

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



# CHAPTER 1

# Chemical Reactions and Equations



1064CH01

**C**onsider 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.

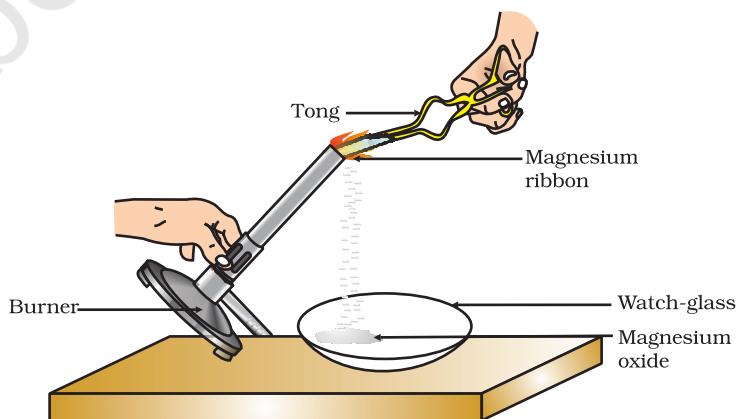
In all the above situations, the nature and the identity of the initial substance have somewhat changed. We have already learnt about physical and chemical changes of matter in our previous classes. Whenever a chemical change occurs, we can say that a chemical reaction has taken place.

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?

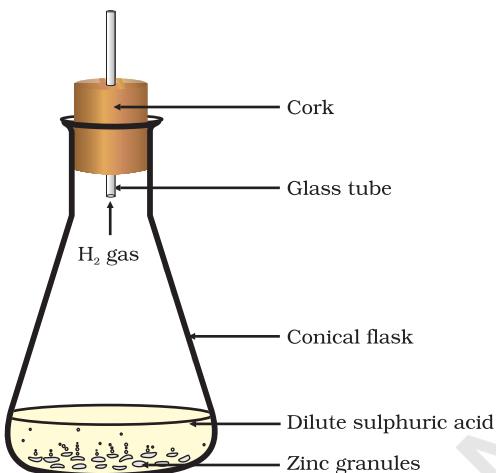


**Figure 1.1**  
Burning of a magnesium ribbon in air and collection of magnesium oxide in a watch-glass

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?



**Figure 1.2**

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

### 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?

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 –

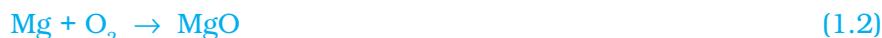


The substances that undergo chemical change in the reaction (1.1), magnesium and oxygen, are the reactants. The new substance is magnesium oxide, formed during the reaction, as a product.

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 –



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 –



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



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 –



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



**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  $\text{Fe}_3\text{O}_4$  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 $\text{H}_2\text{O}$ )	4 (in $\text{Fe}_3\text{O}_4$ )
(ii) To balance	$1 \times 4$	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  $4 \text{ H}_2\text{O}$  and not  $\text{H}_2\text{O}_4$  or  $(\text{H}_2\text{O})_4$ . Now the partly balanced equation becomes –



**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 $4 \text{ H}_2\text{O}$ )	2 (in $\text{H}_2$ )
(ii) To balance	8	$2 \times 4$

The equation would be –



**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 $\text{Fe}_3\text{O}_4$ )
(ii) To balance	$1 \times 3$	3

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



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

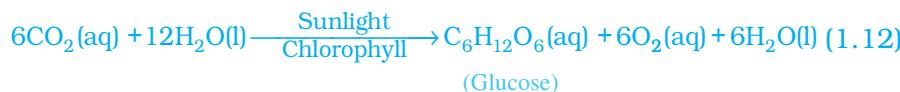
The balanced Eq. (1.9) becomes



Note that the symbol (g) is used with  $\text{H}_2\text{O}$  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 –



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

# Q U E S T I O N S

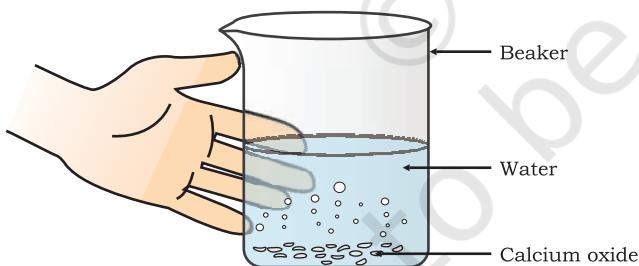
1. Why should a magnesium ribbon be cleaned before burning in air?
  2. Write the balanced equation for the following chemical reactions.
    - (i) Hydrogen + Chlorine → Hydrogen chloride
    - (ii) Barium chloride + Aluminium sulphate → Barium sulphate + Aluminium chloride
    - (iii) Sodium + Water → 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



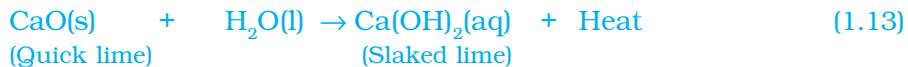
*Figure 1.3*

## 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?

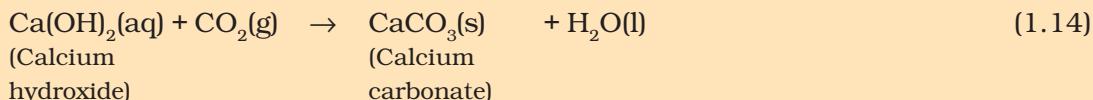
Calcium oxide reacts vigorously with water to produce slaked lime (calcium hydroxide) releasing a large amount of heat.



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  $\text{CaCO}_3$ .



Let us discuss some more examples of combination reactions.

- (i) Burning of coal



- (ii) Formation of water from  $\text{H}_2(\text{g})$  and  $\text{O}_2(\text{g})$



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.

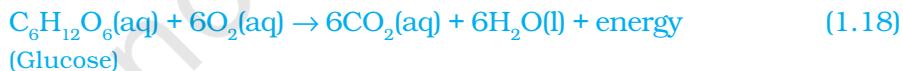
Other examples of exothermic reactions are –

- (i) Burning of natural gas



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

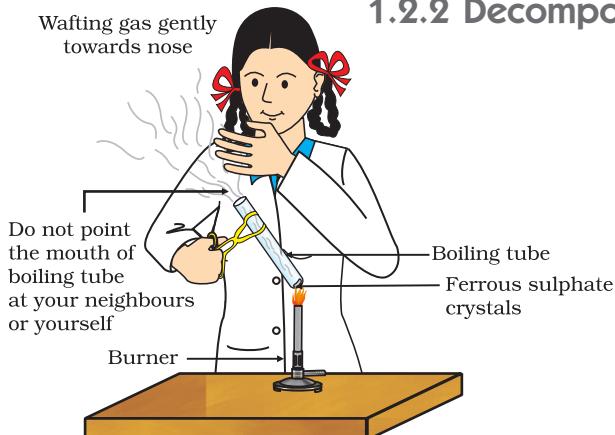
We all know that we need energy to stay alive. We get this energy from the food we eat. During digestion, food is broken down into simpler substances. For example, rice, potatoes and bread contain carbohydrates. These carbohydrates are broken down to form glucose. This glucose combines with oxygen in 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.



- (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.

## 1.2.2 Decomposition Reaction



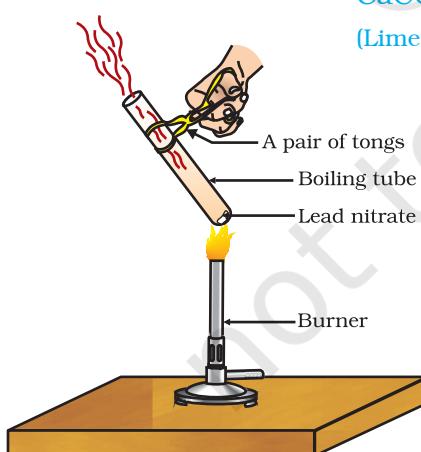
**Figure 1.4**  
Correct way of heating  
the boiling tube  
containing crystals  
of ferrous sulphate  
and of smelling the  
odour

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



In this reaction you can observe that a single reactant breaks down to give simpler products. This is a decomposition reaction. Ferrous sulphate crystals ( $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$ ) lose water when heated and the colour of the crystals changes. It then decomposes to ferric oxide ( $\text{Fe}_2\text{O}_3$ ), sulphur dioxide ( $\text{SO}_2$ ) and sulphur trioxide ( $\text{SO}_3$ ). Ferric oxide is a solid, while  $\text{SO}_2$  and  $\text{SO}_3$  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.



**Figure 1.5**  
Heating of lead nitrate and emission of nitrogen dioxide

## 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.

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

## 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 ( $\text{NO}_2$ ). The reaction that takes place is –



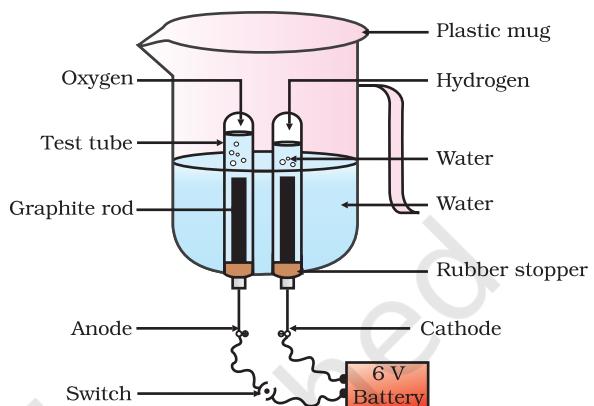
Let us perform some more decomposition reactions as given in Activities 1.7 and 1.8.

## 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.

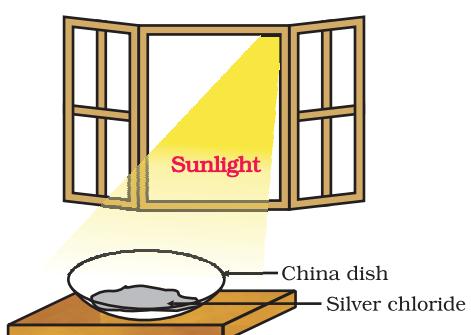
  - What happens in each case?
  - Which gas is present in each test tube?



**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*



Silver bromide also behaves in the same way.



The above reactions are used in black and white photography.

What form of energy is causing these decomposition reactions?

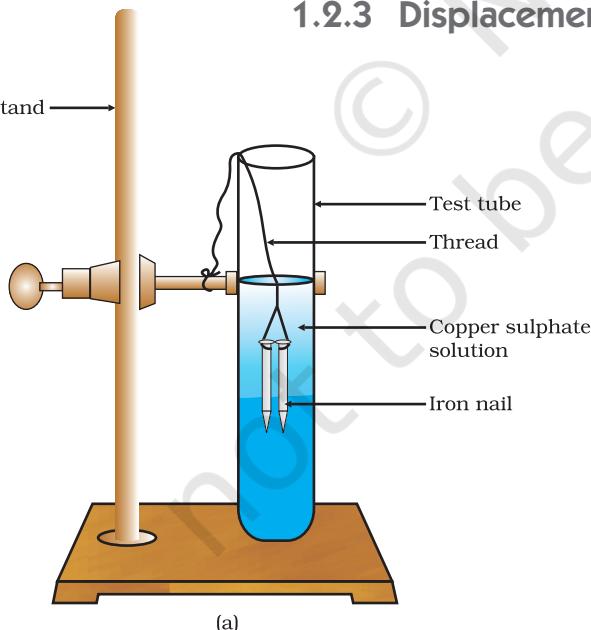
We have seen that the decomposition reactions require energy either in the form of heat, light or electricity for breaking down the reactants. Reactions in which energy is absorbed are known as endothermic reactions.

### 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?

## Q U E S T I O N S

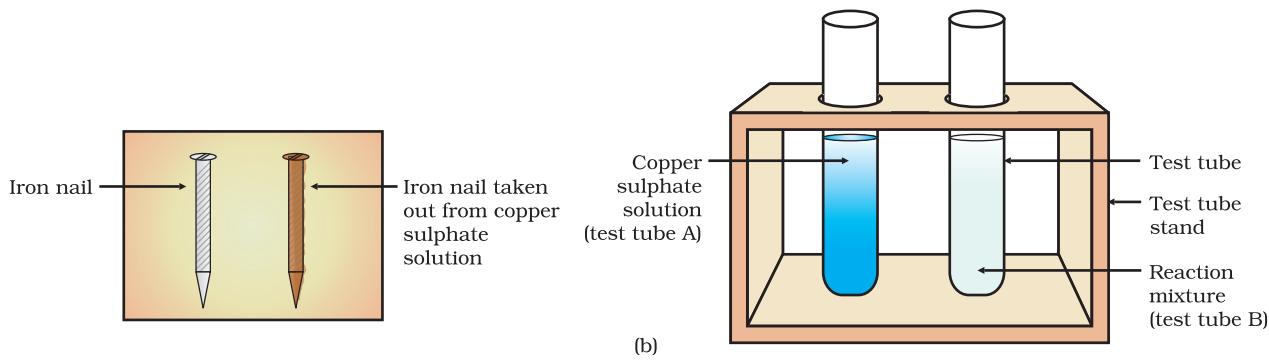
1. A solution of a substance 'X' is used for whitewashing.
  - (i) Name the substance 'X' and write its formula.
  - (ii) Write the reaction of the substance 'X' named in (i) above with water.
2. Why is the amount of gas collected in one of the test tubes in Activity 1.7 double of the amount collected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



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

### Activity 1.9

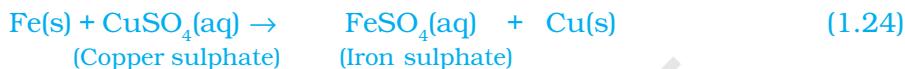
- Take three iron nails and clean them by rubbing with sand paper.
- Take two test tubes marked as (A) and (B). In each test tube, take about 10 mL copper sulphate solution.
- Tie two iron nails with a thread and immerse them carefully in the copper sulphate solution in test tube B for about 20 minutes [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

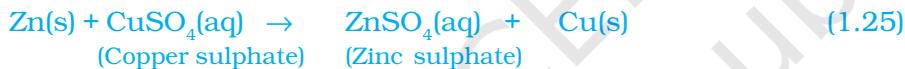
Why does the iron nail become brownish in colour and the blue colour of copper sulphate solution fades?

The following chemical reaction takes place in this Activity—



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



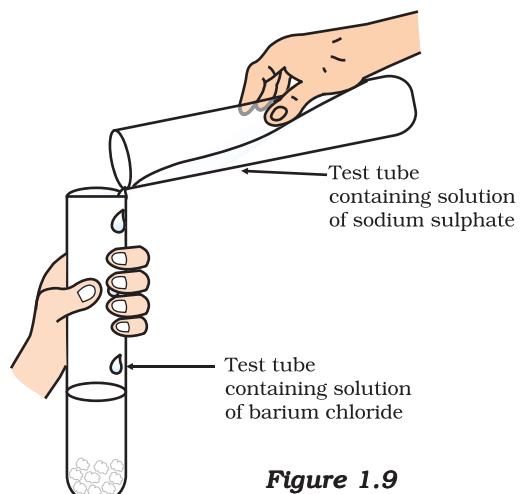
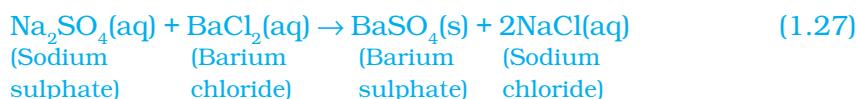
Zinc and lead are more reactive elements than copper. They displace copper from its compounds.

## 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?

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.



## **Figure 1.9**

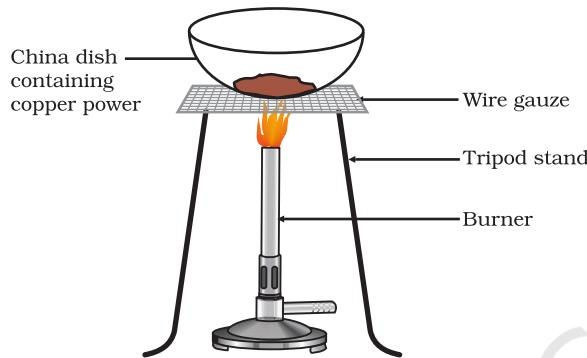
### *Formation of barium sulphate and sodium chloride*

What causes this? The white precipitate of  $\text{BaSO}_4$  is formed by the reaction of  $\text{SO}_4^{2-}$  and  $\text{Ba}^{2+}$ . 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.

- What was the colour of the precipitate formed? Can you name the compound precipitated?
- Write the balanced chemical equation for this reaction.
- 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.



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



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.



Some other examples of redox reactions are:



In reaction (1.31) carbon is oxidised to CO and ZnO is reduced to Zn. In reaction (1.32) HCl is oxidised to Cl<sub>2</sub> whereas MnO<sub>2</sub> is reduced to MnCl<sub>2</sub>.

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)  $4\text{Na(s)} + \text{O}_2\text{(g)} \rightarrow 2\text{Na}_2\text{O(s)}$
  - (ii)  $\text{CuO(s)} + \text{H}_2\text{(g)} \rightarrow \text{Cu(s)} + \text{H}_2\text{O(l)}$



## What you have learnt

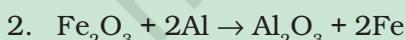
- A complete chemical equation represents the reactants, products and their physical states symbolically.
- A chemical equation is balanced so that the numbers of atoms of each type involved in a chemical reaction are the same on the reactant and product sides of the equation. Equations must always be 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?



- (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



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 filings? Tick the correct answer.
- (a) Hydrogen gas and iron chloride are produced.  
(b) Chlorine gas and iron hydroxide are produced.  
(c) No reaction takes place.  
(d) Iron salt and water are produced.
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 sulphur 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)  $\text{HNO}_3 + \text{Ca}(\text{OH})_2 \rightarrow \text{Ca}(\text{NO}_3)_2 + \text{H}_2\text{O}$   
(b)  $\text{NaOH} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + \text{H}_2\text{O}$   
(c)  $\text{NaCl} + \text{AgNO}_3 \rightarrow \text{AgCl} + \text{NaNO}_3$   
(d)  $\text{BaCl}_2 + \text{H}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + \text{HCl}$
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

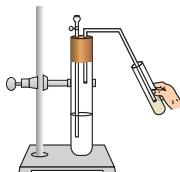
Perform the following activity.

- 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	Phenolphthalein 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?

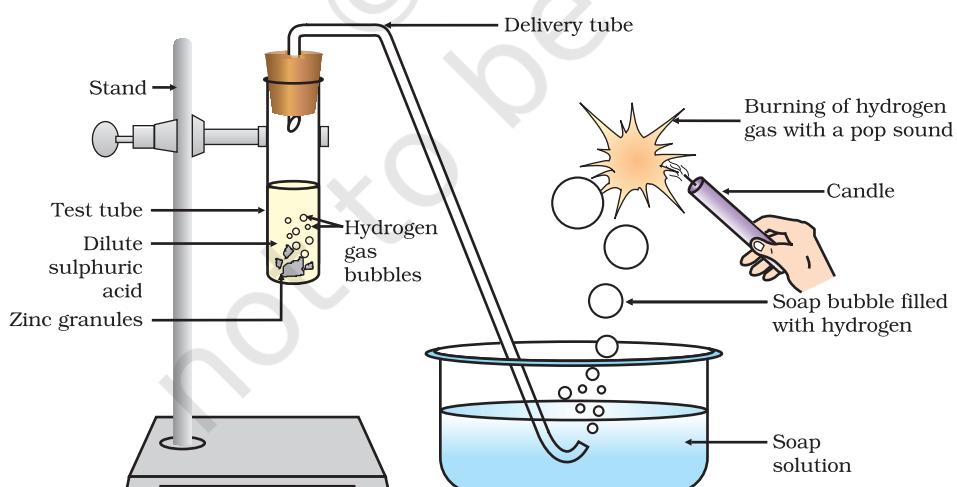
Let us do some more activities to understand the chemical properties of acids and bases.

## 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,  $\text{HNO}_3$  and  $\text{CH}_3\text{COOH}$ .
- 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 –

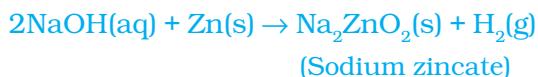
Acid + Metal → Salt + Hydrogen gas

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

## Activity 2.4

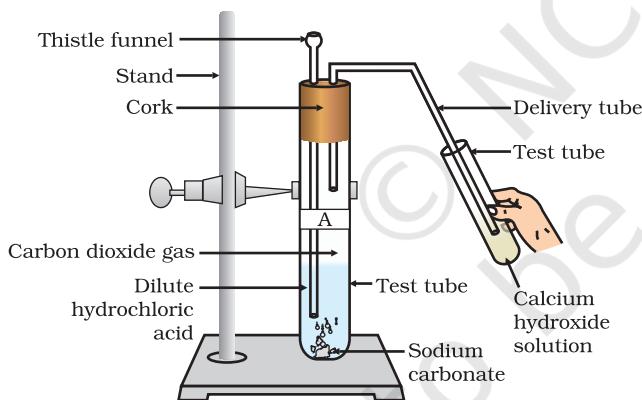
- Place a few pieces of granulated zinc metal in a test tube.
  - Add 2 mL of sodium hydroxide solution and warm the contents of the test tube.
  - Repeat the rest of the steps as in Activity 2.3 and record your observations.

The reaction that takes place can be written as follows.



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?



**Figure 2.2**  
Passing carbon dioxide gas through calcium hydroxide solution

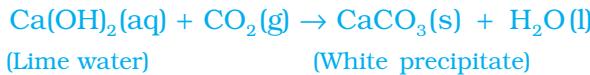
## Activity 2.5

- Take two test tubes, label them as A and B.
  - Take about 0.5 g of sodium carbonate ( $\text{Na}_2\text{CO}_3$ ) in test tube A and about 0.5 g of sodium hydrogencarbonate ( $\text{NaHCO}_3$ ) 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.

The reactions occurring in the above Activity are written as –



On passing the carbon dioxide gas evolved through lime water,



On passing excess carbon dioxide the following reaction takes place:



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 –

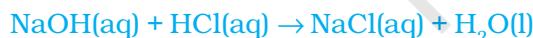


## 2.1.4 How do Acids and Bases React with each other?

### 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 –



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 –



## 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 –



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

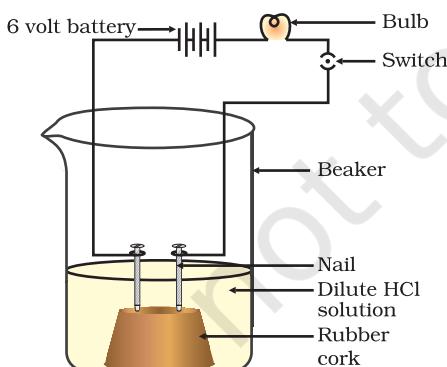
You saw the reaction between carbon dioxide and calcium hydroxide (lime water) in Activity 2.5. Calcium hydroxide, which is a base, reacts with carbon dioxide to produce a salt and water. Since this is similar to the reaction between a base and an acid, we can conclude that non-metallic oxides are acidic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

## Q U E S T I O N S

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?

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<sup>+</sup> ion as cation and anion such as Cl<sup>-</sup> in HCl, NO<sub>3</sub><sup>-</sup> in HNO<sub>3</sub>, SO<sub>4</sub><sup>2-</sup> in H<sub>2</sub>SO<sub>4</sub>, CH<sub>3</sub>COO<sup>-</sup> in CH<sub>3</sub>COOH. Since the cation present in acids is H<sup>+</sup>, this suggests that acids produce hydrogen ions, H<sup>+</sup>(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:
  - dry HCl gas
  - HCl solution?

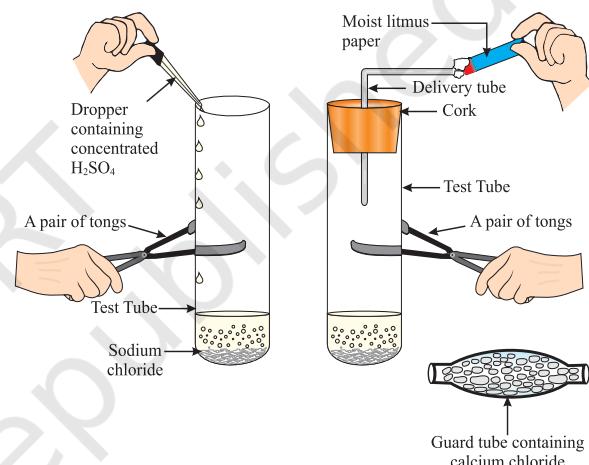


Figure 2.4 Preparation of HCl gas

**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<sup>+</sup> ion from HCl molecules cannot occur in the absence of water.



Hydrogen ions cannot exist alone, but they exist after combining with water molecules. Thus hydrogen ions must always be shown as H<sup>+</sup>(aq) or hydronium ion (H<sub>3</sub>O<sup>+</sup>).



We have seen that acids give H<sub>3</sub>O<sup>+</sup> or H<sup>+</sup>(aq) ion in water. Let us see what happens when a base is dissolved in water.



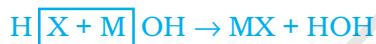


Bases generate hydroxide ( $\text{OH}^-$ ) ions in water. Bases which are soluble in water are called alkalis.

## 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  $\text{H}^+(\text{aq})$  and all bases generate  $\text{OH}^-(\text{aq})$ , we can view the neutralisation reaction as follows –



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  $\text{H}_2\text{SO}_4$  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 ( $\text{H}_3\text{O}^+/\text{OH}^-$ ) 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<sub>3</sub>, 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<sub>3</sub>O<sup>+</sup>) affected when a solution of an acid is diluted?
6. How is the concentration of hydroxide ions (OH<sup>-</sup>) affected when excess base is dissolved in a solution of sodium hydroxide?



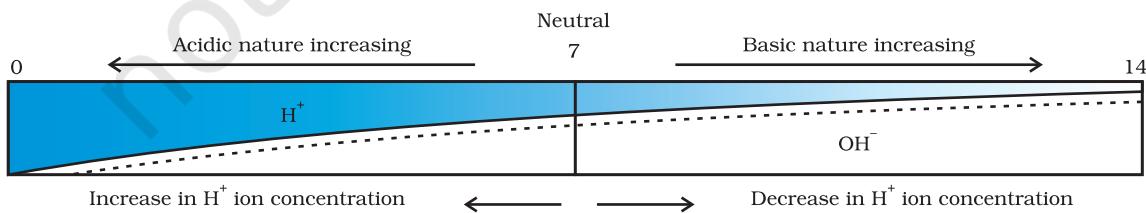
### 2.3 HOW STRONG ARE ACID OR BASE SOLUTIONS?

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<sup>+</sup> or OH<sup>-</sup> 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<sup>-</sup> 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.

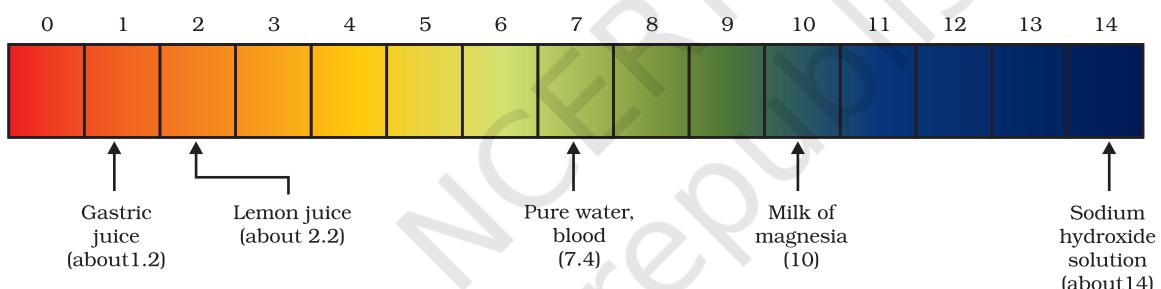


**Figure 2.6** Variation of pH with the change in concentration of H<sup>+</sup>(aq) and OH<sup>-</sup>(aq) ions

**Table 2.2****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?

S. No.	Solution	Colour of pH paper	Approximate 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.

### **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.

## 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

## Q U E S T I O N S

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  $\text{H}^+(\text{aq})$  ions have on the nature of the solution?
3. Do basic solutions also have  $\text{H}^+(\text{aq})$  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<sub>2</sub>SO<sub>4</sub> 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?



### 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.



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.

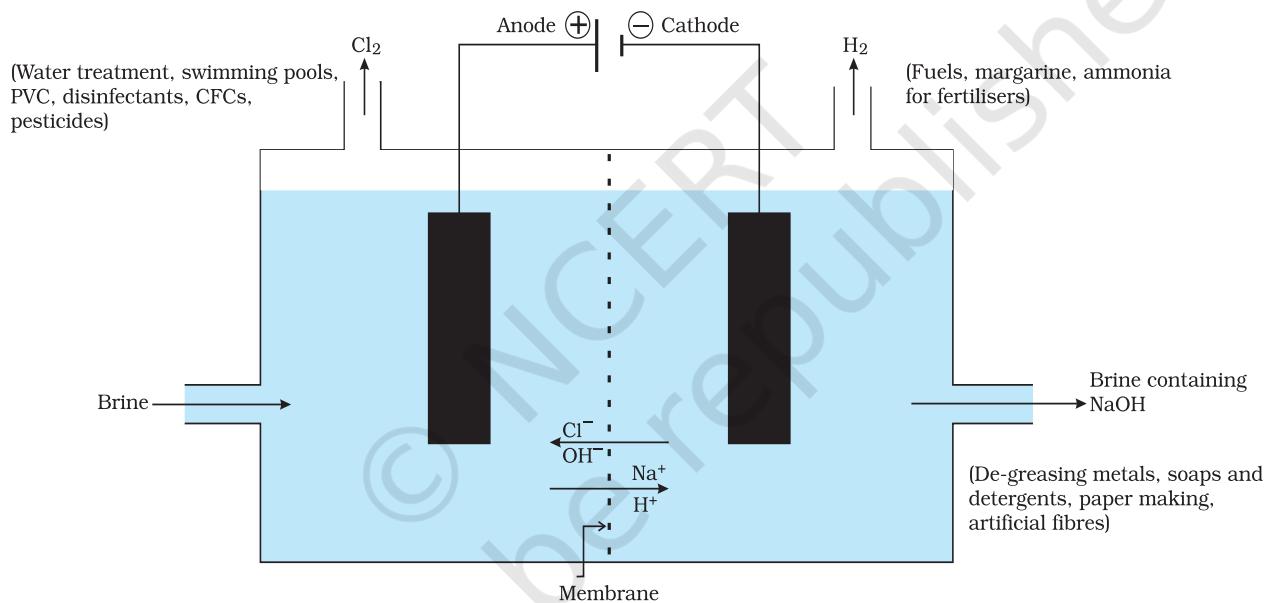


Figure 2.8 Important products from the chlor-alkali process

### 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 [ $\text{Ca}(\text{OH})_2$ ]. Bleaching powder is represented as  $\text{Ca}(\text{ClO})_2$ , though the actual composition is quite complex.

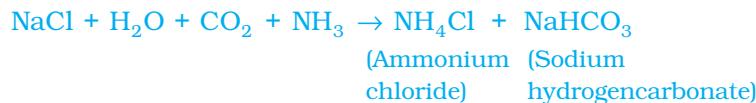


### **Bleaching powder is used –**

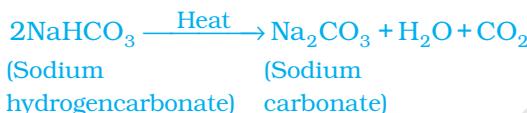
- (i) for bleaching cotton and linen in the textile industry, for bleaching wood pulp in paper factories and for bleaching washed clothes in laundry;
- (ii) as an oxidising agent in many chemical industries; and
- (iii) to make drinking water free from germs.

### **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 ( $\text{NaHCO}_3$ ). It is produced using sodium chloride as one of the raw materials.



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 –



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 –

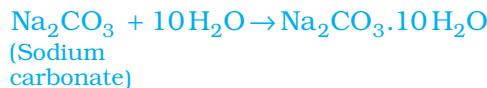


Carbon dioxide produced during the reaction can cause bread or cake to rise making them soft and spongy.

- (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  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$  (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.



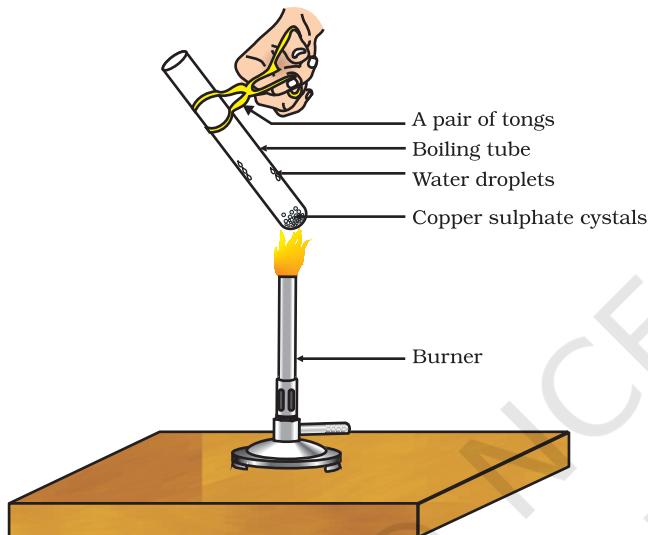
What does  $10\text{H}_2\text{O}$  signify? Does it make  $\text{Na}_2\text{CO}_3$  wet? We will address this question in the next section.

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?



### 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?

**Figure 2.9**  
Removing water of crystallisation

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  $\text{Cu SO}_4 \cdot 5\text{H}_2\text{O}$ . Now you would be able to answer the question whether the molecule of  $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$  is wet.

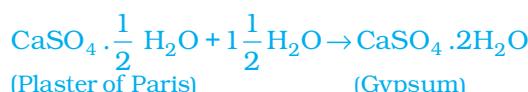
One other salt, which possesses water of crystallisation is gypsum. It has two water molecules as water of crystallisation. It has the chemical formula  $\text{CaSO}_4 \cdot 2\text{H}_2\text{O}$ . Let us look into the use of this salt.

### Plaster of Paris

On heating gypsum at 373 K, it loses water molecules and becomes

calcium sulphate hemihydrate ( $\text{CaSO}_4 \cdot \frac{1}{2} \text{H}_2\text{O}$ ). This is called Plaster of

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.



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  $\text{CaSO}_4$  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'?

## Q U E S T I O N S

1. What is the common name of the compound  $\text{Ca}(\text{ClO})_2$ ?
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  $\text{H}^+(\text{aq})$  ions in solution. Formation of  $\text{OH}^-(\text{aq})$  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.

## E X E R C I S E S

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 ( $\text{CH}_3\text{COOH}$ ) 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

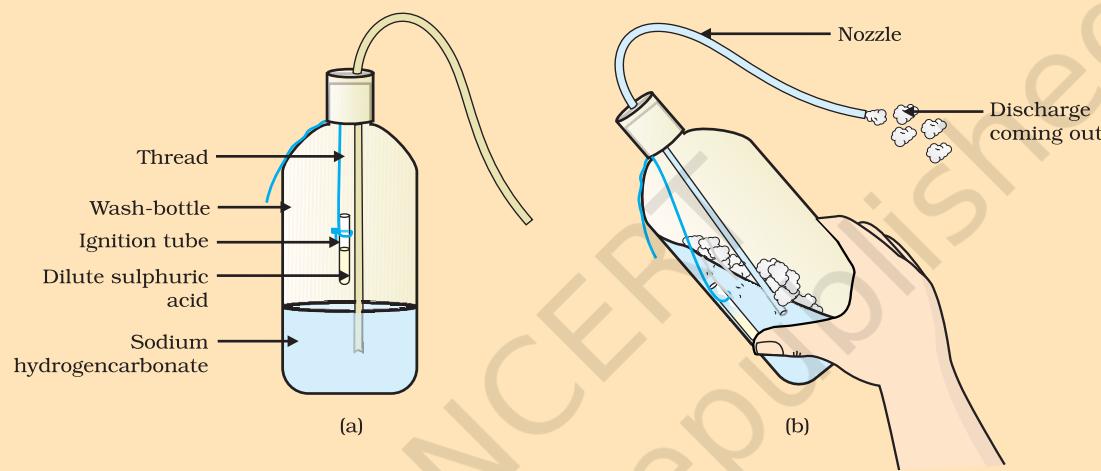
### (I) Prepare your own indicator

- 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*.

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

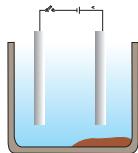
The reaction of acids with metal hydrogencarbonates is used in the fire extinguishers which produce carbon dioxide.

- Take 20 mL of sodium hydrogencarbonate ( $\text{NaHCO}_3$ ) 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



1064CH03

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?

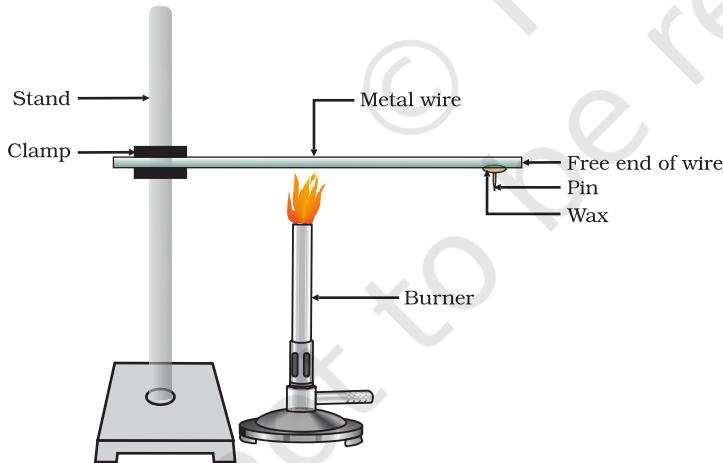
### Activity 3.4

- 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.

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 3.5

- 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?

### 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?

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.

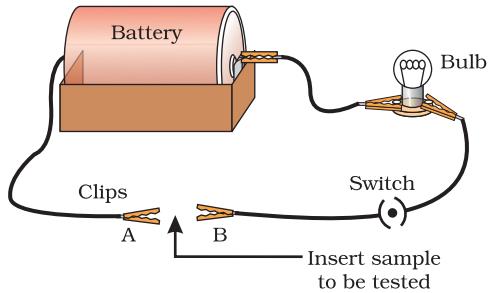
Compile your observations regarding metals and non-metals in Table 3.1.

**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 –

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



**Figure 3.2**  
Metals are good conductors of electricity.

- 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.

## Q U E S T I O N S

- 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

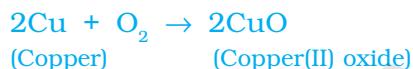
**CAUTION:** The following activity needs the teacher's assistance. It would be better if students wear eye protection.

- 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?

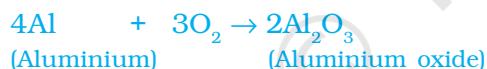
Almost all metals combine with oxygen to form metal oxides.



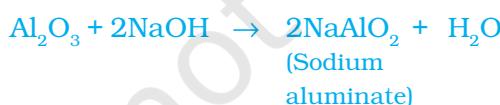
For example, when copper is heated in air, it combines with oxygen to form copper(II) oxide, a black oxide.



Similarly, aluminium forms aluminium oxide.



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 –



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 –



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.

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.

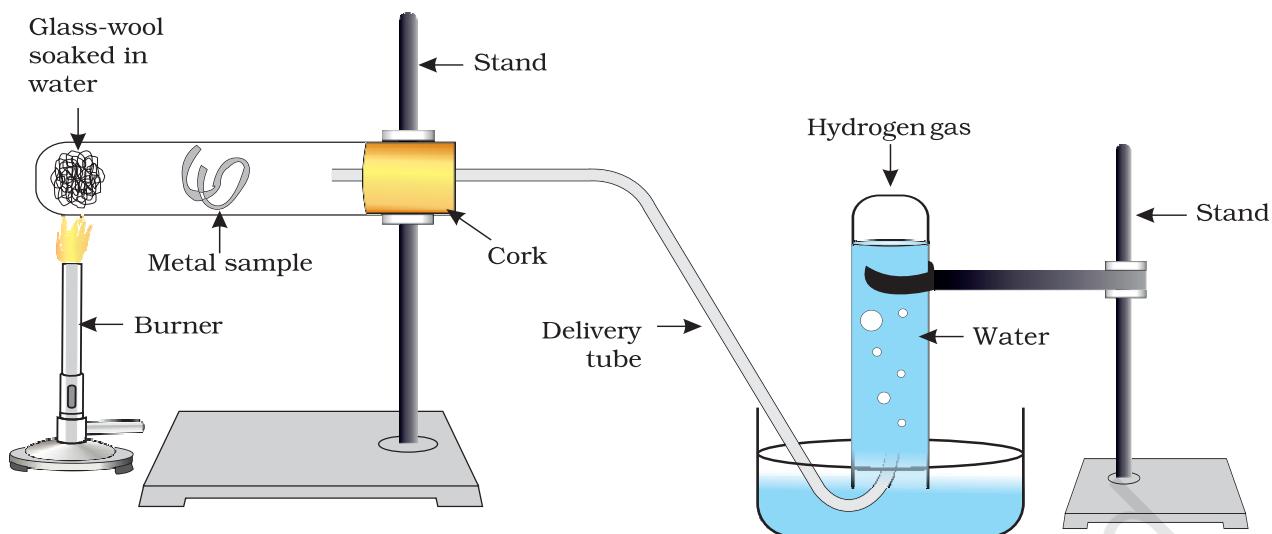
After performing Activity 3.9, you must have observed that sodium is the most reactive of the samples of metals taken here. The reaction of magnesium is less vigorous implying that it is not as reactive as sodium. But burning in oxygen does not help us to decide about the reactivity of zinc, iron, copper 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.



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.



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



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.



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  $\text{HNO}_3$  is a strong oxidising agent. It oxidises the  $\text{H}_2$  produced to water and itself gets reduced to any of the nitrogen oxides ( $\text{N}_2\text{O}$ ,  $\text{NO}$ ,  $\text{NO}_2$ ). But magnesium (Mg) and manganese (Mn) react with very dilute  $\text{HNO}_3$  to evolve  $\text{H}_2$  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  $\text{Mg} > \text{Al} > \text{Zn} > \text{Fe}$ . 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.



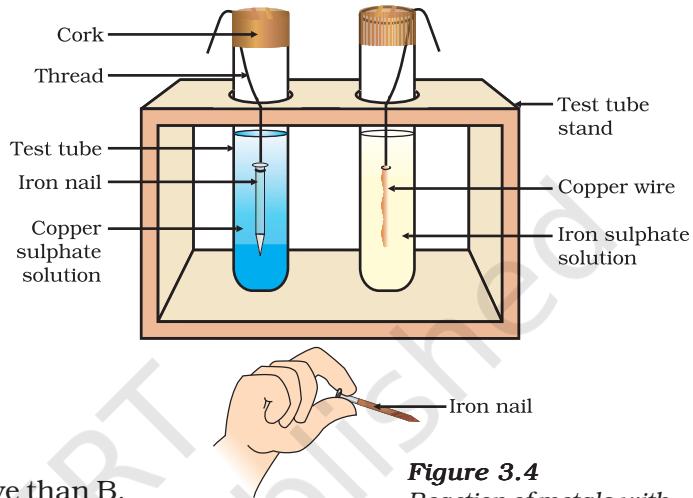
Which metal, copper or iron, is more reactive according to your observations in Activity 3.12?

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



**Figure 3.4**  
Reaction of metals with salt solutions

## Q U E S T I O N S

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  $\text{H}_2\text{SO}_4$ .
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.

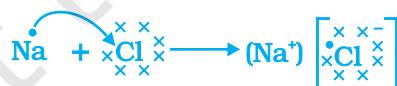
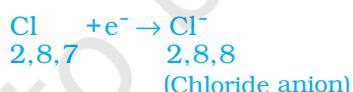
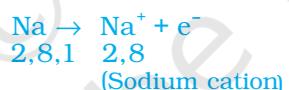
Let us have a look at the electronic configuration of noble gases and some metals and non-metals.

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  $\text{Na}^+$ . 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	

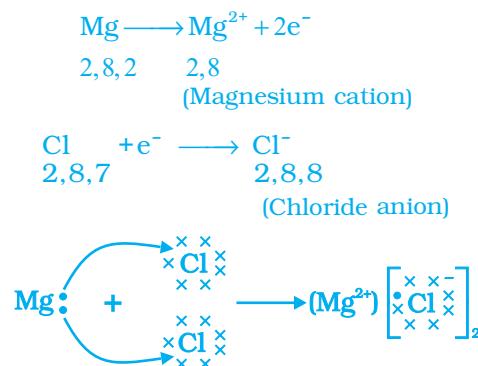
and it requires one more electron to complete its octet. If sodium and chlorine were to react, the electron lost by sodium could be taken up by chlorine. After gaining an electron, the chlorine atom gets a unit negative charge, because its nucleus has 17 protons and there are 18 electrons in its K, L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



**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 ( $\text{NaCl}$ ). 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).

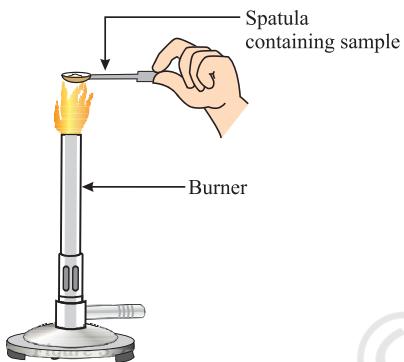


**Figure 3.6** Formation of magnesium chloride

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  $\text{MgCl}_2$ ?

### 3.3.1 Properties of Ionic Compounds

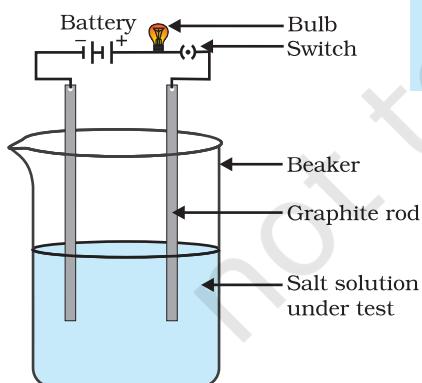
To learn about the properties of ionic compounds, let us perform the following Activity:



**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 <sub>2</sub>	1045	1900
CaO	2850	3120
MgCl <sub>2</sub>	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 electrostatic forces of attraction between the oppositely charged ions are overcome due to the heat. Thus, the ions move freely and conduct electricity.

## Q U E S T I O N S

1.
  - (i) Write the electron-dot structures for sodium, oxygen and magnesium.
  - (ii) Show the formation of  $\text{Na}_2\text{O}$  and  $\text{MgO}$  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

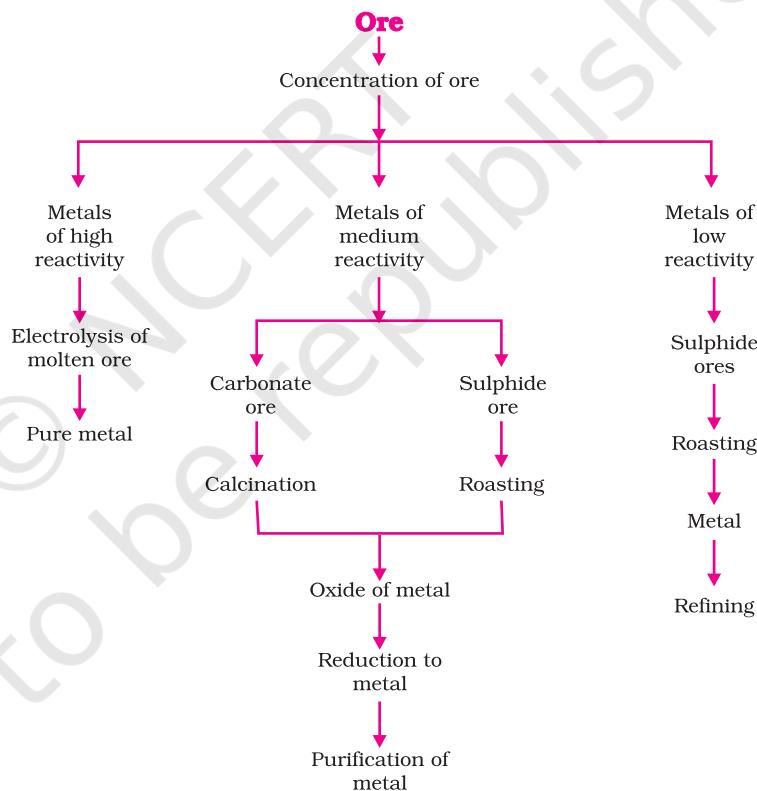
K	
Na	
Ca	Electrolysis
Mg	
Al	
Zn	
Fe	Reduction using carbon
Pb	
Cu	
Ag	Found in native state
Au	

**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

used for removing the gangue from the ore are based on the differences between the physical or chemical properties of the gangue and the ore. Different separation techniques are accordingly employed.

### 3.4.3 Extracting Metals Low in the Activity Series

Metals low in the activity series are very unreactive. The oxides of these metals can be reduced to metals by heating alone. For example, cinnabar ( $\text{HgS}$ ) is an ore of mercury. When it is heated in air, it is first converted into mercuric oxide ( $\text{HgO}$ ). Mercuric oxide is then reduced to mercury on further heating.



Similarly, copper which is found as  $\text{Cu}_2\text{S}$  in nature can be obtained from its ore by just heating in air.



### 3.4.4 Extracting Metals in the Middle of the Activity Series

The metals in the middle of the activity series such as iron, zinc, lead, copper, are moderately reactive. These are usually present as sulphides or carbonates in nature. It is easier to obtain a metal from its oxide, as compared to its sulphides and carbonates. Therefore, prior to reduction, the metal sulphides and carbonates must be converted into metal oxides. The sulphide ores are converted into oxides by heating strongly in the presence of excess air. This process is known as roasting. The carbonate ores are changed into oxides by heating strongly in limited air. This process is known as calcination. The chemical reaction that takes place during roasting and calcination of zinc ores can be shown as follows –

Roasting



Calcination



The metal oxides are then reduced to the corresponding metals by using suitable reducing agents such as carbon. For example, when zinc oxide is heated with carbon, it is reduced to metallic zinc.



You are already familiar with the process of oxidation and reduction explained in the first Chapter. Obtaining metals from their compounds is also a reduction process.

Besides using carbon (coke) to reduce metal oxides to metals, sometimes displacement reactions can also be used. The highly reactive metals such as sodium, calcium, aluminium, etc., are used as reducing

agents because they can displace metals of lower reactivity from their compounds. For example, when manganese dioxide is heated with aluminium powder, the following reaction takes place –



**Figure 3.11**

Thermit process for joining railway tracks

Can you identify the substances that are getting oxidised and reduced?

These displacement reactions are highly exothermic. The amount of heat evolved is so large that the metals are produced in the molten state. In fact, the reaction of iron(III) oxide ( $\text{Fe}_2\text{O}_3$ ) with aluminium is used to join railway tracks or cracked machine parts. This reaction is known as the thermit reaction.

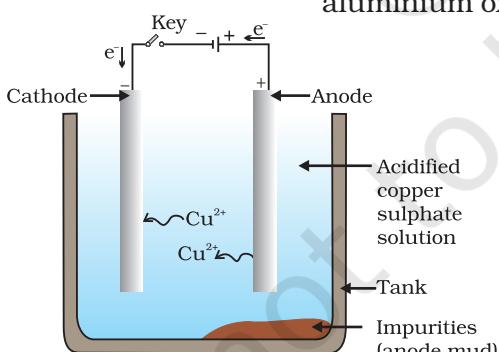


### 3.4.5 Extracting Metals towards the Top of the Activity Series

The metals high up in the reactivity series are very reactive. They cannot be obtained from their compounds by heating with carbon. For example, carbon cannot reduce the oxides of sodium, magnesium, calcium, aluminium, etc., to the respective metals. This is because these metals have more affinity for oxygen than carbon. These metals are obtained by electrolytic reduction. For example, sodium, magnesium and calcium are obtained by the electrolysis of their molten chlorides. The metals are deposited at the cathode (the negatively charged electrode), whereas, chlorine is liberated at the anode (the positively charged electrode). The reactions are –



Similarly, aluminium is obtained by the electrolytic reduction of aluminium oxide.



**Figure 3.12**

Electrolytic refining of copper. The electrolyte is a solution of acidified copper sulphate. The anode is impure copper, whereas, the cathode is a strip of pure copper. On passing electric current, pure copper is deposited on the cathode.

### 3.4.6 Refining of Metals

The metals produced by various reduction processes described above are not very pure. They contain impurities, which must be removed to obtain pure metals. The most widely used method for refining impure metals is electrolytic refining.

**Electrolytic Refining:** Many metals, such as copper, zinc, tin, nickel, silver, gold, etc., are refined electrolytically. In this process, the impure metal is made the anode and a thin strip of pure metal is made the cathode. A solution of the metal salt is used as an electrolyte. The apparatus is set up as shown in Fig. 3.12. On passing the current through the electrolyte, the pure metal from the anode dissolves into the electrolyte. An equivalent amount of pure

metal from the electrolyte is deposited on the cathode. The soluble impurities go into the solution, whereas, the insoluble impurities settle down at the bottom of the anode and are known as anode mud.

## QUESTIONS

1. Define the following terms.  
(i) Mineral      (ii) Ore      (iii) Gangue
2. Name two metals which are found in nature in the free state.
3. What chemical process is used for obtaining a metal from its oxide?



## 3.5 CORROSION

You have learnt the following about corrosion in Chapter 1 –

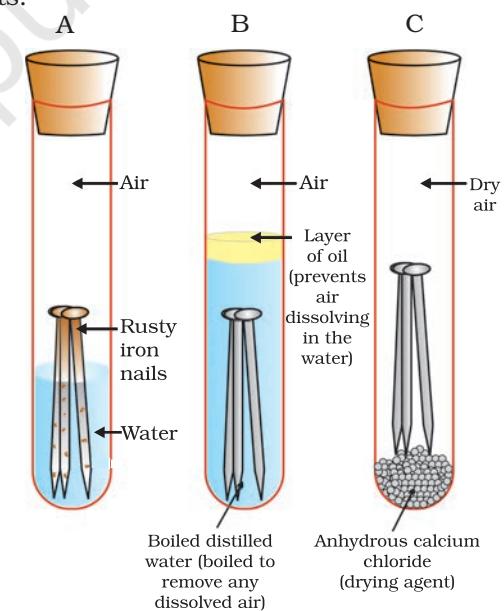
- Silver articles become black after some time when exposed to air. This is because it reacts with sulphur in the air to form a coating of silver sulphide.
- Copper reacts with moist carbon dioxide in the air and slowly loses its shiny brown surface and gains a green coat. This green substance is basic copper carbonate.
- Iron when exposed to moist air for a long time acquires a coating of a brown flaky substance called rust.

Let us find out the conditions under which iron rusts.

### Activity 3.14

- Take three test tubes and place clean iron nails in each of them.
- Label these test tubes A, B and C. Pour some water in test tube A and cork it.
- Pour boiled distilled water in test tube B, add about 1 mL of oil and cork it. The oil will float on water and prevent the air from dissolving in the water.
- Put some anhydrous calcium chloride in test tube C and cork it. Anhydrous calcium chloride will absorb the moisture, if any, from the air. Leave these test tubes for a few days and then observe (Fig. 3.13).

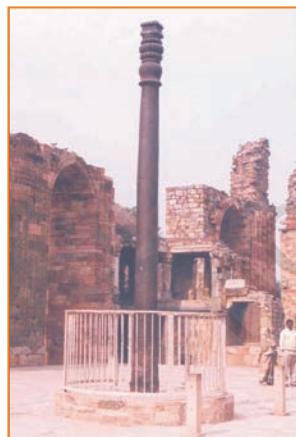
You will observe that iron nails rust in test tube A, but they do not rust in test tubes B and C. In the test tube A, the nails are exposed to both air and water. In the test tube B, the nails are exposed to only water, and the nails in test tube C are exposed to dry air. What does this tell us about the conditions under which iron articles rust?



**Figure 3.13**

Investigating the conditions under which iron rusts. In tube A, both air and water are present. In tube B, there is no air dissolved in the water. In tube C, the air is dry.

## Do You Know?



Iron pillar at Delhi

## More to Know!

The iron pillar near the Qutub Minar in Delhi was built more than 1600 years ago by the iron workers of India. They had developed a process which prevented iron from rusting. For its quality of rust resistance it has been examined by scientists from all parts of the world. The iron pillar is 8 m high and weighs 6 tonnes (6000 kg).

### 3.5.1 Prevention of Corrosion

The rusting of iron can be prevented by painting, oiling, greasing, galvanising, chrome plating, anodising or making alloys.

Galvanisation is a method of protecting steel and iron from rusting by coating them with a thin layer of zinc. The galvanised article is protected against rusting even if the zinc coating is broken. Can you reason this out?

Alloying is a very good method of improving the properties of a metal. We can get the desired properties by this method. For example, iron is the most widely used metal. But it is never used in its pure state. This is because pure iron is very soft and stretches easily when hot. But, if it is mixed with a small amount of carbon (about 0.05 %), it becomes hard and strong. When iron is mixed with nickel and chromium, we get stainless steel, which is hard and does not rust. Thus, if iron is mixed with some other substance, its properties change. In fact, the properties of any metal can be changed if it is mixed with some other substance. The substance added may be a metal or a non-metal. An alloy is a homogeneous mixture of two or more metals, or a metal and a non-metal. It is prepared by first melting the primary metal, and then, dissolving the other elements in it in definite proportions. It is then cooled to room temperature.

Pure gold, known as 24 carat gold, is very soft. It is, therefore, not suitable for making jewellery. It is alloyed with either silver or copper to make it hard. Generally, in India, 22 carat gold is used for making ornaments. It means that 22 parts of pure gold is alloyed with 2 parts of either copper or silver.

If one of the metals is mercury, then the alloy is known as an amalgam. The electrical conductivity and melting point of an alloy is less than that of pure metals. For example, brass, an alloy of copper and zinc (Cu and Zn), and bronze, an alloy of copper and tin (Cu and Sn), are not good conductors of electricity whereas copper is used for making electrical circuits. Solder, an alloy of lead and tin (Pb and Sn), has a low melting point and is used for welding electrical wires together.

## The wonder of ancient Indian metallurgy

The iron pillar near the Qutub Minar in Delhi was built more than 1600 years ago by the iron workers of India. They had developed a process which prevented iron from rusting. For its quality of rust resistance it has been examined by scientists from all parts of the world. The iron pillar is 8 m high and weighs 6 tonnes (6000 kg).

## Q U E S T I O N S

1. Metallic oxides of zinc, magnesium and copper were heated with the following metals.

Metal	Zinc	Magnesium	Copper
Zinc oxide			
Magnesium oxide			
Copper oxide			



In which cases will you find displacement reactions taking place?

2. Which metals do not corrode easily?  
3. What are alloys?

### What you have learnt

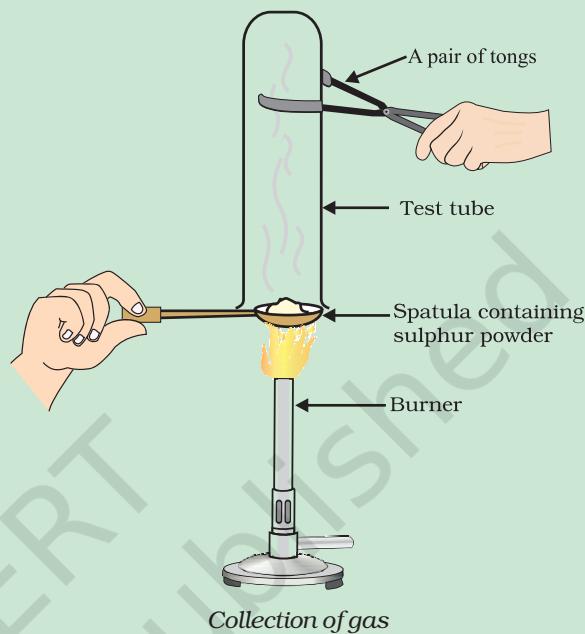
- Elements can be classified as metals and non-metals.
- Metals are lustrous, malleable, ductile and are good conductors of heat and electricity. They are solids at room temperature, except mercury which is a liquid.
- Metals can form positive ions by losing electrons to non-metals.
- Metals combine with oxygen to form basic oxides. Aluminium oxide and zinc oxide show the properties of both basic as well as acidic oxides. These oxides are known as amphoteric oxides.
- Different metals have different reactivities with water and dilute acids.
- A list of common metals arranged in order of their decreasing reactivity is known as an activity series.
- Metals above hydrogen in the Activity series can displace hydrogen from dilute acids.
- A more reactive metal displaces a less reactive metal from its salt solution.
- Metals occur in nature as free elements or in the form of their compounds.
- The extraction of metals from their ores and then refining them for use is known as metallurgy.
- An alloy is a homogeneous mixture of two or more metals, or a metal and a non-metal.
- The surface of some metals, such as iron, is corroded when they are exposed to moist air for a long period of time. This phenomenon is known as corrosion.
- Non-metals have properties opposite to that of metals. They are neither malleable nor ductile. They are bad conductors of heat and electricity, except for graphite, which conducts electricity.

- Non-metals form negatively charged ions by gaining electrons when reacting with metals.
- Non-metals form oxides which are either acidic or neutral.
- Non-metals do not displace hydrogen from dilute acids. They react with hydrogen to form hydrides.

## E X E R C I S E S

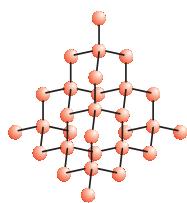
1. Which of the following pairs will give displacement reactions?
  - (a) NaCl solution and copper metal
  - (b) MgCl<sub>2</sub> solution and aluminium metal
  - (c) FeSO<sub>4</sub> solution and silver metal
  - (d) AgNO<sub>3</sub> solution and copper metal.
2. Which of the following methods is suitable for preventing an iron frying pan from rusting?
  - (a) Applying grease
  - (b) Applying paint
  - (c) Applying a coating of zinc
  - (d) All of the above.
3. An element reacts with oxygen to give a compound with a high melting point. This compound is also soluble in water. The element is likely to be
  - (a) calcium
  - (b) carbon
  - (c) silicon
  - (d) iron.
4. Food cans are coated with tin and not with zinc because
  - (a) zinc is costlier than tin.
  - (b) zinc has a higher melting point than tin.
  - (c) zinc is more reactive than tin.
  - (d) zinc is less reactive than tin.
5. You are given a hammer, a battery, a bulb, wires and a switch.
  - (a) How could you use them to distinguish between samples of metals and non-metals?
  - (b) Assess the usefulness of these tests in distinguishing between metals and non-metals.
6. What are amphoteric oxides? Give two examples of amphoteric oxides.
7. Name two metals which will displace hydrogen from dilute acids, and two metals which will not.

8. In the electrolytic refining of a metal M, what would you take as the anode, the cathode and the electrolyte?
9. Pratyush took sulphur powder on a spatula and heated it. He collected the gas evolved by inverting a test tube over it, as shown in figure below.
  - (a) What will be the action of gas on
    - (i) dry litmus paper?
    - (ii) moist litmus paper?
  - (b) Write a balanced chemical equation for the reaction taking place.
10. State two ways to prevent the rusting of iron.
11. What type of oxides are formed when non-metals combine with oxygen?
12. Give reasons
  - (a) Platinum, gold and silver are used to make jewellery.
  - (b) Sodium, potassium and lithium are stored under oil.
  - (c) Aluminium is a highly reactive metal, yet it is used to make utensils for cooking.
  - (d) Carbonate and sulphide ores are usually converted into oxides during the process of extraction.
13. You must have seen tarnished copper vessels being cleaned with lemon or tamarind juice. Explain why these sour substances are effective in cleaning the vessels.
14. Differentiate between metal and non-metal on the basis of their chemical properties.
15. A man went door to door posing as a goldsmith. He promised to bring back the glitter of old and dull gold ornaments. An unsuspecting lady gave a set of gold bangles to him which he dipped in a particular solution. The bangles sparkled like new but their weight was reduced drastically. The lady was upset but after a futile argument the man beat a hasty retreat. Can you play the detective to find out the nature of the solution he had used?
16. Give reasons why copper is used to make hot water tanks and not steel (an alloy of iron).





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## CHAPTER 4

# Carbon and its Compounds

In the last Chapter, we came to know many compounds of importance to us. In this Chapter we will study about some more interesting compounds and their properties. Also, we shall be learning about carbon, an element which is of immense significance to us in both its elemental form and in the combined form.

### Activity 4.1

- Make a list of ten things you have used or consumed since the morning.
- Compile this list with the lists made by your classmates and then sort the items into the adjacent Table.
- If there are items which are made up of more than one material, put them into both the relevant columns of the table.

Things made of metal	Things made of glass/clay	Others

Look at the items that come in the last column of the above table filled by you – your teacher will be able to tell you that most of them are made up of compounds of carbon. Can you think of a method to test this? What would be the product if a compound containing carbon is burnt? Do you know of any test to confirm this?

Food, clothes, medicines, books, or many of the things that you listed are all based on this versatile element carbon. In addition, all living structures are carbon based. The amount of carbon present in the earth's crust and in the atmosphere is quite meagre. The earth's crust has only 0.02% carbon in the form of minerals (like carbonates, hydrogen-carbonates, coal and petroleum) and the atmosphere has 0.03% of carbon dioxide. In spite of this small amount of carbon available in nature, the importance of carbon seems to be immense. In this Chapter, we will know about the properties of carbon which make carbon so important to us.

### 4.1 BONDING IN CARBON – THE COVALENT BOND

In the previous Chapter, we have studied the properties of ionic compounds. We saw that ionic compounds have high melting and boiling points and conduct electricity in solution or in the molten state. We also

saw how the nature of bonding in ionic compounds explains these properties. Let us now study the properties of some carbon compounds.

Most carbon compounds are poor conductors of electricity as we have seen in Chapter 2. From the data given in Table 4.1 on the boiling and melting points of the carbon compounds, we find that these compounds have low melting and boiling points as compared to ionic compounds (Chapter 3). We can conclude that the forces of attraction between the molecules are not very strong. Since these compounds are largely non-conductors of electricity, we can conclude that the bonding in these compounds does not give rise to any ions.

In Class IX, we learnt about the combining capacity of various elements and how it depends on the number of valence electrons. Let us now look at the electronic configuration of carbon. The atomic number of carbon is 6. What would be the distribution of electrons in various shells of carbon? How many valence electrons will carbon have?

We know that the reactivity of elements is explained as their tendency to attain a completely filled outer shell, that is, attain noble gas configuration. Elements forming ionic compounds achieve this by either gaining or losing electrons from the outermost shell. In the case of carbon, it has four electrons in its outermost shell and needs to gain or lose four electrons to attain noble gas configuration. If it were to gain or lose electrons –

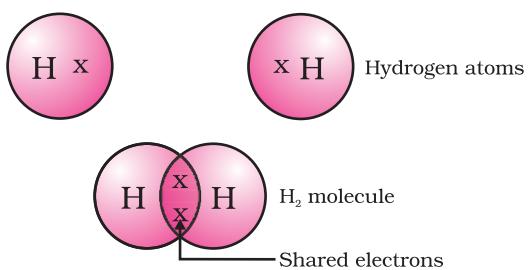
- It could gain four electrons forming  $C^{4-}$  anion. But it would be difficult for the nucleus with six protons to hold on to ten electrons, that is, four extra electrons.
- It could lose four electrons forming  $C^{4+}$  cation. But it would require a large amount of energy to remove four electrons leaving behind a carbon cation with six protons in its nucleus holding on to just two electrons.

Carbon overcomes this problem by sharing its valence electrons with other atoms of carbon or with atoms of other elements. Not just carbon, but many other elements form molecules by sharing electrons in this manner. The shared electrons 'belong' to the outermost shells of both the atoms and lead to both atoms attaining the noble gas configuration. Before going on to compounds of carbon, let us look at some simple molecules formed by the sharing of valence electrons.

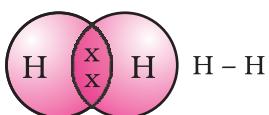
The simplest molecule formed in this manner is that of hydrogen. As you have learnt earlier, the atomic number of hydrogen is 1. Hence hydrogen has one electron in its K shell and it requires one more electron to fill the K shell. So two hydrogen atoms share their electrons to form a molecule of hydrogen,  $H_2$ . This allows each hydrogen atom to attain the

**Table 4.1** Melting points and boiling points of some compounds of carbon

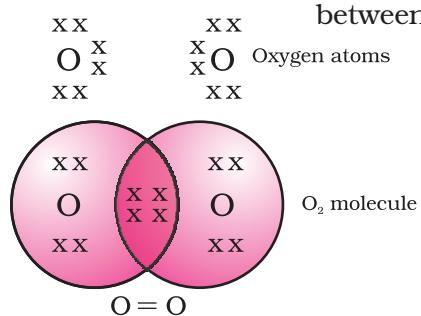
Compound	Melting point (K)	Boiling point (K)
Acetic acid ( $CH_3COOH$ )	290	391
Chloroform ( $CHCl_3$ )	209	334
Ethanol ( $CH_3CH_2OH$ )	156	351
Methane ( $CH_4$ )	90	111



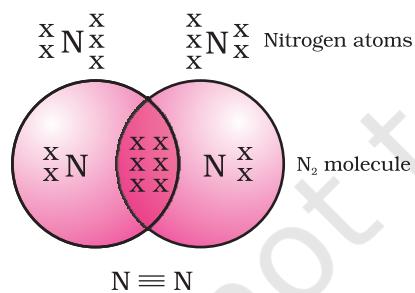
**Figure 4.1**  
A molecule of hydrogen



**Figure 4.2**  
Single bond between two hydrogen atoms



**Figure 4.3**  
Double bond between two oxygen atoms



**Figure 4.4**  
Triple bond between two nitrogen atoms

electronic configuration of the nearest noble gas, helium, which has two electrons in its K shell. We can depict this using dots or crosses to represent valence electrons (Fig. 4.1).

The shared pair of electrons is said to constitute a single covalent bond between the two hydrogen atoms. A single covalent bond is also represented by a line between the two atoms, as shown in Fig. 4.2.

The atomic number of chlorine is 17. What would be its electronic configuration and its valency? Chlorine forms a diatomic molecule,  $\text{Cl}_2$ . Can you draw the electron dot structure for this molecule? Note that only the valence shell electrons need to be depicted.

In the case of oxygen, we see the formation of a double bond between two oxygen atoms. This is because an atom of oxygen has six electrons in its L shell (the atomic number of oxygen is eight) and it requires two more electrons to complete its octet. So each atom of oxygen shares two electrons with another atom of oxygen to give us the structure shown in Fig. 4.3. The two electrons contributed by each oxygen atom give rise to two shared pairs of electrons. This is said to constitute a double bond between the two atoms.

Can you now depict a molecule of water showing the nature of bonding between one oxygen atom and two hydrogen atoms? Does the molecule have single bonds or double bonds?

What would happen in the case of a diatomic molecule of nitrogen? Nitrogen has the atomic number 7. What would be its electronic configuration and its combining capacity? In order to attain an octet, each nitrogen atom in a molecule of nitrogen contributes three electrons giving rise to three shared pairs of electrons. This is said to constitute a triple bond between the two atoms. The electron dot structure of  $\text{N}_2$  and its triple bond can be depicted as in Fig. 4.4.

A molecule of ammonia has the formula  $\text{NH}_3$ . Can you draw the electron dot structure for this molecule showing how all four atoms achieve noble gas configuration? Will the molecule have single, double or triple bonds?

Let us now take a look at methane, which is a compound of carbon. Methane is widely used as a fuel and is a major component of bio-gas and Compressed Natural Gas (CNG). It is also one of the simplest compounds formed by carbon. Methane has a formula  $\text{CH}_4$ . Hydrogen, as you know, has a valency of 1. Carbon is tetravalent because it has four valence electrons. In order to achieve noble gas configuration, carbon shares these electrons with four atoms of hydrogen as shown in Fig. 4.5.

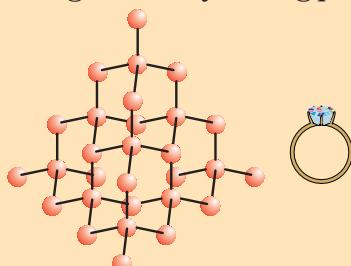
Such bonds which are formed by the sharing of an electron pair between two atoms are known as covalent bonds. Covalently bonded molecules are seen to have strong bonds within the molecule, but intermolecular forces are weak. This gives rise to the low melting and boiling

points of these compounds. Since the electrons are shared between atoms and no charged particles are formed, such covalent compounds are generally poor conductors of electricity.

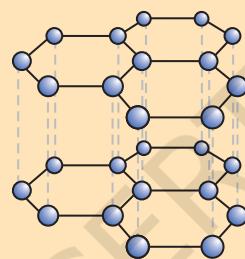


### Allotropes of carbon

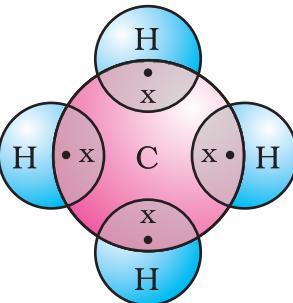
The element carbon occurs in different forms in nature with widely varying physical properties. Both diamond and graphite are formed by carbon atoms, the difference lies in the manner in which the carbon atoms are bonded to one another. In diamond, each carbon atom is bonded to four other carbon atoms forming a rigid three-dimensional structure. In graphite, each carbon atom is bonded to three other carbon atoms in the same plane giving a hexagonal array. One of these bonds is a double-bond, and thus the valency of carbon is satisfied. Graphite structure is formed by the hexagonal arrays being placed in layers one above the other.



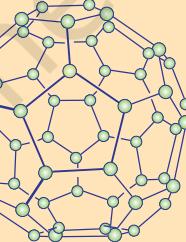
The structure of diamond



The structure of graphite



**Figure 4.5**  
Electron dot structure for methane



The structure of C-60  
Buckminsterfullerene

### More to Know!

These two different structures result in diamond and graphite having very different physical properties even though their chemical properties are the same. Diamond is the hardest substance known while graphite is smooth and slippery. Graphite is also a very good conductor of electricity unlike other non-metals that you studied in the previous Chapter.

Diamonds can be synthesised by subjecting pure carbon to very high pressure and temperature. These synthetic diamonds are small but are otherwise indistinguishable from natural diamonds.

Fullerenes form another class of carbon allotropes. The first one to be identified was C-60 which has carbon atoms arranged in the shape of a football. Since this looked like the geodesic dome designed by the US architect Buckminster Fuller, the molecule was named fullerene.

## Q U E S T I O N S

- What would be the electron dot structure of carbon dioxide which has the formula  $\text{CO}_2$ ?
- What would be the electron dot structure of a molecule of sulphur which is made up of eight atoms of sulphur? (**Hint** – The eight atoms of sulphur are joined together in the form of a ring.)



## 4.2 VERSATILE NATURE OF CARBON

We have seen the formation of covalent bonds by the sharing of electrons in various elements and compounds. We have also seen the structure of a simple carbon compound, methane. In the beginning of the Chapter, we saw how many things we use contain carbon. In fact, we ourselves are made up of carbon compounds. The numbers of carbon compounds whose formulae are known to chemists was recently estimated to be in millions! This outnumbers by a large margin the compounds formed by all the other elements put together. Why is it that this property is seen in carbon and no other element? The nature of the covalent bond enables carbon to form a large number of compounds. Two factors noticed in the case of carbon are –

- (i) Carbon has the unique ability to form bonds with other atoms of carbon, giving rise to large molecules. This property is called catenation. These compounds may have long chains of carbon, branched chains of carbon or even carbon atoms arranged in rings. In addition, carbon atoms may be linked by single, double or triple bonds. Compounds of carbon, which are linked by only single bonds between the carbon atoms are called saturated compounds. Compounds of carbon having double or triple bonds between their carbon atoms are called unsaturated compounds.

No other element exhibits the property of catenation to the extent seen in carbon compounds. Silicon forms compounds with hydrogen which have chains of upto seven or eight atoms, but these compounds are very reactive. The carbon-carbon bond is very strong and hence stable. This gives us the large number of compounds with many carbon atoms linked to each other.

- (ii) Since carbon has a valency of four, it is capable of bonding with four other atoms of carbon or atoms of some other mono-valent element. Compounds of carbon are formed with oxygen, hydrogen, nitrogen, sulphur, chlorine and many other elements giving rise to compounds with specific properties which depend on the elements other than carbon present in the molecule.

Again the bonds that carbon forms with most other elements are very strong making these compounds exceptionally stable. One reason for the formation of strong bonds by carbon is its small size. This enables the nucleus to hold on to the shared pairs of electrons strongly. The bonds formed by elements having bigger atoms are much weaker.

### Organic compounds

The two characteristic features seen in carbon, that is, tetravalency and catenation, put together give rise to a large number of compounds. Many have the same non-carbon atom or group of atoms attached to different carbon chains. These compounds were initially extracted from natural substances and it was thought that these carbon compounds or organic compounds could only be formed within a living system. That is, it was postulated that a 'vital force' was necessary for their synthesis. Friedrich Wöhler disproved this in 1828 by preparing urea from ammonium cyanate. But carbon compounds, except for carbides, oxides of carbon, carbonate and hydrogencarbonate salts continue to be studied under organic chemistry.

#### 4.2.1 Saturated and Unsaturated Carbon Compounds

We have already seen the structure of methane. Another compound formed between carbon and hydrogen is ethane with a formula of  $C_2H_6$ . In order to arrive at the structure of simple carbon compounds, the first step is to link the carbon atoms together with a single bond (Fig. 4.6a) and then use the hydrogen atoms to satisfy the remaining valencies of carbon (Fig. 4.6b). For example, the structure of ethane is arrived in the following steps –



Step 1

**Figure 4.6 (a)** Carbon atoms linked together with a single bond

Three valencies of each carbon atom remain unsatisfied, so each is bonded to three hydrogen atoms giving:



Step 2

**Figure 4.6 (b)** Each carbon atom bonded to three hydrogen atoms

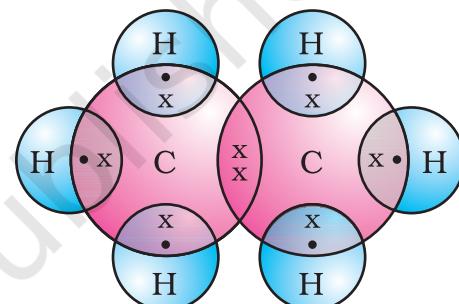
The electron dot structure of ethane is shown in Fig. 4.6(c).

Can you draw the structure of propane, which has the molecular formula  $C_3H_8$  in a similar manner? You will see that the valencies of all the atoms are satisfied by single bonds between them. Such carbon compounds are called saturated compounds. These compounds are normally not very reactive.

However, another compound of carbon and hydrogen has the formula  $C_2H_4$  and is called ethene. How can this molecule be depicted? We follow the same step-wise approach as above.

Carbon–carbon atoms linked together with a single bond (Step 1).

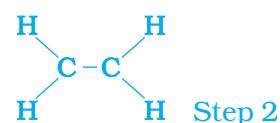
We see that one valency per carbon atom remains unsatisfied (Step 2). This can be satisfied only if there is a double bond between the two carbons (Step 3).



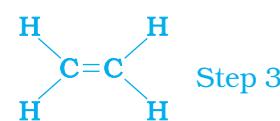
**Figure 4.6 (c)** Electron dot structure of ethane



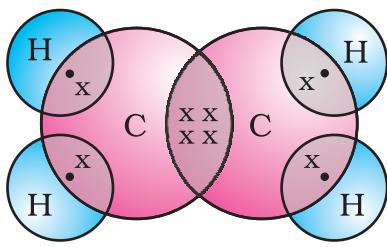
Step 1



Step 2



Step 3



**Figure 4.7**  
Structure of ethene

The electron dot structure for ethene is given in Fig. 4.7. Yet another compound of hydrogen and carbon has the formula  $C_2H_2$  and is called ethyne. Can you draw the electron dot structure for ethyne? How many bonds are necessary between the two carbon atoms in order to satisfy their valencies? Such compounds of carbon having double or triple bonds between the carbon atoms are known as unsaturated carbon compounds and they are more reactive than the saturated carbon compounds.

## 4.2.2 Chains, Branches and Rings

In the earlier section, we mentioned the carbon compounds methane, ethane and propane, containing respectively 1, 2 and 3 carbon atoms. Such 'chains' of carbon atoms can contain many more carbon atoms. The names and structures of six of these are given in Table 4.2.

**Table 4.2** Formulae and structures of saturated compounds of carbon and hydrogen

No. of C atoms	Name	Formula	Structure
1	Methane	$CH_4$	<pre>       H               H-C-H               H     </pre>
2	Ethane	$C_2H_6$	<pre>       H   H                   H-C-C-H                   H   H     </pre>
3	Propane	$C_3H_8$	<pre>       H   H   H                       H-C-C-C-H                       H   H   H     </pre>
4	Butane	$C_4H_{10}$	<pre>       H   H   H   H                           H-C-C-C-C-H                       H   H   H     </pre>
5	Pentane	$C_5H_{12}$	<pre>       H   H   H   H   H                           H-C-C-C-C-C-H                       H   H   H     </pre>
6	Hexane	$C_6H_{14}$	<pre>       H   H   H   H   H   H                           H-C-C-C-C-C-C-H                       H   H   H     </pre>

But, let us take another look at butane. If we make the carbon 'skeleton' with four carbon atoms, we see that two different possible 'skeletons' are –



**Figure 4.8 (a)** Two possible carbon-skeletons

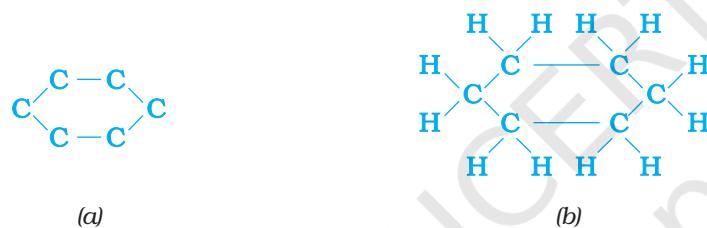
Filling the remaining valencies with hydrogen gives us –



**Figure 4.8 (b)** Complete molecules for two structures with formula  $C_4H_{10}$

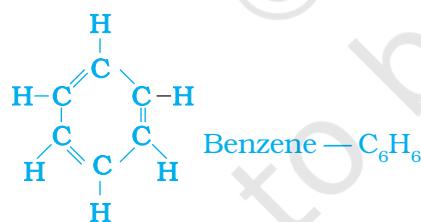
We see that both these structures have the same formula  $C_4H_{10}$ . Such compounds with identical molecular formula but different structures are called structural isomers.

In addition to straight and branched carbon chains, some compounds have carbon atoms arranged in the form of a ring. For example, cyclohexane has the formula  $C_6H_{12}$  and the following structure –



**Figure 4.9** Structure of cyclohexane (a) carbon skeleton (b) complete molecule

Can you draw the electron dot structure for cyclohexane? Straight chain, branched chain and cyclic carbon compounds, all may be saturated or unsaturated. For example, benzene,  $C_6H_6$ , has the following structure –



**Figure 4.10** Structure of benzene

All these carbon compounds which contain only carbon and hydrogen are called hydrocarbons. Among these, the saturated hydrocarbons are called alkanes. The unsaturated hydrocarbons which contain one or more double bonds are called alkenes. Those containing one or more triple bonds are called alkynes.

### 4.2.3 Will you be my Friend?

Carbon seems to be a very friendly element. So far we have been looking at compounds containing carbon and hydrogen only. But carbon also forms

bonds with other elements such as halogens, oxygen, nitrogen and sulphur. In a hydrocarbon chain, one or more hydrogens can be replaced by these elements, such that the valency of carbon remains satisfied. In such compounds, the element replacing hydrogen is referred to as a heteroatom. These heteroatoms are also present in some groups as given in Table 4.3.

**Table 4.3** Some functional groups in carbon compounds

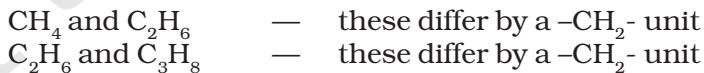
Hetero atom	Class of compounds	Formula of functional group
Cl/Br	Halo- (Chloro/bromo) alkane	$\text{—Cl}$ , $\text{—Br}$ (substitutes for hydrogen atom)
Oxygen	1. Alcohol	$\text{—OH}$
	2. Aldehyde	$\begin{array}{c} \text{H} \\   \\ \text{—C=O} \end{array}$
	3. Ketone	$\begin{array}{c} \text{—C=O} \\   \\ \text{—C—} \end{array}$
	4. Carboxylic acid	$\begin{array}{c} \text{O} \\    \\ \text{—C—OH} \end{array}$

These heteroatoms and the group containing these confer specific properties to the compound, regardless of the length and nature of the carbon chain and hence are called functional groups. Some important functional groups are given in the Table 4.3. Free valency or valencies of the group are shown by the single line. The functional group is attached to the carbon chain through this valency by replacing one hydrogen atom or atoms.

#### 4.2.4 Homologous Series

You have seen that carbon atoms can be linked together to form chains of varying lengths. These chains can be branched also. In addition, hydrogen atom or other atoms on these carbon chains can be replaced by any of the functional groups that we saw above. The presence of a functional group such as alcohol decides the properties of the carbon compound, regardless of the length of the carbon chain. For example, the chemical properties of  $\text{CH}_3\text{OH}$ ,  $\text{C}_2\text{H}_5\text{OH}$ ,  $\text{C}_3\text{H}_7\text{OH}$  and  $\text{C}_4\text{H}_9\text{OH}$  are all very similar. Hence, such a series of compounds in which the same functional group substitutes for hydrogen in a carbon chain is called a homologous series.

Let us look at the homologous series that we saw earlier in Table 4.2. If we look at the formulae of successive compounds, say –



What is the difference between the next pair – propane and butane ( $\text{C}_4\text{H}_{10}$ )?

Can you find out the difference in molecular masses between these pairs (the atomic mass of carbon is 12 u and the atomic mass of hydrogen is 1 u)?

Similarly, take the homologous series for alkenes. The first member of the series is ethene which we have already come across in Section 4.2.1. What is the formula for ethene? The succeeding members have the formula  $\text{C}_3\text{H}_6$ ,  $\text{C}_4\text{H}_8$  and  $\text{C}_5\text{H}_{10}$ . Do these also differ by a  $-\text{CH}_2-$

unit? Do you see any relation between the number of carbon and hydrogen atoms in these compounds? The general formula for alkenes can be written as  $C_nH_{2n}$ , where  $n = 2, 3, 4$ . Can you similarly generate the general formula for alkanes and alkynes?

As the molecular mass increases in any homologous series, a gradation in physical properties is seen. This is because the melting and boiling points increase with increasing molecular mass. Other physical properties such as solubility in a particular solvent also show a similar gradation. But the chemical properties, which are determined solely by the functional group, remain similar in a homologous series.

### Activity 4.2

- Calculate the difference in the formulae and molecular masses for (a)  $CH_3OH$  and  $C_2H_5OH$  (b)  $C_2H_5OH$  and  $C_3H_7OH$ , and (c)  $C_3H_7OH$  and  $C_4H_9OH$ .
- Is there any similarity in these three?
- Arrange these alcohols in the order of increasing carbon atoms to get a family. Can we call this family a homologous series?
- Generate the homologous series for compounds containing up to four carbons for the other functional groups given in Table 4.3.

#### 4.2.5 Nomenclature of Carbon Compounds

The names of compounds in a homologous series are based on the name of the basic carbon chain modified by a “prefix” “phrase before” or “suffix” “phrase after” indicating the nature of the functional group. For example, the names of the alcohols taken in Activity 4.2 are methanol, ethanol, propanol and butanol.

Naming a carbon compound can be done by the following method –

- (i) Identify the number of carbon atoms in the compound. A compound having three carbon atoms would have the name propane.
- (ii) In case a functional group is present, it is indicated in the name of the compound with either a prefix or a suffix (as given in Table 4.4).
- (iii) If the name of the functional group is to be given as a suffix, and the suffix of the functional group begins with a vowel a, e, i, o, u, then the name of the carbon chain is modified by deleting the final ‘e’ and adding the appropriate suffix. For example, a three-carbon chain with a ketone group would be named in the following manner –  
Propane – ‘e’ = propan + ‘one’ = propanone.
- (iv) If the carbon chain is unsaturated, then the final ‘ane’ in the name of the carbon chain is substituted by ‘ene’ or ‘yne’ as given in Table 4.4. For example, a three-carbon chain with a double bond would be called propene and if it has a triple bond, it would be called propyne.

**Table 4.4** Nomenclature of organic compounds

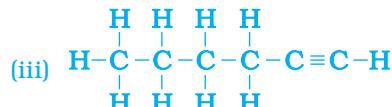
Class of compounds	Prefix/Suffix	Example
1. Halo alkane	Prefix-chloro, bromo, etc.	$  \begin{array}{c} \text{H} & \text{H} & \text{H} \\   &   &   \\ \text{H}-\text{C} & -\text{C} & -\text{C}-\text{Cl} \\   &   &   \\ \text{H} & \text{H} & \text{H} \end{array}  $ Chloropropane
		$  \begin{array}{c} \text{H} & \text{H} & \text{H} \\   &   &   \\ \text{H}-\text{C} & -\text{C} & -\text{C}-\text{Br} \\   &   &   \\ \text{H} & \text{H} & \text{H} \end{array}  $ Bromopropane
2. Alcohol	Suffix - ol	$  \begin{array}{c} \text{H} & \text{H} & \text{H} \\   &   &   \\ \text{H}-\text{C} & -\text{C} & -\text{C}-\text{OH} \\   &   &   \\ \text{H} & \text{H} & \text{H} \end{array}  $ Propanol
3. Aldehyde	Suffix - al	$  \begin{array}{c} \text{H} & \text{H} & \text{H} \\   &   &   \\ \text{H}-\text{C} & -\text{C} & -\text{C}=\text{O} \\   &   &   \\ \text{H} & \text{H} & \text{H} \end{array}  $ Propanal
4. Ketone	Suffix - one	$  \begin{array}{c} \text{H} & \text{H} & \text{H} \\   &   &   \\ \text{H}-\text{C} & -\text{C} & -\text{C}-\text{H} \\    & &   \\ \text{H} & \text{O} & \text{H} \end{array}  $ Propanone
5. Carboxylic acid	Suffix - oic acid	$  \begin{array}{c} \text{H} & \text{H} & \text{O} \\   &   &    \\ \text{H}-\text{C} & -\text{C} & -\text{C}-\text{OH} \\   &   &   \\ \text{H} & \text{H} & \text{H} \end{array}  $ Propanoic acid
6. Alkenes	Suffix - ene	$  \begin{array}{c} \text{H} & \text{H} \\   &   \\ \text{H}-\text{C} & -\text{C}=\text{C} \\   & / \\ \text{H} & \text{H} \end{array}  $ Propene
7. Alkynes	Suffix - yne	$  \begin{array}{c} \text{H} \\   \\ \text{H}-\text{C} & -\text{C}\equiv\text{C}-\text{H} \\   \\ \text{H} \end{array}  $ Propyne

## Q U E S T I O N S

- How many structural isomers can you draw for pentane?
- What are the two properties of carbon which lead to the huge number of carbon compounds we see around us?
- What will be the formula and electron dot structure of cyclopentane?

\*Are structural isomers possible for bromopentane?

5. How would you name the following compounds?



## 4.3 CHEMICAL PROPERTIES OF CARBON COMPOUNDS

In this section we will be studying about some of the chemical properties of carbon compounds. Since most of the fuels we use are either carbon or its compounds, we shall first study combustion.

### 4.3.1 Combustion

Carbon, in all its allotropic forms, burns in oxygen to give carbon dioxide along with the release of heat and light. Most carbon compounds also release a large amount of heat and light on burning. These are the oxidation reactions that you learnt about in the first Chapter –

- (i)  $\text{C} + \text{O}_2 \rightarrow \text{CO}_2 + \text{heat and light}$
  - (ii)  $\text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} + \text{heat and light}$
  - (iii)  $\text{CH}_3\text{CH}_2\text{OH} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} + \text{heat and light}$

Balance the latter two reactions like you learnt in the first Chapter.

## Activity 4.3

**CAUTION:** This Activity needs the teacher's assistance.

- Take some carbon compounds (naphthalene, camphor, alcohol) one by one on a spatula and burn them.
  - Observe the nature of the flame and note whether smoke is produced.
  - Place a metal plate above the flame. Is there a deposition on the plate in case of any of the compounds?

## Activity 4.4

- Light a bunsen burner and adjust the air hole at the base to get different types of flames/presence of smoke.
  - When do you get a yellow, sooty flame?
  - When do you get a blue flame?

Saturated hydrocarbons will generally give a clean flame while unsaturated carbon compounds will give a yellow flame with lots of black smoke. This results in a sooty deposit on the metal plate in Activity 4.3. However, limiting the supply of air results in incomplete combustion of even saturated hydrocarbons giving a sooty flame. The gas/kerosene stove used at home has inlets for air so that a sufficiently oxygen-rich

mixture is burnt to give a clean blue flame. If you observe the bottoms of cooking vessels getting blackened, it means that the air holes are blocked and fuel is getting wasted. Fuels such as coal and petroleum have some amount of nitrogen and sulphur in them. Their combustion results in the formation of oxides of sulphur and nitrogen which are major pollutants in the environment.

### Why do substances burn with or without a flame?

Have you ever observed either a coal or a wood fire? If not, the next time you get a chance, take close note of what happens when the wood or coal starts to burn. You have seen above that a candle or the LPG in the gas stove burns with a flame. However, you will observe the coal or charcoal in an ‘angithi’ sometimes just glows red and gives out heat without a flame. This is because a flame is only produced when gaseous substances burn. When wood or charcoal is ignited, the volatile substances present vapourise and burn with a flame in the beginning.

A luminous flame is seen when the atoms of the gaseous substance are heated and start to glow. The colour produced by each element is a characteristic property of that element. Try and heat a copper wire in the flame of a gas stove and observe its colour. You have seen that incomplete combustion gives soot which is carbon. On this basis, what will you attribute the yellow colour of a candle flame to?

### Formation of coal and petroleum

Coal and petroleum have been formed from biomass which has been subjected to various biological and geological processes. Coal is the remains of trees, ferns, and other plants that lived millions of years ago. These were crushed into the earth, perhaps by earthquakes or volcanic eruptions. They were pressed down by layers of earth and rock. They slowly decayed into coal. Oil and gas are the remains of millions of tiny plants and animals that lived in the sea. When they died, their bodies sank to the sea bed and were covered by silt. Bacteria attacked the dead remains, turning them into oil and gas under the high pressures they were being subjected to. Meanwhile, the silt was slowly compressed into rock. The oil and gas seeped into the porous parts of the rock, and got trapped like water in a sponge. Can you guess why coal and petroleum are called fossil fuels?

### 4.3.2 Oxidation

#### Activity 4.5

- Take about 3 mL of ethanol in a test tube and warm it gently in a water bath.
- Add a 5% solution of alkaline potassium permanganate drop by drop to this solution.
- Does the colour of potassium permanganate persist when it is added initially?
- Why does the colour of potassium permanganate not disappear when excess is added?

You have learnt about oxidation reactions in the first Chapter. Carbon compounds can be easily oxidised on combustion. In addition to this complete oxidation, we have reactions in which alcohols are converted to carboxylic acids –

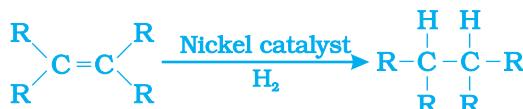


We see that some substances are capable of adding oxygen to others. These substances are known as oxidising agents.

Alkaline potassium permanganate or acidified potassium dichromate are oxidising alcohols to acids, that is, adding oxygen to the starting material. Hence they are known as oxidising agents.

#### 4.3.3 Addition Reaction

Unsaturated hydrocarbons add hydrogen in the presence of catalysts such as palladium or nickel to give saturated hydrocarbons. Catalysts are substances that cause a reaction to occur or proceed at a different rate without the reaction itself being affected. This reaction is commonly used in the hydrogenation of vegetable oils using a nickel catalyst. Vegetable oils generally have long unsaturated carbon chains while animal fats have saturated carbon chains.



You must have seen advertisements stating that some vegetable oils are 'healthy'. Animal fats generally contain saturated fatty acids which are said to be harmful for health. Oils containing unsaturated fatty acids should be chosen for cooking.

#### 4.3.4 Substitution Reaction

Saturated hydrocarbons are fairly unreactive and are inert in the presence of most reagents. However, in the presence of sunlight, chlorine is added to hydrocarbons in a very fast reaction. Chlorine can replace the hydrogen atoms one by one. It is called a substitution reaction because one type of atom or a group of atoms takes the place of another. A number of products are usually formed with the higher homologues of alkanes.



### Q U E S T I O N S

1. Why is the conversion of ethanol to ethanoic acid an oxidation reaction?
2. A mixture of oxygen and ethyne is burnt for welding. Can you tell why a mixture of ethyne and air is not used?



#### 4.4 SOME IMPORTANT CARBON COMPOUNDS – ETHANOL AND ETHANOIC ACID

Many carbon compounds are invaluable to us. But here we shall study the properties of two commercially important compounds – ethanol and ethanoic acid.

#### 4.4.1 Properties of Ethanol

Ethanol is a liquid at room temperature (refer to Table 4.1 for the melting and boiling points of ethanol). Ethanol is commonly called alcohol and is the active ingredient of all alcoholic drinks. In addition, because it is a good solvent, it is also used in medicines such as tincture iodine, cough syrups, and many tonics. Ethanol is also soluble in water in all proportions. Consumption of small quantities of dilute ethanol causes drunkenness. Even though this practice is condemned, it is a socially widespread practice. However, intake of even a small quantity of pure ethanol (called absolute alcohol) can be lethal. Also, long-term consumption of alcohol leads to many health problems.

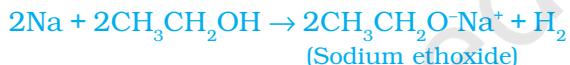
##### **Reactions of Ethanol**

- (i) Reaction with sodium –

#### Activity 4.6

Teacher's demonstration –

- Drop a small piece of sodium, about the size of a couple of grains of rice, into ethanol (absolute alcohol).
- What do you observe?
- How will you test the gas evolved?



Alcohols react with sodium leading to the evolution of hydrogen. With ethanol, the other product is sodium ethoxide. Can you recall which other substances produce hydrogen on reacting with metals?

- (ii) Reaction to give unsaturated hydrocarbon: Heating ethanol at 443 K with excess concentrated sulphuric acid results in the dehydration of ethanol to give ethene –



The concentrated sulphuric acid can be regarded as a dehydrating agent which removes water from ethanol.

#### How do alcohols affect living beings?

When large quantities of ethanol are consumed, it tends to slow metabolic processes and to depress the central nervous system. This results in lack of coordination, mental confusion, drowsiness, lowering of the normal inhibitions, and finally stupor. The individual may feel relaxed without realising that his sense of judgement, sense of timing, and muscular coordination have been seriously impaired.

Unlike ethanol, intake of methanol in very small quantities can cause death. Methanol is oxidised to methanal in the liver. Methanal reacts rapidly with the components of cells. It coagulates the protoplasm, in much the same way an egg is coagulated by cooking. Methanol also affects the optic nerve, causing blindness.

Ethanol is an important industrial solvent. To prevent the misuse of ethanol produced for industrial use, it is made unfit for drinking by adding poisonous substances like methanol to it. Dyes are also added to colour the alcohol blue so that it can be identified easily. This is called denatured alcohol.

## Alcohol as a fuel

Sugarcane plants are one of the most efficient converters of sunlight into chemical energy. Sugarcane juice can be used to prepare molasses which is fermented to give alcohol (ethanol). Some countries now use alcohol as an additive in petrol since it is a cleaner fuel which gives rise to only carbon dioxide and water on burning in sufficient air (oxygen).

## 4.4.2 Properties of Ethanoic Acid

Ethanoic acid is commonly called acetic acid and belongs to a group of acids called carboxylic acids. 5-8% solution of acetic acid in water is called vinegar and is used widely as a preservative in pickles. The melting point of pure ethanoic acid is 290 K and hence it often freezes during winter in cold climates. This gave rise to its name glacial acetic acid.

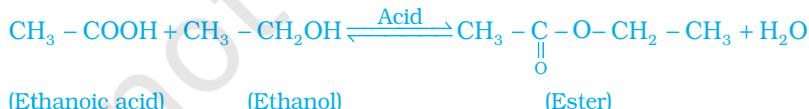
The group of organic compounds called carboxylic acids are obviously characterised by their acidic nature. However, unlike mineral acids like HCl, which are completely ionised, carboxylic acids are weak acids.

## Activity 4.8

- Take 1 mL ethanol (absolute alcohol) and 1 mL glacial acetic acid along with a few drops of concentrated sulphuric acid in a test tube.
  - Warm in a water-bath for at least five minutes as shown in Fig. 4.11.
  - Pour into a beaker containing 20-50 mL of water and smell the resulting mixture.

### Reactions of ethanoic acid:

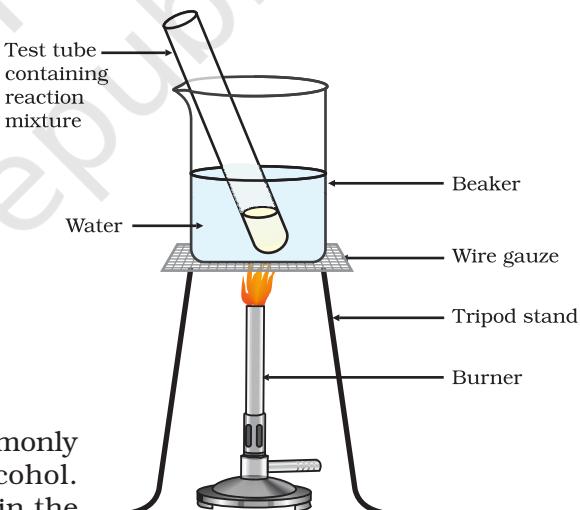
- (i) *Esterification reaction:* Esters are most commonly formed by reaction of an acid and an alcohol. Ethanoic acid reacts with absolute ethanol in the presence of an acid catalyst to give an ester –



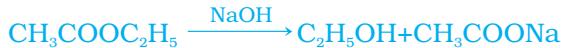
Generally, esters are sweet-smelling substances. These are used in making perfumes and as flavouring agents. On treating with sodium hydroxide, which is an alkali, the ester is converted back to alcohol and sodium salt of carboxylic acid. This reaction is known as saponification because it is used in the preparation of soap. Soaps are sodium or potassium salts of long chain carboxylic acid.

## Activity 4.7

- Compare the pH of dilute acetic acid and dilute hydrochloric acid using both litmus paper and universal indicator.
  - Are both acids indicated by the litmus test?
  - Does the universal indicator show them as equally strong acids?



**Figure 4.11**  
*Formation of ester*



- (ii) *Reaction with a base:* Like mineral acids, ethanoic acid reacts with a base such as sodium hydroxide to give a salt (sodium ethanoate or commonly called sodium acetate) and water:



### How does ethanoic acid react with carbonates and hydrogencarbonates?

Let us perform an activity to find out.

### Activity 4.9

- Set up the apparatus as shown in Chapter 2, Activity 2.5.
- Take a spatula full of sodium carbonate in a test tube and add 2 mL of dilute ethanoic acid.
- What do you observe?
- Pass the gas produced through freshly prepared lime-water. What do you observe?
- Can the gas produced by the reaction between ethanoic acid and sodium carbonate be identified by this test?
- Repeat this Activity with sodium hydrogencarbonate instead of sodium carbonate.

- (iii) *Reaction with carbonates and hydrogencarbonates:* Ethanoic acid reacts with carbonates and hydrogencarbonates to give rise to a salt, carbon dioxide and water. The salt produced is commonly called sodium acetate.

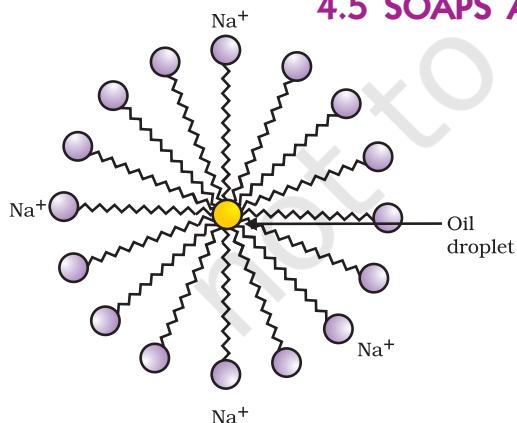


### QUESTIONS

1. How would you distinguish experimentally between an alcohol and a carboxylic acid?
2. What are oxidising agents?



### 4.5 SOAPS AND DETERGENTS



**Figure 4.12**  
Formation of micelles

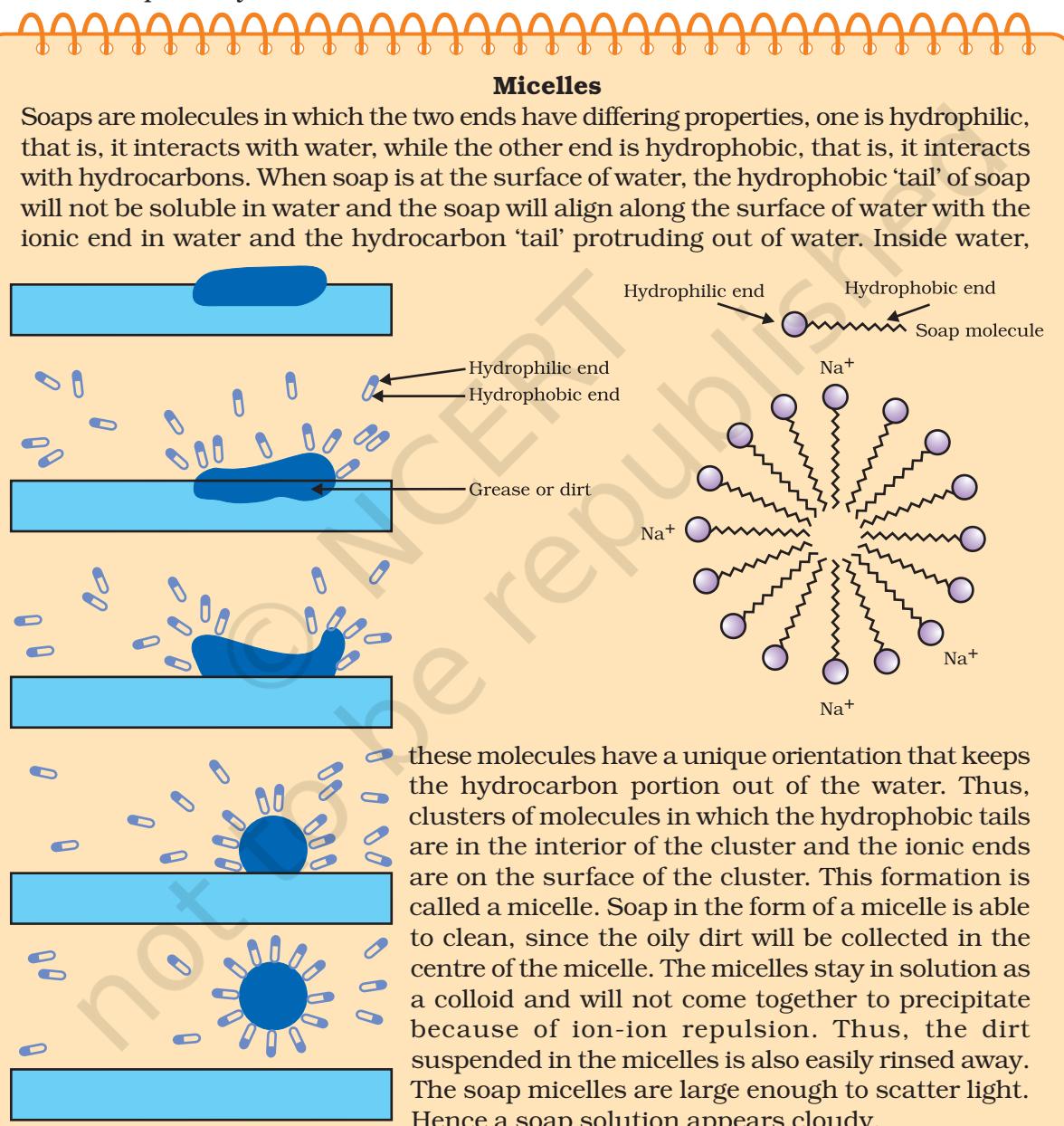
### Activity 4.10

- Take about 10 mL of water each in two test tubes.
- Add a drop of oil (cooking oil) to both the test tubes and label them as A and B.
- To test tube B, add a few drops of soap solution.
- Now shake both the test tubes vigorously for the same period of time.
- Can you see the oil and water layers separately in both the test tubes immediately after you stop shaking them?
- Leave the test tubes undisturbed for some time and observe. Does the oil layer separate out? In which test tube does this happen first?

This activity demonstrates the effect of soap in cleaning. Most dirt is oily in nature and as you know, oil does not dissolve in water. The molecules of soap are sodium or potassium salts of long-chain carboxylic acids. The ionic-end of soap interacts with water while the carbon chain interacts with oil. The soap molecules, thus form structures called micelles (see Fig. 4.12) where one end of the molecules is towards the oil droplet while the ionic-end faces outside. This forms an emulsion in water. The soap micelle thus helps in pulling out the dirt in water and we can wash our clothes clean (Fig. 4.13).

Can you draw the structure of the micelle that would be formed if you dissolve soap in a hydrocarbon?

## More to Know!



these molecules have a unique orientation that keeps the hydrocarbon portion out of the water. Thus, clusters of molecules in which the hydrophobic tails are in the interior of the cluster and the ionic ends are on the surface of the cluster. This formation is called a micelle. Soap in the form of a micelle is able to clean, since the oily dirt will be collected in the centre of the micelle. The micelles stay in solution as a colloid and will not come together to precipitate because of ion-ion repulsion. Thus, the dirt suspended in the micelles is also easily rinsed away. The soap micelles are large enough to scatter light. Hence a soap solution appears cloudy.

**Figure 4.13** Effect of soap in cleaning

## Activity 4.11

- Take about 10 mL of distilled water (or rain water) and 10 mL of hard water (from a tubewell or hand-pump) in separate test tubes.
- Add a couple of drops of soap solution to both.
- Shake the test tubes vigorously for an equal period of time and observe the amount of foam formed.
- In which test tube do you get more foam?
- In which test tube do you observe a white curdy precipitate?  
*Note for the teacher:* If hard water is not available in your locality, prepare some hard water by dissolving hydrogencarbonates/sulphates/chlorides of calcium or magnesium in water.

## Activity 4.12

- Take two test tubes with about 10 mL of hard water in each.
- Add five drops of soap solution to one and five drops of detergent solution to the other.
- Shake both test tubes for the same period.
- Do both test tubes have the same amount of foam?
- In which test tube is a curdy solid formed?

Have you ever observed while bathing that foam is formed with difficulty and an insoluble substance (scum) remains after washing with water? This is caused by the reaction of soap with the calcium and magnesium salts, which cause the hardness of water. Hence you need to use a larger amount of soap. This problem is overcome by using another class of compounds called detergents as cleansing agents. Detergents are generally sodium salts of sulphonic acids or ammonium salts with chlorides or bromides ions, etc. Both have long hydrocarbon chain. The charged ends of these compounds do not form insoluble precipitates with the calcium and magnesium ions in hard water. Thus, they remain effective in hard water. Detergents are usually used to make shampoos and products for cleaning clothes.

## Q U E S T I O N S

1. Would you be able to check if water is hard by using a detergent?
2. People use a variety of methods to wash clothes. Usually after adding the soap, they 'beat' the clothes on a stone, or beat it with a paddle, scrub with a brush or the mixture is agitated in a washing machine. Why is agitation necessary to get clean clothes?



## What you have learnt

- Carbon is a versatile element that forms the basis for all living organisms and many of the things we use.
- This large variety of compounds is formed by carbon because of its tetravalency and the property of catenation that it exhibits.
- Covalent bonds are formed by the sharing of electrons between two atoms so that both can achieve a completely filled outermost shell.
- Carbon forms covalent bonds with itself and other elements such as hydrogen, oxygen, sulphur, nitrogen and chlorine.
- Carbon also forms compounds containing double and triple bonds between carbon atoms. These carbon chains may be in the form of straight chains, branched chains or rings.
- The ability of carbon to form chains gives rise to a homologous series of compounds in which the same functional group is attached to carbon chains of different lengths.
- The functional groups such as alcohols, aldehydes, ketones and carboxylic acids bestow characteristic properties to the carbon compounds that contain them.
- Carbon and its compounds are some of our major sources of fuels.
- Ethanol and ethanoic acid are carbon compounds of importance in our daily lives.
- The action of soaps and detergents is based on the presence of both hydrophobic and hydrophilic groups in the molecule and this helps to emulsify the oily dirt and hence its removal.

## E X E R C I S E S

1. Ethane, with the molecular formula  $C_2H_6$  has
  - (a) 6 covalent bonds.
  - (b) 7 covalent bonds.
  - (c) 8 covalent bonds.
  - (d) 9 covalent bonds.
2. Butanone is a four-carbon compound with the functional group
  - (a) carboxylic acid.
  - (b) aldehyde.
  - (c) ketone.
  - (d) alcohol.
3. While cooking, if the bottom of the vessel is getting blackened on the outside, it means that
  - (a) the food is not cooked completely.
  - (b) the fuel is not burning completely.
  - (c) the fuel is wet.
  - (d) the fuel is burning completely.

4. Explain the nature of the covalent bond using the bond formation in  $\text{CH}_3\text{Cl}$ .
5. Draw the electron dot structures for
  - (a) ethanoic acid.
  - (b)  $\text{H}_2\text{S}$ .
  - (c) propanone.
  - (d)  $\text{F}_2$ .
6. What is an homologous series? Explain with an example.
7. How can ethanol and ethanoic acid be differentiated on the basis of their physical and chemical properties?
8. Why does micelle formation take place when soap is added to water? Will a micelle be formed in other solvents such as ethanol also?
9. Why are carbon and its compounds used as fuels for most applications?
10. Explain the formation of scum when hard water is treated with soap.
11. What change will you observe if you test soap with litmus paper (red and blue)?
12. What is hydrogenation? What is its industrial application?
13. Which of the following hydrocarbons undergo addition reactions:  
 $\text{C}_2\text{H}_6$ ,  $\text{C}_3\text{H}_8$ ,  $\text{C}_3\text{H}_6$ ,  $\text{C}_2\text{H}_2$  and  $\text{CH}_4$ .
14. Give a test that can be used to differentiate between saturated and unsaturated hydrocarbons.
15. Explain the mechanism of the cleaning action of soaps.

## Group Activity

- I Use molecular model kits to make models of the compounds you have learnt in this Chapter.
- II
  - Take about 20 mL of castor oil/cotton seed oil/linseed oil/soyabean oil in a beaker. Add 30 mL of 20 % sodium hydroxide solution. Heat the mixture with continuous stirring for a few minutes till the mixture thickens. Add 5-10 g of common salt to this. Stir the mixture well and allow it to cool.
  - You can cut out the soap in fancy shapes. You can also add perfume to the soap before it sets.