from HCl molecules cannot occur in the absence of water.

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

Hydrogen ions cannot exist alone, but they exist after combining with water molecules. Thus hydrogen ions must always be shown as $H^+(aq)$ or hydronium ion (H_3O^+).

$$H^+ + H_2O \rightarrow H_3O^+$$

We have seen that acids give H_3O^+ or $H^+(aq)$ ion in water. Let us see what happens when a base is dissolved in water.

$$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$$

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$$KOH(s) \xrightarrow{H_2O} K^+(aq) + OH^-(aq)$$

$$Mg(OH)_2(s) \xrightarrow{H_2O} Mg^{2+}(aq) + 2OH^-(aq)$$

Bases generate hydroxide (OH^-) ions in water. Bases which are soluble in water are called alkalis.

Do You Know?

All bases do not dissolve in water. An alkali is a base that dissolves in water. They are soapy to touch, bitter and corrosive. Never taste or touch them as they may cause harm. Which of the bases in the Table 2.1 are alkalis?

Now as we have identified that all acids generate $H^+(aq)$ and all bases generate $OH^-(aq)$, we can view the neutralisation reaction as follows –

$$\text{Acid} + \text{Base} \rightarrow \text{Salt} + \text{Water}$$

$$H\boxed{X + M}OH \rightarrow MX + HOH$$

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Let us see what is involved when water is mixed with an acid or a base.

Figure 2.5
Warning sign displayed on containers containing concentrated acids and bases

Activity 2.10

- Take 10 mL water in a beaker.
- Add a few drops of concentrated H_2SO_4 to it and swirl the beaker slowly.
- Touch the base of the beaker.
- Is there a change in temperature?
- Is this an exothermic or endothermic process?
- Repeat the above Activity with sodium hydroxide pellets and record your observations.

The process of dissolving an acid or a base in water is a highly exothermic one. Care must be taken while mixing concentrated nitric acid or sulphuric acid with water. The acid must always be added slowly to water with constant stirring. If water is added to a concentrated acid, the heat generated may cause the mixture to splash out and cause burns. The glass container may also break due to excessive local heating. Look out for the warning sign (shown in Fig. 2.5) on the can of concentrated sulphuric acid and on the bottle of sodium hydroxide pellets.

Mixing an acid or base with water results in decrease in the concentration of ions (H_3O^+/OH^-) per unit volume. Such a process is called dilution and the acid or the base is said to be diluted.

Q U E S T I O N S

1. Why do HCl, HNO_3, etc., show acidic characters in aqueous solutions while solutions of compounds like alcohol and glucose do not show acidic character?
2. Why does an aqueous solution of an acid conduct electricity?
3. Why does dry HCl gas not change the colour of the dry litmus paper?
4. While diluting an acid, why is it recommended that the acid should be added to water and not water to the acid?
5. How is the concentration of hydronium ions (H_3O^+) affected when a solution of an acid is diluted?
6. How is the concentration of hydroxide ions (OH^-) affected when excess base is dissolved in a solution of sodium hydroxide?

2.3 HOW STRONG ARE ACID OR BASE SOLUTIONS?

We know how acid-base indicators can be used to distinguish between an acid and a base. We have also learnt in the previous section about dilution and decrease in concentration of H^+ or OH^- ions in solutions. Can we quantitatively find the amount of these ions present in a solution? Can we judge how strong a given acid or base is?

We can do this by making use of a universal indicator, which is a mixture of several indicators. The universal indicator shows different colours at different concentrations of hydrogen ions in a solution.

A scale for measuring hydrogen ion concentration in a solution, called pH scale has been developed. The p in pH stands for '*potenz*' in German, meaning power. On the pH scale we can measure pH generally from 0 (very acidic) to 14 (very alkaline). pH should be thought of simply as a number which indicates the acidic or basic nature of a solution. Higher the hydronium ion concentration, lower is the pH value.

The pH of a neutral solution is 7. Values less than 7 on the pH scale represent an acidic solution. As the pH value increases from 7 to 14, it represents an increase in OH^- ion concentration in the solution, that is, increase in the strength of alkali (Fig. 2.6). Generally paper impregnated with the universal indicator is used for measuring pH.

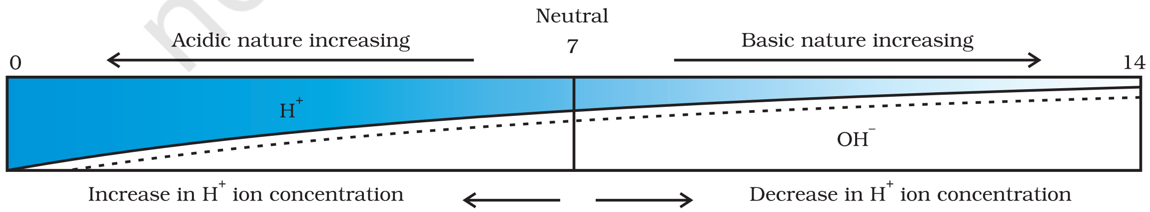

Figure 2.6 *Variation of* pH *with the change in concentration of* H^+(aq) *and* OH^-(aq) *ions*

Activity 2.11

- Test the pH values of solutions given in Table 2.2.
- Record your observations.
- What is the nature of each substance on the basis of your observations?

Table 2.2

S. No.	Solution	Colour of pH paper	Approx-imate pH value	Nature of substance
1	Saliva (before meal)			
2	Saliva (after meal)			
3	Lemon juice			
4	Colourless aerated drink			
5	Carrot juice			
6	Coffee			
7	Tomato juice			
8	Tap water			
9	1M NaOH			
10	1M HCl			

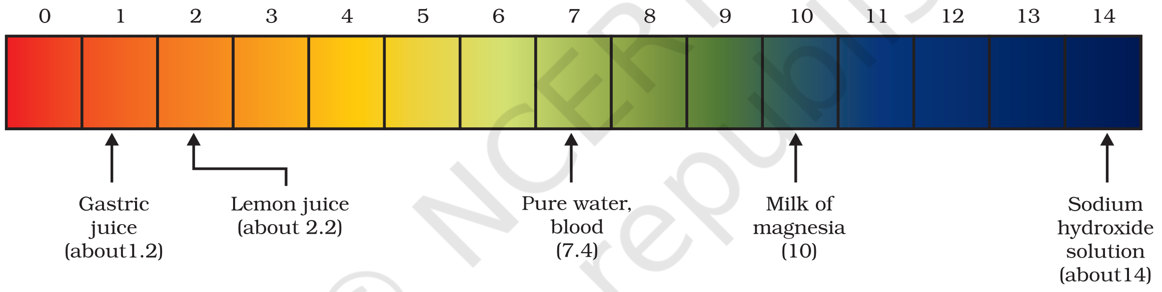

Figure 2.7 pH *of some common substances shown on a* pH *paper (colours are only a rough guide)*

The strength of acids and bases depends on the number of H^+ ions and OH^- ions produced, respectively. If we take hydrochloric acid and acetic acid of the same concentration, say one molar, then these produce different amounts of hydrogen ions. Acids that give rise to more H^+ ions are said to be strong acids, and acids that give less H^+ ions are said to be weak acids. Can you now say what weak and strong bases are?

2.3.1 Importance of pH in Everyday Life

Are plants and animals pH sensitive?

Our body works within the pH range of 7.0 to 7.8. Living organisms can survive only in a narrow range of pH change. When pH of rain water is less than 5.6, it is called acid rain. When acid rain flows into the rivers, it lowers the pH of the river water. The survival of aquatic life in such rivers becomes difficult.

Do You Know?

Acids in other planets

The atmosphere of venus is made up of thick white and yellowish clouds of sulphuric acid. Do you think life can exist on this planet?

What is the pH of the soil in your backyard?

Plants require a specific pH range for their healthy growth. To find out the pH required for the healthy growth of a plant, you can collect the soil from various places and check the pH in the manner described below in Activity 2.12. Also, you can note down which plants are growing in the region from which you have collected the soil.

Activity 2.12

- Put about 2 g soil in a test tube and add 5 mL water to it.
- Shake the contents of the test tube.
- Filter the contents and collect the filtrate in a test tube.
- Check the pH of this filtrate with the help of universal indicator paper.
- What can you conclude about the ideal soil pH for the growth of plants in your region?

pH in our digestive system

It is very interesting to note that our stomach produces hydrochloric acid. It helps in the digestion of food without harming the stomach. During indigestion the stomach produces too much acid and this causes pain and irritation. To get rid of this pain, people use bases called antacids. One such remedy must have been suggested by you at the beginning of this Chapter. These antacids neutralise the excess acid. Magnesium hydroxide (Milk of magnesia), a mild base, is often used for this purpose.

pH change as the cause of tooth decay

Tooth decay starts when the pH of the mouth is lower than 5.5. Tooth enamel, made up of calcium hydroxyapatite (a crystalline form of calcium phosphate) is the hardest substance in the body. It does not dissolve in water, but is corroded when the pH in the mouth is below 5.5. Bacteria present in the mouth produce acids by degradation of sugar and food particles remaining in the mouth after eating. The best way to prevent this is to clean the mouth after eating food. Using toothpastes, which are generally basic, for cleaning the teeth can neutralise the excess acid and prevent tooth decay.

Self defence by animals and plants through chemical warfare

Have you ever been stung by a honey-bee? Bee-sting leaves an acid which causes pain and irritation. Use of a mild base like baking soda on the stung area gives relief. Stinging hair of nettle leaves inject methanoic acid causing burning pain.

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Do You Know?

Nature provides neutralisation options

Nettle is a herbaceous plant which grows in the wild. Its leaves have stinging hair, which cause painful stings when touched accidentally. This is due to the methanoic acid secreted by them. A traditional remedy is rubbing the area with the leaf of the dock plant, which often grows beside the nettle in the wild. Can you guess the nature of the dock plant? So next time you know what to look out for if you accidentally touch a nettle plant while trekking. Are you aware of any other effective traditional remedies for such stings?

Table 2.3 Some naturally occurring acids

Natural source	Acid	Natural source	Acid
Vinegar	Acetic acid	Sour milk (Curd)	Lactic acid
Orange	Citric acid	Lemon	Citric acid
Tamarind	Tartaric acid	Ant sting	Methanoic acid
Tomato	Oxalic acid	Nettle sting	Methanoic acid

QUESTIONS

1. You have two solutions, A and B. The pH of solution A is 6 and pH of solution B is 8. Which solution has more hydrogen ion concentration? Which of this is acidic and which one is basic?
2. What effect does the concentration of $H^+(aq)$ ions have on the nature of the solution?
3. Do basic solutions also have $H^+(aq)$ ions? If yes, then why are these basic?
4. Under what soil condition do you think a farmer would treat the soil of his fields with quick lime (calcium oxide) or slaked lime (calcium hydroxide) or chalk (calcium carbonate)?

2.4 MORE ABOUT SALTS

In the previous sections we have seen the formation of salts during various reactions. Let us understand more about their preparation, properties and uses.

2.4.1 Family of Salts

Activity 2.13

- Write the chemical formulae of the salts given below. Potassium sulphate, sodium sulphate, calcium sulphate, magnesium sulphate, copper sulphate, sodium chloride, sodium nitrate, sodium carbonate and ammonium chloride.

- Identify the acids and bases from which the above salts may be obtained.
- Salts having the same positive or negative radicals are said to belong to a family. For example, NaCl and Na_2SO_4 belong to the family of sodium salts. Similarly, NaCl and KCl belong to the family of chloride salts. How many families can you identify among the salts given in this Activity?

2.4.2 pH of Salts

Activity 2.14

- Collect the following salt samples – sodium chloride, potassium nitrate, aluminium chloride, zinc sulphate, copper sulphate, sodium acetate, sodium carbonate and sodium hydrogencarbonate (some other salts available can also be taken).
- Check their solubility in water (use distilled water only).
- Check the action of these solutions on litmus and find the pH using a pH paper.
- Which of the salts are acidic, basic or neutral?
- Identify the acid or base used to form the salt.
- Report your observations in Table 2.4.

Salts of a strong acid and a strong base are neutral with pH value of 7. On the other hand, salts of a strong acid and weak base are acidic with pH value less than 7 and those of a strong base and weak acid are basic in nature, with pH value more than 7.

Table 2.4

Salt	pH	Acid used	Base used

2.4.3 Chemicals from Common Salt

By now you have learnt that the salt formed by the combination of hydrochloric acid and sodium hydroxide solution is called sodium chloride. This is the salt that you use in food. You must have observed in the above Activity that it is a neutral salt.

Seawater contains many salts dissolved in it. Sodium chloride is separated from these salts. Deposits of solid salt are also found in several parts of the world. These large crystals are often brown due to impurities. This is called rock salt. Beds of rock salt were formed when seas of bygone ages dried up. Rock salt is mined like coal.

You must have heard about Mahatma Gandhi's *Dandi March*. Did you know that sodium chloride was such an important symbol in our struggle for freedom?

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Common salt — A raw material for chemicals

The common salt thus obtained is an important raw material for various materials of daily use, such as sodium hydroxide, baking soda, washing soda, bleaching powder and many more. Let us see how one substance is used for making all these different substances.

Sodium hydroxide

When electricity is passed through an aqueous solution of sodium chloride (called brine), it decomposes to form sodium hydroxide. The process is called the chlor-alkali process because of the products formed– chlor for chlorine and alkali for sodium hydroxide.

$$2NaCl(aq) + 2H_2O(l) \rightarrow 2NaOH(aq) + Cl_2(g) + H_2(g)$$

Chlorine gas is given off at the anode, and hydrogen gas at the cathode. Sodium hydroxide solution is formed near the cathode. The three products produced in this process are all useful. Figure 2.8 shows the different uses of these products.

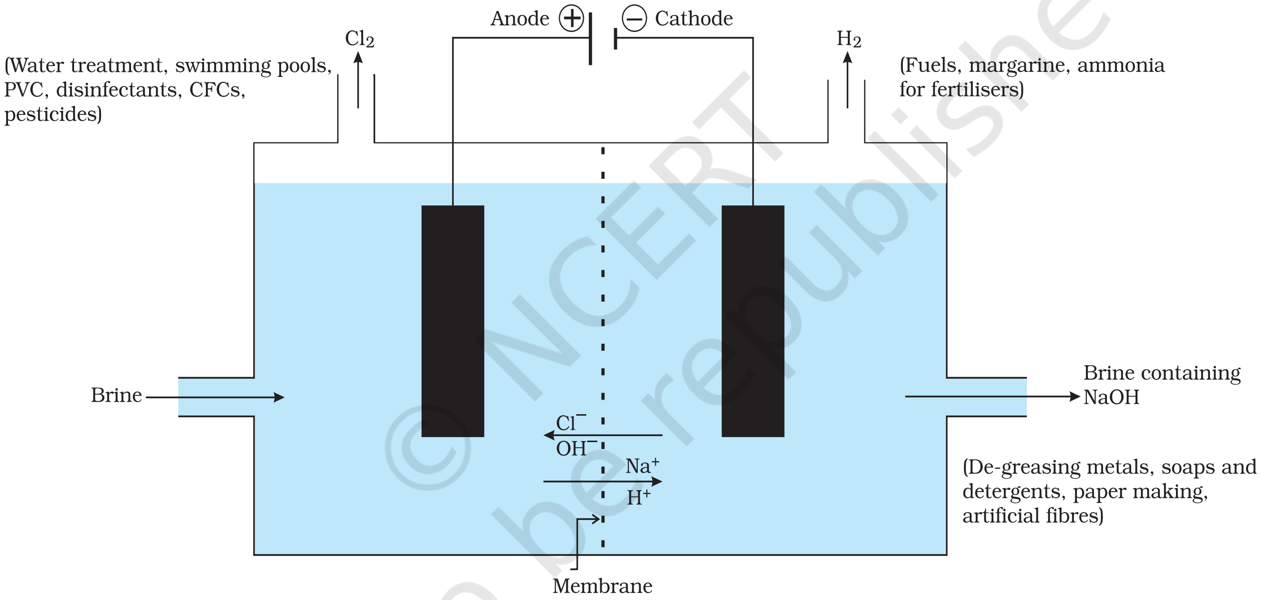

Figure 2.8 *Important products from the chlor-alkali process*

Bleaching powder

You have already come to know that chlorine is produced during the electrolysis of aqueous sodium chloride (brine). This chlorine gas is used for the manufacture of bleaching powder. Bleaching powder is produced by the action of chlorine on dry slaked lime [$Ca(OH)_2$]. Bleaching powder is represented as $CaOCl_2$, though the actual composition is quite complex.

$$Ca(OH)_2 + Cl_2 \rightarrow CaOCl_2 + H_2O$$

Bleaching powder is used –

(i) for bleaching cotton and linen in the textile industry, for bleaching wood pulp in paper factories and for bleaching washed clothes in laundry;

(ii) as an oxidising agent in many chemical industries; and

(iii) to make drinking water free from germs.

Baking soda

The baking soda is commonly used in the kitchen for making tasty crispy pakoras, etc. Sometimes it is added for faster cooking. The chemical name of the compound is sodium hydrogencarbonate ($NaHCO_3$). It is produced using sodium chloride as one of the raw materials.

$$NaCl + H_2O + CO_2 + NH_3 \rightarrow \underset{\text{(Ammonium chloride)}}{NH_4Cl} + \underset{\text{(Sodium hydrogencarbonate)}}{NaHCO_3}$$

Did you check the pH of sodium hydrogencarbonate in Activity 2.14? Can you correlate why it can be used to neutralise an acid? It is a mild non-corrosive basic salt. The following reaction takes place when it is heated during cooking –

$$\underset{\text{(Sodium hydrogencarbonate)}}{2NaHCO_3} \xrightarrow{\text{Heat}} \underset{\text{(Sodium carbonate)}}{Na_2CO_3} + H_2O + CO_2$$

Sodium hydrogencarbonate has got various uses in the household.

Uses of Baking soda

(i) For making baking powder, which is a mixture of baking soda (sodium hydrogencarbonate) and a mild edible acid such as tartaric acid. When baking powder is heated or mixed in water, the following reaction takes place –

$$NaHCO_3 + \underset{\text{(From any acid)}}{H^+} \rightarrow CO_2 + H_2O + \text{Sodium salt of acid}$$

Carbon dioxide produced during the reaction can cause bread or cake to rise making them soft and spongy.

(ii) Sodium hydrogencarbonate is also an ingredient in antacids. Being alkaline, it neutralises excess acid in the stomach and provides relief.

(iii) It is also used in soda-acid fire extinguishers.

Washing soda

Another chemical that can be obtained from sodium chloride is $Na_2CO_3.10H_2O$ (washing soda). You have seen above that sodium carbonate can be obtained by heating baking soda; recrystallisation of sodium carbonate gives washing soda. It is also a basic salt.

$$\underset{\text{(Sodium carbonate)}}{Na_2CO_3} + 10\,H_2O \rightarrow Na_2CO_3.10\,H_2O$$

What does $10H_2O$ signify? Does it make Na_2CO_3 wet? We will address this question in the next section.

Sodium carbonate and sodium hydrogencarbonate are useful chemicals for many industrial processes as well.

Uses of washing soda

(i) Sodium carbonate (washing soda) is used in glass, soap and paper industries.
(ii) It is used in the manufacture of sodium compounds such as borax.
(iii) Sodium carbonate can be used as a cleaning agent for domestic purposes.
(iv) It is used for removing permanent hardness of water.

2.4.4 Are the Crystals of Salts really Dry?

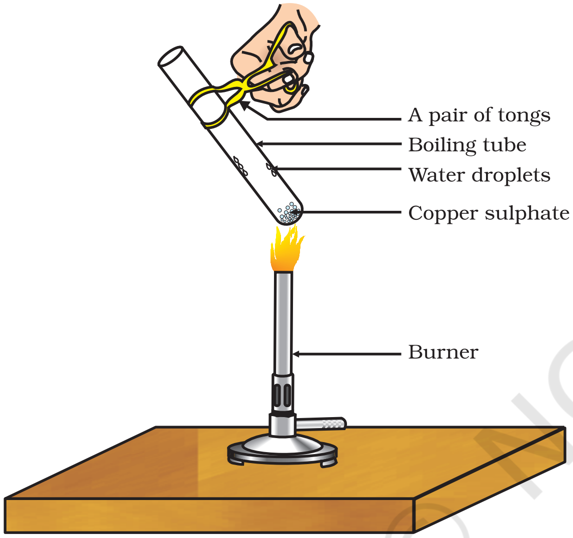

Figure 2.9
Removing water of crystallisation

Activity 2.15

- Heat a few crystals of copper sulphate in a dry boiling tube.
- What is the colour of the copper sulphate after heating?
- Do you notice water droplets in the boiling tube? Where have these come from?
- Add 2-3 drops of water on the sample of copper sulphate obtained after heating.
- What do you observe? Is the blue colour of copper sulphate restored?

Copper sulphate crystals which seem to be dry contain water of crystallisation. When we heat the crystals, this water is removed and the salt turns white.

If you moisten the crystals again with water, you will find that blue colour of the crystals reappears.

Water of crystallisation is the fixed number of water molecules present in one formula unit of a salt. Five water molecules are present in one formula unit of copper sulphate. Chemical formula for hydrated copper sulphate is $Cu\ SO_4.\ 5H_2O$. Now you would be able to answer the question whether the molecule of $Na_2CO_3.10H_2O$ is wet.

One other salt, which possesses water of crystallisation is gypsum. It has two water molecules as water of cyrstallisation. It has the chemical formula $CaSO_4.2H_2O$. Let us look into the use of this salt.

Plaster of Paris

On heating gypsum at 373 K, it loses water molecules and becomes calcium sulphate hemihydrate $(CaSO_4 . \frac{1}{2} H_2O)$. This is called Plaster of

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Paris, the substance which doctors use as plaster for supporting fractured bones in the right position. Plaster of Paris is a white powder and on mixing with water, it changes to gypsum once again giving a hard solid mass.

$$\underset{\text{(Plaster of Paris)}}{CaSO_4 . \frac{1}{2} H_2O} + 1\frac{1}{2} H_2O \rightarrow \underset{\text{(Gypsum)}}{CaSO_4 .2H_2O}$$

Note that only half a water molecule is shown to be attached as water of crystallisation. How can you get half a water molecule? It is written in this form because two formula units of $CaSO_4$ share one molecule of water. Plaster of Paris is used for making toys, materials for decoration and for making surfaces smooth. Try to find out why is calcium sulphate hemihydrate called 'Plaster of Paris' ?

QUESTIONS

1. What is the common name of the compound $CaOCl_2$?
2. Name the substance which on treatment with chlorine yields bleaching powder.
3. Name the sodium compound which is used for softening hard water.
4. What will happen if a solution of sodium hydrocarbonate is heated? Give the equation of the reaction involved.
5. Write an equation to show the reaction between Plaster of Paris and water.

What you have learnt

- Acid-base indicators are dyes or mixtures of dyes which are used to indicate the presence of acids and bases.
- Acidic nature of a substance is due to the formation of $H^+(aq)$ ions in solution. Formation of $OH^-(aq)$ ions in solution is responsible for the basic nature of a substance.
- When an acid reacts with a metal, hydrogen gas is evolved and a corresponding salt is formed.
- When a base reacts with a metal, along with the evolution of hydrogen gas a salt is formed which has a negative ion composed of the metal and oxygen.
- When an acid reacts with a metal carbonate or metal hydrogencarbonate, it gives the corresponding salt, carbon dioxide gas and water.
- Acidic and basic solutions in water conduct electricity because they produce hydrogen and hydroxide ions respectively.

- The strength of an acid or an alkali can be tested by using a scale called the pH scale (0-14) which gives the measure of hydrogen ion concentration in a solution.
- A neutral solution has a pH of exactly 7, while an acidic solution has a pH less than 7 and a basic solution a pH more than 7.
- Living beings carry out their metabolic activities within an optimal pH range.
- Mixing concentrated acids or bases with water is a highly exothermic process.
- Acids and bases neutralise each other to form corresponding salts and water.
- Water of crystallisation is the fixed number of water molecules present in one formula unit of a salt.
- Salts have various uses in everyday life and in industries.

EXERCISES

1. A solution turns red litmus blue, its pH is likely to be
   (a) 1 (b) 4 (c) 5 (d) 10
2. A solution reacts with crushed egg-shells to give a gas that turns lime-water milky. The solution contains
   (a) NaCl (b) HCl (c) LiCl (d) KCl
3. 10 mL of a solution of NaOH is found to be completely neutralised by 8 mL of a given solution of HCl. If we take 20 mL of the same solution of NaOH, the amount HCl solution (the same solution as before) required to neutralise it will be
   (a) 4 mL (b) 8 mL (c) 12 mL (d) 16 mL
4. Which one of the following types of medicines is used for treating indigestion?
   (a) Antibiotic
   (b) Analgesic
   (c) Antacid
   (d) Antiseptic
5. Write word equations and then balanced equations for the reaction taking place when –
   (a) dilute sulphuric acid reacts with zinc granules.
   (b) dilute hydrochloric acid reacts with magnesium ribbon.
   (c) dilute sulphuric acid reacts with aluminium powder.
   (d) dilute hydrochloric acid reacts with iron filings.
6. Compounds such as alcohols and glucose also contain hydrogen but are not categorised as acids. Describe an Activity to prove it.
7. Why does distilled water not conduct electricity, whereas rain water does?

8. Why do acids not show acidic behaviour in the absence of water?
9. Five solutions A,B,C,D and E when tested with universal indicator showed pH as 4,1,11,7 and 9, respectively. Which solution is
   (a) neutral?
   (b) strongly alkaline?
   (c) strongly acidic?
   (d) weakly acidic?
   (e) weakly alkaline?
   Arrange the pH in increasing order of hydrogen-ion concentration.
10. Equal lengths of magnesium ribbons are taken in test tubes A and B. Hydrochloric acid (HCl) is added to test tube A, while acetic acid (CH_3COOH) is added to test tube B. Amount and concentration taken for both the acids are same. In which test tube will the fizzing occur more vigorously and why?
11. Fresh milk has a pH of 6. How do you think the pH will change as it turns into curd? Explain your answer.
12. A milkman adds a very small amount of baking soda to fresh milk.
    (a) Why does he shift the pH of the fresh milk from 6 to slightly alkaline?
    (b) Why does this milk take a long time to set as curd?
13. Plaster of Paris should be stored in a moisture-proof container. Explain why?
14. What is a neutralisation reaction? Give two examples.
15. Give two important uses of washing soda and baking soda.

Group Activity

(I) **Prepare your own indicator**

- Crush beetroot in a mortar.
- Add sufficient water to obtain the extract.
- Filter the extract by the procedure learnt by you in earlier classes.
- Collect the filtrate to test the substances you may have tasted earlier.
- Arrange four test tubes in a test tube stand and label them as A,B,C and D. Pour 2 mL each of lemon juice solution, soda-water, vinegar and baking soda solution in them respectively.
- Put 2-3 drops of the beetroot extract in each test tube and note the colour change if any. Write your observation in a Table.
- You can prepare indicators by using other natural materials like extracts of red cabbage leaves, coloured petals of some flowers such as *Petunia*, *Hydrangea* and *Geranium*.

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(II) **Preparing a soda-acid fire extinguisher**

The reaction of acids with metal hydrogencarbonates is used in the fire extinguishers which produce carbon dioxide.

- Take 20 mL of sodium hydrogencarbonate ($NaHCO_3$) solution in a wash-bottle.
- Suspend an ignition tube containing dilute sulphuric acid in the wash-bottle (Fig. 2.10).
- Close the mouth of the wash-bottle.
- Tilt the wash-bottle so that the acid from the ignition tube mixes with the sodium hydrogencarbonate solution below.
- You will notice discharge coming out of the nozzle.
- Direct this discharge on a burning candle. What happens?

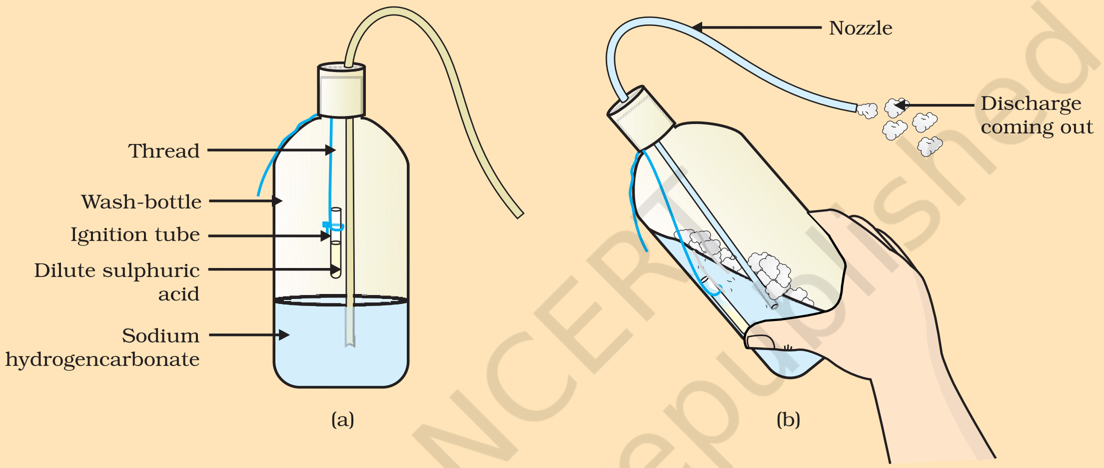
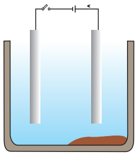


Figure 2.10 *(a) Ignition tube containing dilute sulphuric acid suspended in a wash-bottle containing sodium hydrogencarbonate, (b) Discharge coming out of the nozzle*

CHAPTER 3



Metals and Non-metals

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In Class IX you have learnt about various elements. You have seen that elements can be classified as metals or non-metals on the basis of their properties.

- Think of some uses of metals and non-metals in your daily life.
- What properties did you think of while categorising elements as metals or non-metals?
- How are these properties related to the uses of these elements?

Let us look at some of these properties in detail.

3.1 PHYSICAL PROPERTIES

3.1.1 Metals

The easiest way to start grouping substances is by comparing their physical properties. Let us study this with the help of the following activities. For performing Activities 3.1 to 3.6, collect the samples of following metals – iron, copper, aluminium, magnesium, sodium, lead, zinc and any other metal that is easily available.

Activity 3.1

- Take samples of iron, copper, aluminium and magnesium. Note the appearance of each sample.
- Clean the surface of each sample by rubbing them with sand paper and note their appearance again.

Metals, in their pure state, have a shining surface. This property is called metallic lustre.

Activity 3.2

- Take small pieces of iron, copper, aluminium, and magnesium. Try to cut these metals with a sharp knife and note your observations.
- Hold a piece of sodium metal with a pair of tongs. CAUTION: Always handle sodium metal with care. Dry it by pressing between the folds of a filter paper.
- Put it on a watch-glass and try to cut it with a knife.
- What do you observe?

You will find that metals are generally hard. The hardness varies from metal to metal.

Activity 3.3

- Take pieces of iron, zinc, lead and copper.
- Place any one metal on a block of iron and strike it four or five times with a hammer. What do you observe?
- Repeat with other metals.
- Record the change in the shape of these metals.

You will find that some metals can be beaten into thin sheets. This property is called malleability. Did you know that gold and silver are the most malleable metals?

Activity 3.4

- List the metals whose wires you have seen in daily life.

The ability of metals to be drawn into thin wires is called ductility. Gold is the most ductile metal. You will be surprised to know that a wire of about 2 km length can be drawn from one gram of gold.

It is because of their malleability and ductility that metals can be given different shapes according to our needs.

Can you name some metals that are used for making cooking vessels? Do you know why these metals are used for making vessels? Let us do the following Activity to find out the answer.

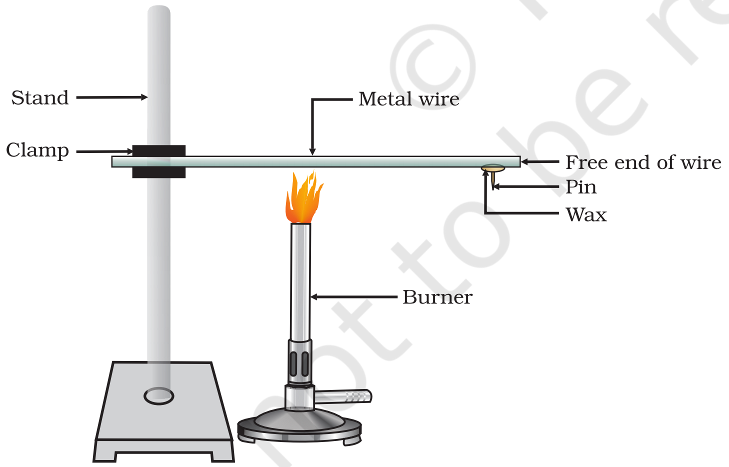

Figure 3.1
Metals are good conductors of heat.

Activity 3.5

- Take an aluminium or copper wire. Clamp this wire on a stand, as shown in Fig. 3.1.
- Fix a pin to the free end of the wire using wax.
- Heat the wire with a spirit lamp, candle or a burner near the place where it is clamped.
- What do you observe after some time?
- Note your observations. Does the metal wire melt?

The above activity shows that metals are good conductors of heat and have high melting points. The best conductors of heat are silver and copper. Lead and mercury are comparatively poor conductors of heat.

Do metals also conduct electricity? Let us find out.

Activity 3.6

- Set up an electric circuit as shown in Fig. 3.2.
- Place the metal to be tested in the circuit between terminals A and B as shown.
- Does the bulb glow? What does this indicate?

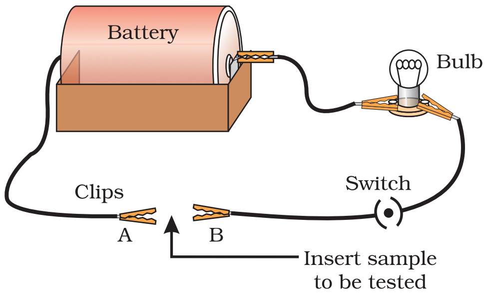

Figure 3.2
Metals are good conductors of electricity.

You must have seen that the wires that carry current in your homes have a coating of polyvinylchloride (PVC) or a rubber-like material. Why are electric wires coated with such substances?

What happens when metals strike a hard surface? Do they produce a sound? The metals that produce a sound on striking a hard surface are said to be sonorous. Can you now say why school bells are made of metals?

3.1.2 Non-metals

In the previous Class you have learnt that there are very few non-metals as compared to metals. Some of the examples of non-metals are carbon, sulphur, iodine, oxygen, hydrogen, etc. The non-metals are either solids or gases except bromine which is a liquid.

Do non-metals also have physical properties similar to that of metals? Let us find out.

Activity 3.7

- Collect samples of carbon (coal or graphite), sulphur and iodine.
- Carry out the Activities 3.1 to 3.4 and 3.6 with these non-metals and record your observations.

Compile your observations regarding metals and non-metals in Table 3.1.

Table 3.1

Element	Symbol	Type of surface	Hardness	Malleability	Ductility	Conducts Electricity	Sonority

On the bases of the observations recorded in Table 3.1, discuss the general physical properties of metals and non-metals in the class. You must have concluded that we cannot group elements according to their physical properties alone, as there are many exceptions. For example –

(i) All metals except mercury exist as solids at room temperature. In Activity 3.5, you have observed that metals have high melting

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points but gallium and caesium have very low melting points. These two metals will melt if you keep them on your palm.

(ii) Iodine is a non-metal but it is lustrous.

(iii) Carbon is a non-metal that can exist in different forms. Each form is called an allotrope. Diamond, an allotrope of carbon, is the hardest natural substance known and has a very high melting and boiling point. Graphite, another allotrope of carbon, is a conductor of electricity.

(iv) Alkali metals (lithium, sodium, potassium) are so soft that they can be cut with a knife. They have low densities and low melting points.

Elements can be more clearly classified as metals and non-metals on the basis of their chemical properties.

Activity 3.8

- Take a magnesium ribbon and some sulphur powder.
- Burn the magnesium ribbon. Collect the ashes formed and dissolve them in water.
- Test the resultant solution with both red and blue litmus paper.
- Is the product formed on burning magnesium acidic or basic?
- Now burn sulphur powder. Place a test tube over the burning sulphur to collect the fumes produced.
- Add some water to the above test tube and shake.
- Test this solution with blue and red litmus paper.
- Is the product formed on burning sulphur acidic or basic?
- Can you write equations for these reactions?

Most non-metals produce acidic oxides when dissolve in water. On the other hand, most metals, give rise to basic oxides. You will be learning more about these metal oxides in the next section.

QUESTIONS

1. Give an example of a metal which
   (i) is a liquid at room temperature.
   (ii) can be easily cut with a knife.
   (iii) is the best conductor of heat.
   (iv) is a poor conductor of heat.
2. Explain the meanings of malleable and ductile.

3.2 CHEMICAL PROPERTIES OF METALS

We will learn about the chemical properties of metals in the following Sections 3.2.1 to 3.2.4. For this, collect the samples of following metals – aluminium, copper, iron, lead, magnesium, zinc and sodium.

3.2.1 What happens when Metals are burnt in Air?

You have seen in Activity 3.8 that magnesium burns in air with a dazzling white flame. Do all metals react in the same manner? Let us check by performing the following Activity.

Activity 3.9

CAUTION: The following activity needs the teacher's assistance. It would be better if students wear eye protection.

- Hold any of the samples taken above with a pair of tongs and try burning over a flame. Repeat with the other metal samples.
- Collect the product if formed.
- Let the products and the metal surface cool down.
- Which metals burn easily?
- What flame colour did you observe when the metal burnt?
- How does the metal surface appear after burning?
- Arrange the metals in the decreasing order of their reactivity towards oxygen.
- Are the products soluble in water?

Almost all metals combine with oxygen to form metal oxides.

$$\text{Metal} + \text{Oxygen} \rightarrow \text{Metal oxide}$$

For example, when copper is heated in air, it combines with oxygen to form copper(II) oxide, a black oxide.

$$\underset{\text{(Copper)}}{2Cu} + O_2 \rightarrow \underset{\text{(Copper(II) oxide)}}{2CuO}$$

Similarly, aluminium forms aluminium oxide.

$$\underset{\text{(Aluminium)}}{4Al} + 3O_2 \rightarrow \underset{\text{(Aluminium oxide)}}{2Al_2O_3}$$

Recall from Chapter 2, how copper oxide reacts with hydrochloric acid. We have learnt that metal oxides are basic in nature. But some metal oxides, such as aluminium oxide, zinc oxide show both acidic as well as basic behaviour. Such metal oxides which react with both acids as well as bases to produce salts and water are known as amphoteric oxides. Aluminium oxide reacts in the following manner with acids and bases –

$$Al_2O_3 + 6HCl \rightarrow 2AlCl_3 + 3H_2O$$

$$Al_2O_3 + 2NaOH \rightarrow \underset{\text{(Sodium aluminate)}}{2NaAlO_2} + H_2O$$

Most metal oxides are insoluble in water but some of these dissolve in water to form alkalis. Sodium oxide and potassium oxide dissolve in water to produce alkalis as follows –

$$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$$

$$K_2O(s) + H_2O(l) \rightarrow 2KOH(aq)$$

We have observed in Activity 3.9 that all metals do not react with oxygen at the same rate. Different metals show different reactivities towards oxygen. Metals such as potassium and sodium react so vigorously that they catch fire if kept in the open. Hence, to protect them and to prevent accidental fires, they are kept immersed in kerosene oil. At ordinary temperature, the surfaces of metals such as magnesium, aluminium, zinc, lead, etc., are covered with a thin layer of oxide. The protective oxide layer prevents the metal from further oxidation. Iron does not burn on heating but iron filings burn vigorously when sprinkled in the flame of the burner. Copper does not burn, but the hot metal is coated with a black coloured layer of copper(II) oxide. Silver and gold do not react with oxygen even at high temperatures.

Do You Know?

Anodising is a process of forming a thick oxide layer of aluminium. Aluminium develops a thin oxide layer when exposed to air. This aluminium oxide coat makes it resistant to further corrosion. The resistance can be improved further by making the oxide layer thicker. During anodising, a clean aluminium article is made the anode and is electrolysed with dilute sulphuric acid. The oxygen gas evolved at the anode reacts with aluminium to make a thicker protective oxide layer. This oxide layer can be dyed easily to give aluminium articles an attractive finish.

After performing Activity 3.9, you must have observed that sodium is the most reactive of the samples of metals taken here. The reaction of magnesium is less vigorous implying that it is not as reactive as sodium. But burning in oxygen does not help us to decide about the reactivity of zinc, iron, copper or lead. Let us see some more reactions to arrive at a conclusion about the order of reactivity of these metals.

3.2.2 What happens when Metals react with Water?

Activity 3.10

CAUTION: This Activity needs the teacher's assistance.

- Collect the samples of the same metals as in Activity 3.9.
- Put small pieces of the samples separately in beakers half-filled with cold water.
- Which metals reacted with cold water? Arrange them in the increasing order of their reactivity with cold water.
- Did any metal produce fire on water?
- Does any metal start floating after some time?
- Put the metals that did not react with cold water in beakers half-filled with hot water.
- For the metals that did not react with hot water, arrange the apparatus as shown in Fig. 3.3 and observe their reaction with steam.
- Which metals did not react even with steam?
- Arrange the metals in the decreasing order of reactivity with water.

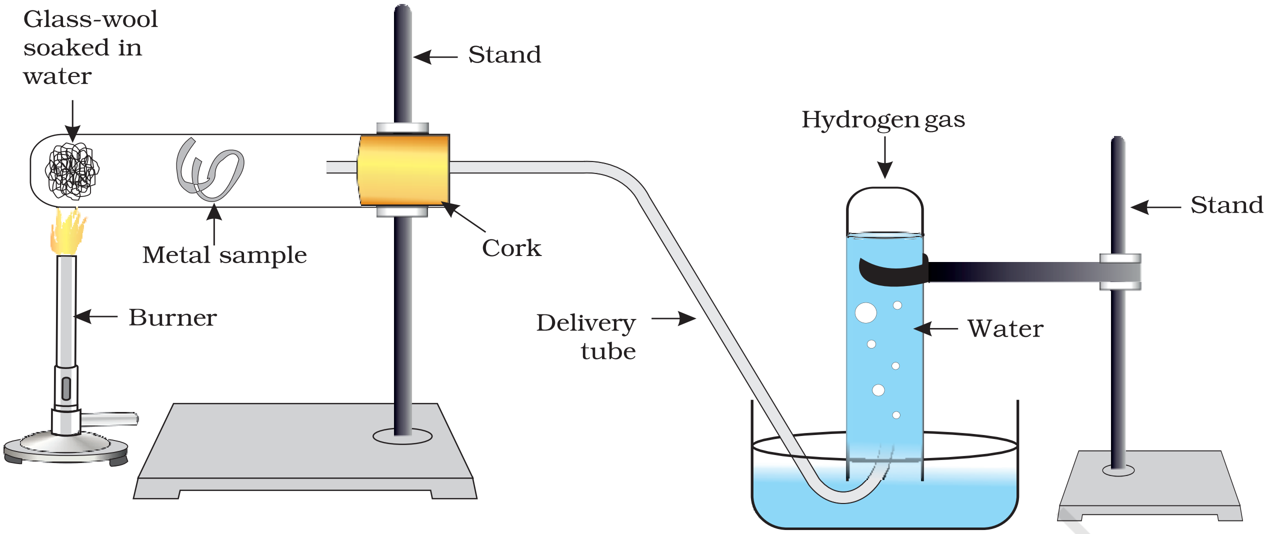

Figure 3.3 *Action of steam on a metal*

Metals react with water and produce a metal oxide and hydrogen gas. Metal oxides that are soluble in water dissolve in it to further form metal hydroxide. But all metals do not react with water.

Metal + Water → Metal oxide + Hydrogen
Metal oxide + Water → Metal hydroxide

Metals like potassium and sodium react violently with cold water. In case of sodium and potassium, the reaction is so violent and exothermic that the evolved hydrogen immediately catches fire.

$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g) + \text{heat energy}$
$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g) + \text{heat energy}$

The reaction of calcium with water is less violent. The heat evolved is not sufficient for the hydrogen to catch fire.

$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

Calcium starts floating because the bubbles of hydrogen gas formed stick to the surface of the metal.

Magnesium does not react with cold water. It reacts with hot water to form magnesium hydroxide and hydrogen. It also starts floating due to the bubbles of hydrogen gas sticking to its surface.

Metals like aluminium, iron and zinc do not react either with cold or hot water. But they react with steam to form the metal oxide and hydrogen.

$2Al(s) + 3H_2O(g) \rightarrow Al_2O_3(s) + 3H_2(g)$
$3Fe(s) + 4H_2O(g) \rightarrow Fe_3O_4(s) + 4H_2(g)$

Metals such as lead, copper, silver and gold do not react with water at all.

3.2.3 What happens when Metals react with Acids?

You have already learnt that metals react with acids to give a salt and hydrogen gas.

Metal + Dilute acid → Salt + Hydrogen

But do all metals react in the same manner? Let us find out.

Activity 3.11

- Collect all the metal samples except sodium and potassium again. If the samples are tarnished, rub them clean with sand paper. CAUTION: Do not take sodium and potassium as they react vigorously even with cold water.
- Put the samples separately in test tubes containing dilute hydrochloric acid.
- Suspend thermometers in the test tubes, so that their bulbs are dipped in the acid.
- Observe the rate of formation of bubbles carefully.
- Which metals reacted vigorously with dilute hydrochloric acid?
- With which metal did you record the highest temperature?
- Arrange the metals in the decreasing order of reactivity with dilute acids.

Write equations for the reactions of magnesium, aluminium, zinc and iron with dilute hydrochloric acid.

Hydrogen gas is not evolved when a metal reacts with nitric acid. It is because HNO_3 is a strong oxidising agent. It oxidises the H_2 produced to water and itself gets reduced to any of the nitrogen oxides (N_2O, NO, NO_2). But magnesium (Mg) and manganese (Mn) react with very dilute HNO_3 to evolve H_2 gas.

You must have observed in Activity 3.11, that the rate of formation of bubbles was the fastest in the case of magnesium. The reaction was also the most exothermic in this case. The reactivity decreases in the order Mg > Al > Zn > Fe. In the case of copper, no bubbles were seen and the temperature also remained unchanged. This shows that copper does not react with dilute HCl.

Do You Know?

Aqua regia, (Latin for 'royal water') is a freshly prepared mixture of concentrated hydrochloric acid and concentrated nitric acid in the ratio of 3:1. It can dissolve gold, even though neither of these acids can do so alone. *Aqua regia* is a highly corrosive, fuming liquid. It is one of the few reagents that is able to dissolve gold and platinum.

3.2.4 How do Metals react with Solutions of other Metal Salts?

Activity 3.12

- Take a clean wire of copper and an iron nail.
- Put the copper wire in a solution of iron sulphate and the iron nail in a solution of copper sulphate taken in test tubes (Fig. 3.4).
- Record your observations after 20 minutes.

- In which test tube did you find that a reaction has occurred?
- On what basis can you say that a reaction has actually taken place?
- Can you correlate your observations for the Activities 3.9, 3.10 and 3.11?
- Write a balanced chemical equation for the reaction that has taken place.
- Name the type of reaction.

Reactive metals can displace less reactive metals from their compounds in solution or molten form.

We have seen in the previous sections that all metals are not equally reactive. We checked the reactivity of various metals with oxygen, water and acids. But all metals do not react with these reagents. So we were not able to put all the metal samples we had collected in decreasing order of their reactivity. Displacement reactions studied in Chapter 1 give better evidence about the reactivity of metals. It is simple and easy if metal A displaces metal B from its solution, it is more reactive than B.

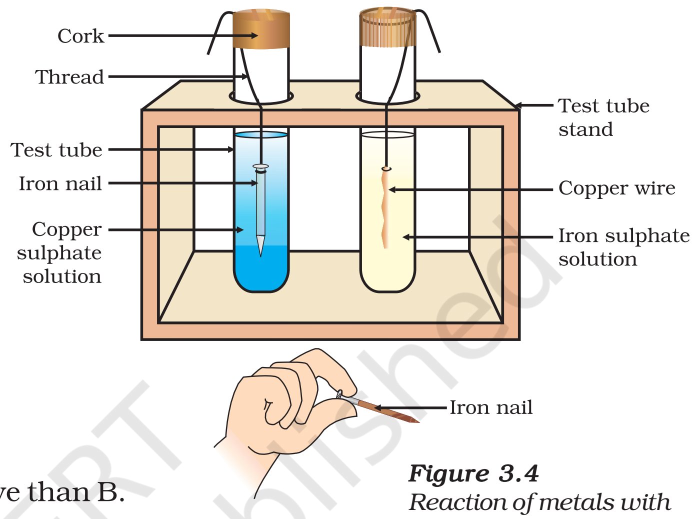

Figure 3.4
Reaction of metals with salt solutions

Metal A + Salt solution of B → Salt solution of A + Metal B

Which metal, copper or iron, is more reactive according to your observations in Activity 3.12?

3.2.5 The Reactivity Series

The reactivity series is a list of metals arranged in the order of their decreasing activities. After performing displacement experiments (Activities 1.9 and 3.12), the following series, (Table 3.2) known as the reactivity or activity series has been developed.

Table 3.2 Activity series : Relative reactivities of metals

K	Potassium	Most reactive
Na	Sodium	
Ca	Calcium	
Mg	Magnesium	
Al	Aluminium	
Zn	Zinc	Reactivity decreases
Fe	Iron	
Pb	Lead	
[H]	[Hydrogen]	
Cu	Copper	
Hg	Mercury	
Ag	Silver	
Au	Gold	Least reactive

QUESTIONS

1. Why is sodium kept immersed in kerosene oil?
2. Write equations for the reactions of
   (i) iron with steam
   (ii) calcium and potassium with water
3. Samples of four metals A, B, C and D were taken and added to the following solution one by one. The results obtained have been tabulated as follows.

Metal	Iron(II) sulphate	Copper(II) sulphate	Zinc sulphate	Silver nitrate
A	No reaction	Displacement		
B	Displacement		No reaction	
C	No reaction	No reaction	No reaction	Displacement
D	No reaction	No reaction	No reaction	No reaction

   Use the Table above to answer the following questions about metals A, B, C and D.
   (i) Which is the most reactive metal?
   (ii) What would you observe if B is added to a solution of Copper(II) sulphate?
   (iii) Arrange the metals A, B, C and D in the order of decreasing reactivity.
4. Which gas is produced when dilute hydrochloric acid is added to a reactive metal? Write the chemical reaction when iron reacts with dilute H_2SO_4.
5. What would you observe when zinc is added to a solution of iron(II) sulphate? Write the chemical reaction that takes place.

3.3 HOW DO METALS AND NON-METALS REACT?

In the above activities, you saw the reactions of metals with a number of reagents. Why do metals react in this manner? Let us recall what we learnt about the electronic configuration of elements in Class IX. We learnt that noble gases, which have a completely filled valence shell, show little chemical activity. We, therefore, explain the reactivity of elements as a tendency to attain a completely filled valence shell.

Let us have a look at the electronic configuration of noble gases and some metals and non-metals.

We can see from Table 3.3 that a sodium atom has one electron in its outermost shell. If it loses the electron from its M shell then its L shell now becomes the outermost shell and that has a stable octet. The nucleus of this atom still has 11 protons but the number of electrons has become 10, so there is a net positive charge giving us a sodium cation Na^+. On the other hand chlorine has seven electrons in its outermost shell

Table 3.3 Electronic configurations of some elements

Type of element	Element	Atomic number	Number of electrons in shells			
			K	L	M	N
Noble gases	Helium (He)	2	2			
	Neon (Ne)	10	2	8		
	Argon (Ar)	18	2	8	8	
Metals	Sodium (Na)	11	2	8	1	
	Magnesium (Mg)	12	2	8	2	
	Aluminium (Al)	13	2	8	3	
	Potassium (K)	19	2	8	8	1
	Calcium (Ca)	20	2	8	8	2
Non-metals	Nitrogen (N)	7	2	5		
	Oxygen (O)	8	2	6		
	Fluorine (F)	9	2	7		
	Phosphorus (P)	15	2	8	5	
	Sulphur (S)	16	2	8	6	
	Chlorine (Cl)	17	2	8	7	

and it requires one more electron to complete its octet. If sodium and chlorine were to react, the electron lost by sodium could be taken up by chlorine. After gaining an electron, the chlorine atom gets a unit negative charge, because its nucleus has 17 protons and there are 18 electrons in its K, L and M shells. This gives us a chloride anion Cl^-. So both these elements can have a give-and-take relation between them as follows (Fig. 3.5).

$$\underset{2,8,1}{Na} \rightarrow \underset{\substack{2,8\\(\text{Sodium cation})}}{Na^+} + e^-$$

$$\underset{2,8,7}{Cl} + e^- \rightarrow \underset{\substack{2,8,8\\(\text{Chloride anion})}}{Cl^-}$$

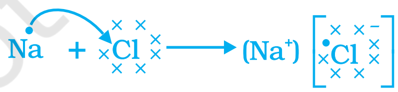

Figure 3.5 *Formation of sodium chloride*

Sodium and chloride ions, being oppositely charged, attract each other and are held by strong electrostatic forces of attraction to exist as sodium chloride (NaCl). It should be noted that sodium chloride does not exist as molecules but aggregates of oppositely charged ions.

Let us see the formation of one more ionic compound, magnesium chloride (Fig. 3.6).

$$Mg \longrightarrow Mg^{2+} + 2e^-$$

2,8,2 → 2,8 (Magnesium cation)

$$Cl + e^- \longrightarrow Cl^-$$

2,8,7 → 2,8,8 (Chloride anion)

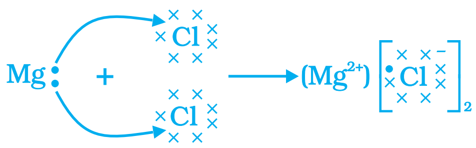

Figure 3.6 *Formation of magnesium chloride*

The compounds formed in this manner by the transfer of electrons from a metal to a non-metal are known as ionic compounds or electrovalent compounds. Can you name the cation and anion present in $MgCl_2$?

3.3.1 Properties of Ionic Compounds

To learn about the properties of ionic compounds, let us perform the following Activity:

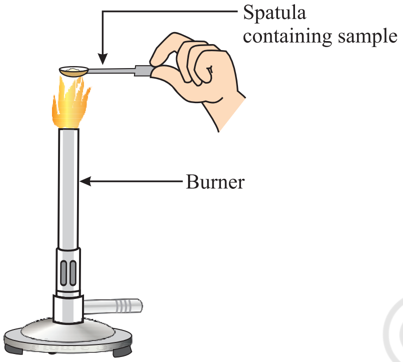

Figure 3.7
Heating a salt sample on a spatula

Activity 3.13

- Take samples of sodium chloride, potassium iodide, barium chloride or any other salt from the science laboratory.
- What is the physical state of these salts?
- Take a small amount of a sample on a metal spatula and heat directly on the flame (Fig. 3.7). Repeat with other samples.
- What did you observe? Did the samples impart any colour to the flame? Do these compounds melt?
- Try to dissolve the samples in water, petrol and kerosene. Are they soluble?
- Make a circuit as shown in Fig. 3.8 and insert the electrodes into a solution of one salt. What did you observe? Test the other salt samples too in this manner.
- What is your inference about the nature of these compounds?

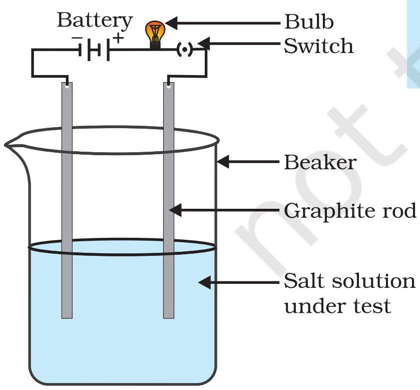

Figure 3.8
Testing the conductivity of a salt solution

Table 3.4 Melting and boiling points of some ionic compounds

Ionic compound	Melting point (K)	Boiling point (K)
NaCl	1074	1686
LiCl	887	1600
$CaCl_2$	1045	1900
CaO	2850	3120
$MgCl_2$	981	1685

You may have observed the following general properties for ionic compounds—

(i) *Physical nature*: Ionic compounds are solids and are somewhat hard because of the strong force of attraction between the positive and negative ions. These compounds are generally brittle and break into pieces when pressure is applied.

(ii) *Melting and Boiling points*: Ionic compounds have high melting and boiling points (see Table 3.4). This is because a considerable amount of energy is required to break the strong inter-ionic attraction.

(iii) *Solubility*: Electrovalent compounds are generally soluble in water and insoluble in solvents such as kerosene, petrol, etc.

(iv) *Conduction of Electricity*: The conduction of electricity through a solution involves the movement of charged particles. A solution of an ionic compound in water contains ions, which move to the opposite electrodes when electricity is passed through the solution. Ionic compounds in the solid state do not conduct electricity because movement of ions in the solid is not possible due to their rigid structure. But ionic compounds conduct electricity in the molten state. This is possible in the molten state since the elecrostatic forces of attraction between the oppositely charged ions are overcome due to the heat. Thus, the ions move freely and conduct electricity.

QUESTIONS

1. (i) Write the electron-dot structures for sodium, oxygen and magnesium.
   (ii) Show the formation of Na_2O and MgO by the transfer of electrons.
   (iii) What are the ions present in these compounds?
2. Why do ionic compounds have high melting points?

3.4 OCCURRENCE OF METALS

The earth's crust is the major source of metals. Seawater also contains some soluble salts such as sodium chloride, magnesium chloride, etc. The elements or compounds, which occur naturally in the earth's crust, are known as minerals. At some places, minerals contain a very high percentage of a particular metal and the metal can be profitably extracted from it. These minerals are called ores.

3.4.1 Extraction of Metals

You have learnt about the reactivity series of metals. Having this knowledge, you can easily understand how a metal is extracted from its ore. Some metals are found in the earth's crust in the free state. Some are found in the form of their compounds. The metals at the bottom of the activity series are the least reactive. They are often found in a free

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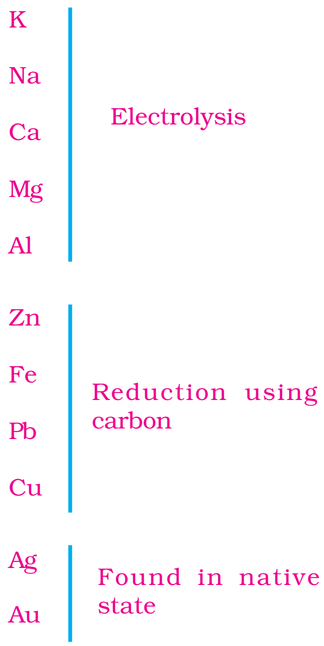

Figure 3.9
Activity series and related metallurgy

state. For example, gold, silver, platinum and copper are found in the free state. Copper and silver are also found in the combined state as their sulphide or oxide ores. The metals at the top of the activity series (K, Na, Ca, Mg and Al) are so reactive that they are never found in nature as free elements. The metals in the middle of the activity series (Zn, Fe, Pb, etc.) are moderately reactive. They are found in the earth's crust mainly as oxides, sulphides or carbonates. You will find that the ores of many metals are oxides. This is because oxygen is a very reactive element and is very abundant on the earth.

Thus on the basis of reactivity, we can group the metals into the following three categories (Fig. 3.9) – (i) Metals of low reactivity; (ii) Metals of medium reactivity; (iii) Metals of high reactivity. Different techniques are to be used for obtaining the metals falling in each category.

Several steps are involved in the extraction of pure metal from ores. A summary of these steps is given in Fig.3.10. Each step is explained in detail in the following sections.

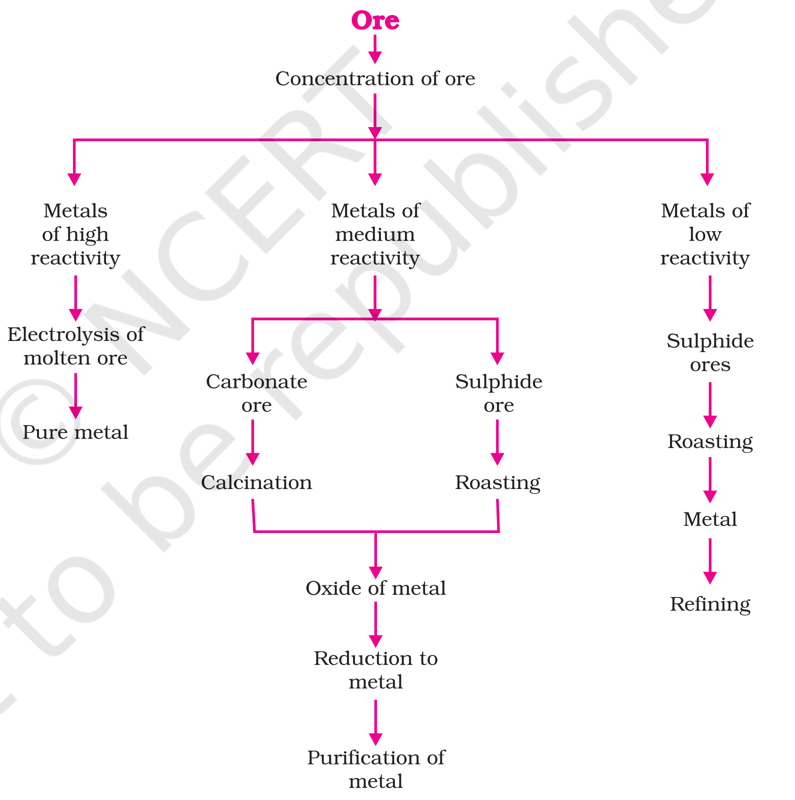

Figure 3.10 *Steps involved in the extraction of metals from ores*

3.4.2 Enrichment of Ores

Ores mined from the earth are usually contaminated with large amounts of impurities such as soil, sand, etc., called gangue. The impurities must be removed from the ore prior to the extraction of the metal. The processes